The Chemistry 221 Companion

Lab Manual, Problem Sets, Lecture Slides and Learning Resources

Dr. Michael A. Russell
Mt. Hood Community College
Fall 2024

Welcome to Chemistry 221!

My name is **Dr. Michael Russell** and I am pleased that you have decided to take Chemistry 221 with me this quarter. I look forward to an exciting term with you!

Here are some hints on how to get the most out of the ***Chemistry 221 Companion***:

- Glance over the **Table of Contents** that follows this introduction. The Table of Contents lists the respective page numbers for each of the sections.

- If you need it: information on **how to construct a graph** can be found in the lab section (with a Roman number "***I***" leading.) A handy **pictorial guide to common glassware**, a ScienceNotes.org **Periodic Table** and a **parts per thousand handout** follow shortly afterwards. The **labs** we will be performing this quarter follow, and **be sure to use the correct lab for your lab section** (section W1 (online) is different from sections 01 and H1 (face to face).)

- The **problem sets** and **Exam Prep worksheets** that we will use this quarter follow the lab section. They are listed with a Roman number "***II***".

- A printed version of the **Lecture slides** that will be covered this quarter can be found next. The Lecture notes use a Roman number "***III***" followed by the Chapter number, then the page number. For example, ***Page III-5-3*** would refer to a PowerPoint note (the "***III***") in Chapter ***5*** (the "***5***"), and the "***3***" refers to the ***third*** page of notes for Chapter 5.

- **Lecture handouts** follow the lecture slides and augment difficult concepts discussed in lecture. The numbering system is similar to the PowerPoint slides system but with a "**IV**". For example, ***Page IV-5-1*** would refer to a Lecture Handout (the "***IV***") in Chapter ***5*** (the "***5***"), and the "***1***" refers to the ***first*** page of lecture handouts for Chapter 5.

- Finally, the **Concept Guides** (which are useful worked examples relating to each of the chapters studied this quarter) might prove useful – they begin with a Roman number **V**. Also, **practice problem sets** (which include answers at the end; they begin with a Roman number **VI**), various **quizzes with answers** (which begin with a Roman number **VII**) and finally **sample quizzes and exams** (with answers, they start with a Roman number **VIII**) follow shortly after. Note that additional quiz and exam testing resources are available on the CH 221 website (http://mhchem.org/221/classroom/qe.htm).

If you have questions throughout the quarter, please do not hesitate to contact me using the contact information below. Good luck with your studying!

Peace,

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(503) 491-7348 - *phone*
AC 2568 - *office*

CHEMISTRY 221 TABLE OF CONTENTS

CH 221 Laboratory Experiments - *http://mhchem.org/lab221*

CH 221 Problem Sets and Exam Prep Worksheets - *http://mhchem.org/pset221*

CH 221 Lecture Slides - *http://mhchem.org/ho221*

CH 221 Handouts - *http://mhchem.org/ho221*

CH 221 Concept Guides - *http://mhchem.org/ho221*

CH 221 Practice Problem Sets and Quizzes *(with answers)* - *http://mhchem.org/221*

More sample quizzes and exams online - http://mhchem.org/221/classroom/qe.htm

Laboratory Safety & Etiquette

Safety is of utmost importance. Work in the laboratory should be a safe experience. It will be safe, however, only if certain safety precautions are followed without exception. Safety is up to you. Everyone working in the chemistry laboratories must follow the following rules. Your instructor will discuss specific safety precautions relevant to each experiment during the pre-lab lecture. Do not hesitate to consult with your instructor if you have questions regarding any safety precautions. Failure to observe laboratory safety rules and procedures may result in injury to you or to fellow students. Students who do not follow these safety rules (including proper attire) will be asked to leave the laboratory. Repeat offenders may be dropped from the course at the discretion of the instructor.

1. **Appropriate attire:** Appropriate protective clothing must be worn at all times while in the laboratory. It is a good idea not to wear your best clothing to lab since many chemicals can stain, bleach or generate holes in your clothing.
   a. **Safety goggles** approved by the chemistry department must be worn at all times, even if you are wearing prescription glasses. Contact lenses are not recommended in the lab. Various fumes may accumulate under the lenses and injure your eyes. You are responsible for bringing your own pair of safety goggles to lab each week. Students who borrow safety goggles from the instructor will have points deducted from their lab. Students who fail to wear their safety goggles will be reminded once and have points deducted. The second time a student is seen without safety goggles on during a lab period, the student will be asked to leave the laboratory.
   b. **Shirts** must cover the entire upper torso, including the midsection and upper chest area and should be long enough to tuck inside your pants. Cotton t-shirts are fine. Tank tops, scooped neck tops, leotards, sleeveless blouses and tops made of sheer material are not allowed.
   c. **Pants and skirts** must be at least knee length.
   d. **Shoes** must be flat-soled and cover the entire foot. Socks must be worn with shoes. Sandals, open-toe shoes and high heels are not permitted.
   e. **Long hair** (shoulder length and longer) and billowy clothing must be tied back while working in the lab.

2. **Food and Drink:** NO food or drink will be allowed in the laboratory. This includes coffee, water, candy and chewing gum.

3. Working in the laboratory without an instructor present is strictly forbidden. Students must work in instructional laboratories only during regularly scheduled lab periods and then only when supervised by a member of the faculty.

4. Do not perform any unauthorized experiments. If you have an idea for improving an experiment or for a new experiment, consult with your instructor.

5. Wash your hands after every experiment and each time your hands come in contact with chemicals.

6. **Scales:** Never weigh reagents or chemicals directly on a balance or scales. First weigh an empty container or weighing paper, then press tare or "re-zero" to set the mass reading to zero. Then add your reagent to the container or weighing paper and re-weigh. Balances are expensive! Clean any spills immediately! Replace caps on bottles and return to cart when complete.

7. **Fume hoods** should be used when performing experiments that generate an objectionable gas.

8. **Working With Chemicals:**
    a. Never smell or taste anything in the laboratory unless specifically directed by your instructor. Many chemicals are poisons. Use your hand to waft the odor to your nose.
    b. Always **read the label** on all chemical bottles and waste bottles. If you see the wrong chemical, you may have a serious explosion. If unsure, consult with your instructor.
    c. Do not take chemical bottles to your lab bench unless directed by your instructor. Pour the approximate amount you need from the bottle into a small container and take this to your bench.
    d. Always use a metal spatula or scoopula to transfer solid chemicals. Do not use your finger to transfer chemicals. This will directly expose you to the potential hazards of the chemical and might contaminate the remaining chemical in the container.
    e. Do not put excess reagent back into the original bottle. There is always a chance of contaminating the original sample. Ask your instructor how to properly dispose of excess chemicals.
    f. Do not put pipets directly in any reagent bottle. This might result in contamination of the remaining liquid in the bottle. Never mouth pipet any liquid in the lab.
    g. Keep the lids and caps on the chemical bottles. Put the lids back on as soon as you are finished dispensing the material. Many chemicals are sensitive to light or to moisture in the air.
    h. When diluting concentrated acids or bases, add the acid or base slowly into water. Never pour water into acid. The heat generated from adding water to a concentrated acid or base solution can cause the solution to splatter or shatter the glass.

9. **Waste:** Dispose chemical waste in designated containers. Only flush chemicals down the sink if instructed by your instructor. Never pour organic waste down the drain. The waste containers are in the hood for each experiment. Read labels on waste containers to be sure to dispose of waste in the proper container. Disposing waste in the wrong container can generate an unwanted (and unexpected) chemical reaction!

10. **Spills:** Clean up any spills immediately and dispose of the spilled material properly. Check with your instructor on the proper way to clean up any material that you spill.

11. **Chipped or broken glassware** should be thrown in the glass waste container. Report broken glassware to your instructor so that it can be replaced.

12. **Hot objects** will burn! Do not pick up hot objects with your fingers. Use tongs or hot pads. Hot glass will crack if run under cold water. Allow heated glass sufficient time to cool.

13. **Accidents and Emergencies:** Report all injuries and accidents, no matter how minor, to your instructor immediately. Know the location of the fire extinguishes, fire blankets, safety showers, and eyewash stations. Familiarize yourself with two different exits from the lab, in the event of an emergency situation. Accidents are usually minor, but it is best to be prepared for serious trouble.

14. **Be aware of your classmates!** Are they obeying the safety rules? A nearby accident may not hurt or harm him/her but may injure you!

15. Above all else, **ask the instructor if you have any safety related questions!**

Information on Graphs

Graphs: Whenever you create a graph for a chemistry lab, keep the following points in mind:

- If creating a graph by hand, use a large portion of the graph paper to create your graph; small graphs can easily misrepresent data and/or trendlines. If using a computer program like Microsoft Excel, use large graph sizes when creating lab reports (up to the size of one complete page.)
- Plan ahead! Make sure all the data points will fit on the graph but will not be too crowded together horizontally or vertically. Again, use as much of the graph paper as possible when constructing your graph.
- On the vertical axis, label the quantity that is being plotted (i.e. "Time") and put its units in parentheses (i.e. "(seconds)". Do the same on the horizontal axis.
- If you are drawing a **best-fit line** through the data points, do ***not*** connect the dots! Instead, draw a line which has some data points on each side of the line you are drawing... think of your line as an "average" of the data points.
- Never force a graph to go through the origin (i.e. at x=0 and y=0) unless expressly told to do so.
- Examine your graph: are there one or two points which are farther away from the line than the other points? If so, make sure you plotted them correctly.
- Use regression techniques to find the equation for the best fit for your data. **ALWAYS** include the regression equation with the graph itself.
- Linear regression equations should always be accompanied by the **correlation coefficient**, **r**, and not just R^2. To find r from R^2, take the square root of R^2. If the slope is negative, your r value will be negative as well.

An example graph follows:

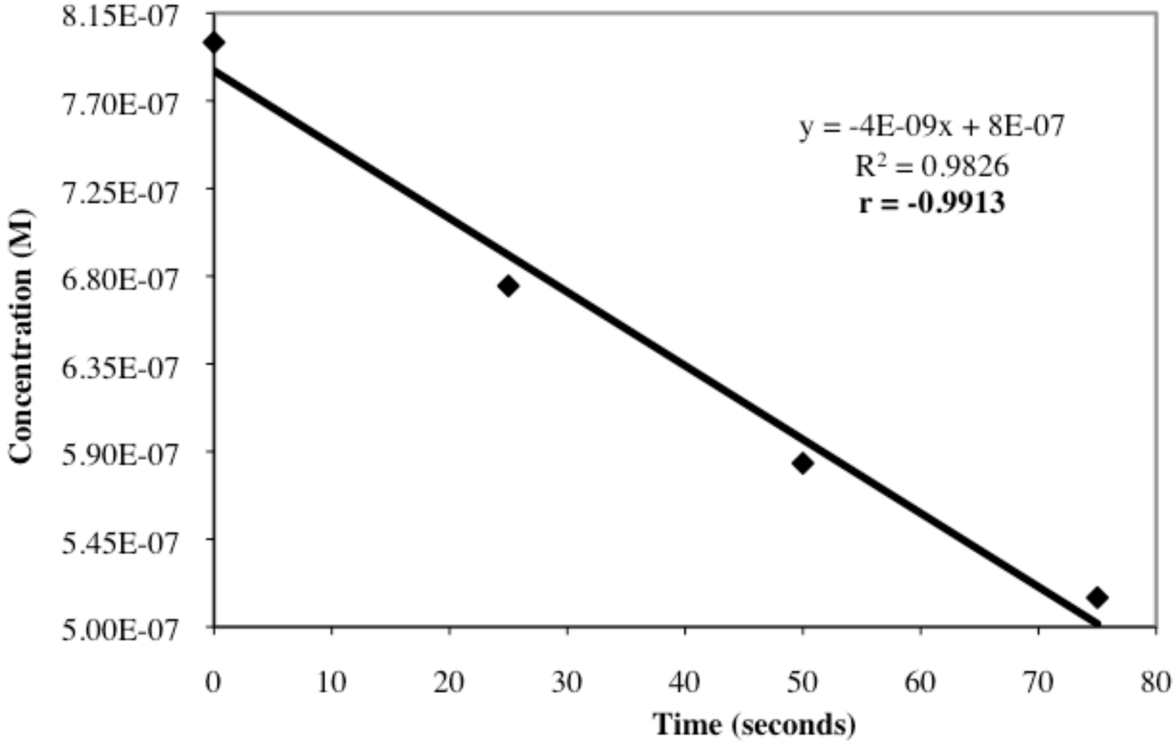
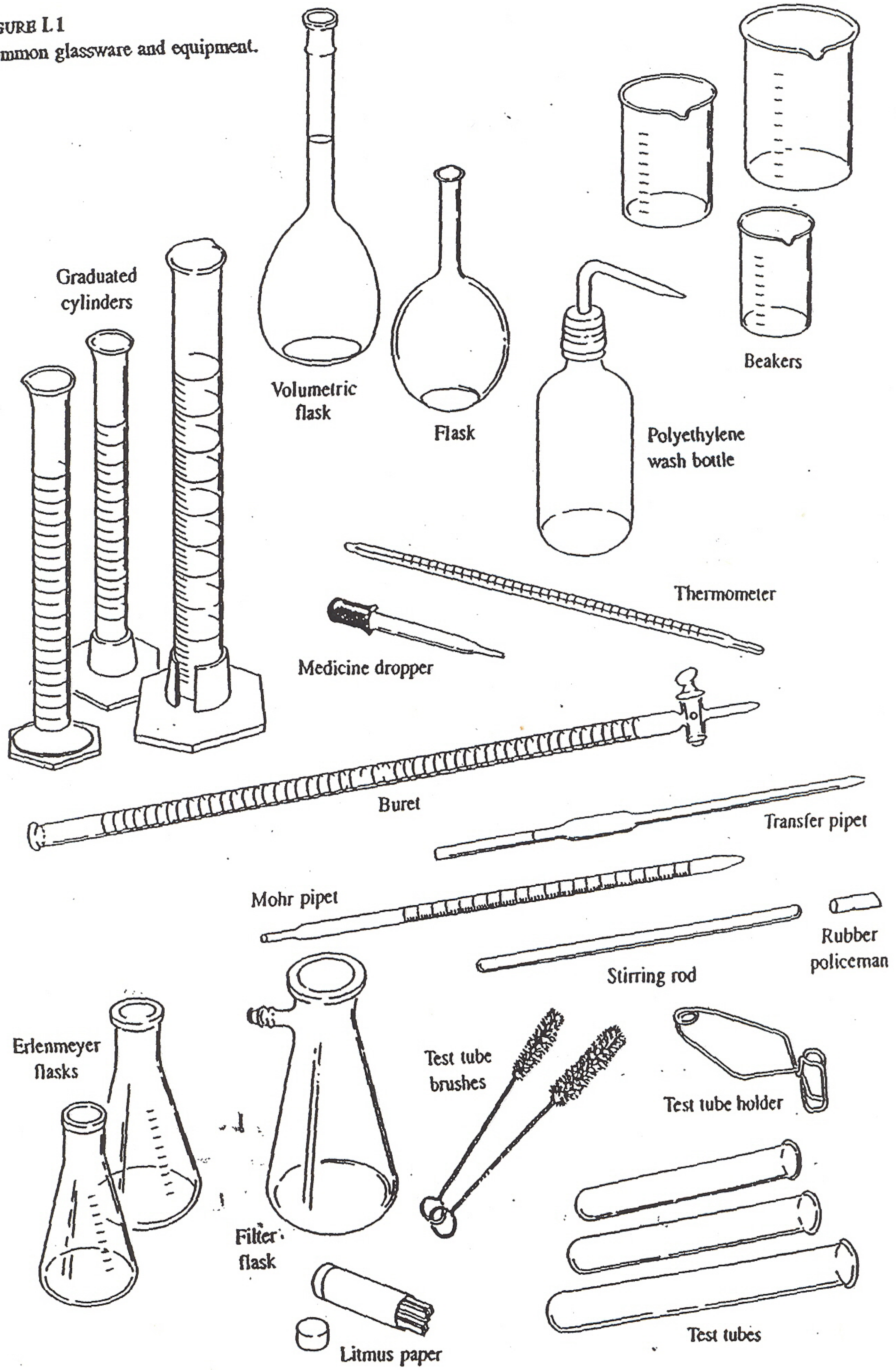
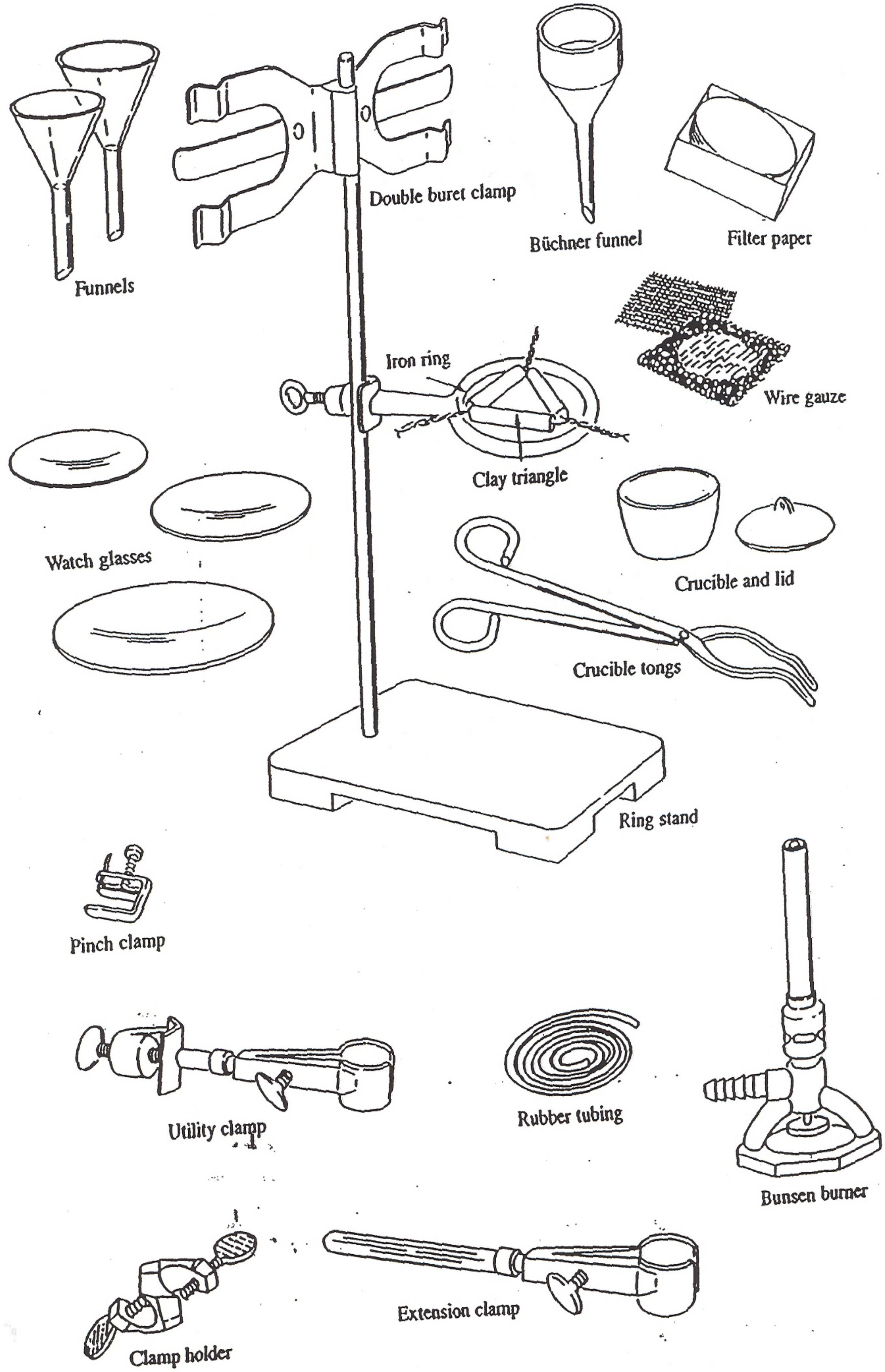

8.15E-07
7.70E-07
7.25E-07
6.80E-07
6.35E-07
5.90E-07
5.45E-07
5.00E-07
Concentration (M)
y = -4E-09x + 8E-07
R2 = 0.9826
r = -0.9913
0
10
20
30
40
50
60
70
80
Time (seconds)

FIGURE I.1
Common glassware and equipment.




Funnels
Double buret clamp
Büchner funnel
Filter paper
Iron ring
Wire gauze
Clay triangle
Watch glasses
Crucible and lid
Crucible tongs
Ring stand
Pinch clamp
Utility clamp
Rubber tubing
Bunsen burner
Extension clamp
Clamp holder

Periodic table from ScienceNotes.org

Periodic Table of the Elements

1 IA 1A	2 IIA 2A	3 IIIB 3B	4 IVB 4B	5 VB 5B	6 VIB 6B	7 VIIB 7B	8	9 VIII 8	10	11 IB 1B	12 IIB 2B	13 IIIA 3A	14 IVA 4A	15 VA 5A	16 VIA 6A	17 VIIA 7A	18 VIIIA 8A
1 H Hydrogen 1.008																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.066	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.631	33 As Arsenic 74.922	34 Se Selenium 78.971	35 Br Bromine 79.904	36 Kr Krypton 83.798
37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium 98.907	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.414	49 In Indium 114.818	50 Sn Tin 118.711	51 Sb Antimony 121.760	52 Te Tellurium 127.6	53 I Iodine 126.904	54 Xe Xenon 131.294
55 Cs Cesium 132.905	56 Ba Barium 137.328	57-71	72 Hf Hafnium 178.49	73 Ta Tantalum 180.948	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.217	78 Pt Platinum 195.085	79 Au Gold 196.967	80 Hg Mercury 200.592	81 Tl Thallium 204.383	82 Pb Lead 207.2	83 Bi Bismuth 208.980	84 Po Polonium [208.982]	85 At Astatine 209.987	86 Rn Radon 222.018
87 Fr Francium 223.020	88 Ra Radium 226.025	89-103	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [269]	109 Mt Meitnerium [278]	110 Ds Darmstadtium [281]	111 Rg Roentgenium [280]	112 Cn Copernicium [285]	113 Nh Nihonium [286]	114 Fl Flerovium [289]	115 Mc Moscovium [289]	116 Lv Livermorium [293]	117 Ts Tennessine [294]	118 Og Oganesson [294]

Lanthanide Series	57 La Lanthanum 138.905	58 Ce Cerium 140.116	59 Pr Praseodymium 140.908	60 Nd Neodymium 144.243	61 Pm Promethium 144.913	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.925	66 Dy Dysprosium 162.500	67 Ho Holmium 164.930	68 Er Erbium 167.259	69 Tm Thulium 168.934	70 Yb Ytterbium 173.055	71 Lu Lutetium 174.967
Actinide Series	89 Ac Actinium 227.028	90 Th Thorium 232.038	91 Pa Protactinium 231.036	92 U Uranium 238.029	93 Np Neptunium 237.048	94 Pu Plutonium 244.064	95 Am Americium 243.061	96 Cm Curium 247.070	97 Bk Berkelium 247.070	98 Cf Californium 251.080	99 Es Einsteinium [254]	100 Fm Fermium 257.095	101 Md Mendelevium 258.1	102 No Nobelium 259.101	103 Lr Lawrencium [262]

Parts per Thousand (ppt) Guide

Parts per thousand (ppt), also known as the "relative standard deviation", is useful when comparing the uncertainty between different measurements of varying magnitude (i.e. it is a measure of the ***precision*** within an experiment.) **Parts per thousand can be applied to any set of data** where more than one experimental value has been applied – i.e. volumes, percentages, concentrations, etc. We will use parts per thousand often this year, so knowledge of how it works is critical for the successful student.

For the values x_1, x_2 and x_3:

- Take the **average** of the values

$$\text{average} = \frac{\text{sum}}{\#\ \text{of values}} = \frac{x_1 + x_2 + x_3}{3}$$

- Find the **deviation** of each value relative to the average

$$\text{deviation}_1 = \text{absolute value (average} - x_1) = |\,\text{average} - x_1\,|$$
$$\text{deviation}_2 = |\,\text{average} - x_2\,|$$
$$\text{deviation}_3 = |\,\text{average} - x_3\,|$$

- Find the **average deviation** of the deviations

$$\text{average deviation} = \frac{\text{sum of deviations}}{\#\ \text{of values}} = \frac{deviation_1 + deviation_2 + deviation_3}{3}$$

- Calculate the **parts per thousand (ppt)** for the values

$$\text{ppt} = \frac{\text{average deviation}}{\text{average}} * 1000$$

<u>*Example*</u>*:* Calculate the parts per thousand for the values 35.72%, 35.92% and 36.02%

- Average = $\frac{35.72 + 35.92 + 36.02}{3}$ = **35.89 %**
- $\text{Deviation}_1 = |\,35.89 - 35.72\,| = 0.17$
- $\text{Deviation}_2 = |\,35.89 - 35.92\,| = 0.03$
- $\text{Deviation}_3 = |\,35.89 - 36.02\,| = 0.13$
- average deviation = $\frac{0.17 + 0.03 + 0.13}{3}$ = **0.11 %**
- parts per thousand = $\frac{0.11}{35.89} * 1000$ = **3.1** *unitless*

Parts per thousand relates the deviation to the magnitude of the experimental data. Consider these two sets of data each with an average deviation of ±0.010:

Data set 1: 0.250 ± 0.010, ppt = (0.010/0.250) x 1000 = 40 ppt (not very good precision).

Data set 2: 4.50 ± 0.010, ppt = (0.010/4.50) x 1000 = 2 ppt (excellent precision)

Although both scenarios have the same deviation, the relative deviation compared to the data gives very different results. Patience and focus is a virtue in this lab.

Fall 2024 Chemistry 221 *with* Dr. Michael A. Russell

For Section 01, Section H1 and Section W1

CH 221, Mt. Hood Community College, Gresham, Oregon, USA 97030

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Email: **mike.russell@mhcc.edu** *or* **mike.russell@mhchem.org**

Office Hours: Held in AC 2568 **MW 10 AM - 11, MW noon - 1 PM** *and* **F 8 AM - 9**
CH 221 Discord server: **https://discord.gg/XrumtbY**

Chemistry 221 website:
http://mhchem.org/221

Required/Recommended Materials*:*
* "Chemistry" *by* The OpenStax College (978-1-947172-62-3)*,*
available here for free: **http://mhchem.org/text/OpenStaxChem.pdf**
* **Chemistry 221 Companion***, purchase here:* **http://mhcc.edu/bookstore (required)**
* **Scientific calculator** *with at least EXP/EE and ln and log (ideally)* **(required)**
* *iClicker Student App subscription if attending lectures (optional)*

Course Description: This course offers the fundamental basis of chemistry for science, preprofessional, chemistry and engineering majors. A strong emphasis is placed on a mathematical approach. CH 221 covers nomenclature, stoichiometry, thermochemistry and introductory chemical bonding. ***Prerequisites****:* RD090, WR090, each with a grade of "C" or better; or placement above stated course levels. Co-requisite of MTH111 or placement above stated course level. CH151 highly recommended.

Course Philosophy: To be successful, students enrolled in a 200 level chemistry course should complete all assignments before coming to class, attend classes regularly, participate in discussions, and think critically. Homework assignments represent the *minimum* requirement for understanding the principles of chemistry. It is assumed that A and B students will perform enough *unassigned* exercises to master key concepts. I encourage questions in this class, and I expect a considerable amount of work. If you contact me by email, I will respond to you normally within 24 hours; phone messages will be acted upon as soon as possible.

The Honor Principle: All students will be expected to behave with the highest moral and academic integrity while enrolled in this class. Plagiarism, cheating or sharing information on tests or laboratory reports, disruptive behavior, and other related offenses will be dealt with according to the directives stated in the current *Mt. Hood Community College Student Guide*. Offering, asking for, giving or receiving help from a person or website without instructor consent is cheating. Copying and/or sharing any course materials outside this class is not allowed and illegal due to copyright laws.

Grading:	Midterm Exams (2 total, 130 points each)	260 points	*26% of total points*
	Quizzes (6 total, lowest quiz dropped, 20 points each)	100 points	*10%*
	Lecture Final Exam	180 points	*18%*
	Laboratory Final Exam	100 points	*10%*
	Class Presentation	100 points	*10%*
	Problem sets, worksheets, reserve CP topic	50 points	*5%*
	Nine lab experiments (20 points each)	180 points	*18%*
	Lab Completion Bonus	30 points	*3%*
	Total points:	**1000** points	***100%***

Tentative grading distribution: **A**: 89-100% **B**: 78-88% **C**: 65-77% **D**: 55-64% **F**: less than 55%

Opportunities for extra credit are available and explained in the "Extra-Credit Guide" handout.

Exams and Quizzes will be completed exclusively in class (sections 01 and H1) or exclusively online (section W1.) Sections 01 and H1 must turn in assignments in person to avoid a point penalty. Section W1 assignments must be submitted via email to the instructor in a suitable format, and Section W1 must show work on all problems to get full credit.

Labs and Problem Sets will be submitted on campus (sections 01 and H1) or via email (section W1.) Sections 01 and H1 must be present during recitation for full credit, and a stamp system will be used to guarantee on-time attendance if necessary.

Each student will submit a **Class Presentation** this term - see the "Class Presentation FAQ" for more information.

Details regarding grading will be discussed during the first week of the term.

"What's Due This Week" Schedule for CH 221 Fall 2024

***All assignments** can be found on our **website** (http://mhchem.org/221)*

***Assignments are different for section W1 and sections 01 and H1** - contact the instructor if you are unsure which applies to you*

- **Sections 01 and H1** *must bring a printed copy of the lab on the specified day, then turn it in on the deadline during class. Problem sets and labs will be turned in during class in AC 2501; quizzes and exams will be completed during class time. Emailed assignments from Section 01 and H1 students will incur a point penalty, no exceptions.*
- **Section W1** *will **email** all assignments to the instructor as a single PDF file.*

Week	Dates	Assignment
1	9/23 - 9/27	Introduction *to the course*
		01/H1 *Lab:* "Eight Bottles" (Lab #1) *due next week in recitation*
		W1 *Lab:* "Introduce Yourself" (Lab #1) *due Friday, September 27 by 9 AM via email*
2	9/30 - 10/4	*Due:* Problem set #1 *Chapter 1;* **01, H1:** *due in recitation;* **W1:** *due 10/2 by 11:59 PM via email*
		Due: Quiz #1; **01, H1:** *take in recitation;* **W1:** *due 10/4 by 9 AM via email*
		01/H1 *Lab:* "Density (in class)" (Lab #2) *due next week in recitation*
		W1 *Lab:* "Density (online)" (Lab #2) *due Wednesday, October 9 by 11:59 PM*
3	10/7 - 10/11	*Due:* Problem set #2 *Chapter 2 & 3.1;* **01, H1:** *due in recitation;* **W1:** *due 10/9 by 11:59 PM*
		Due: Quiz #2; **01, H1:** *take in recitation;* **W1:** *due 10/11 by 9 AM*
		01/H1 *Lab:* "Chemical Nomenclature" (Lab #3) *due next week in recitation*
		W1 *Lab:* "Chemical Nomenclature" (Lab #3) *due Wednesday, October 16 by 11:59 PM*
		October 11, 9 AM: *Last chance to reserve a Class Presentation element*
4	10/14 - 10/18	*Due:* Problem set #3 *Chapter 2 & 3.1-3.2;* **01, H1:** *due in recitation;* **W1:** *due 10/16 by 11:59 PM*
		Due: Quiz #3**; 01, H1:** *take in recitation;* **W1:** *due 10/18 by 9 AM*
		01/H1 *Lab:* "Empirical Formula (in class)" (Lab #4) *due next week in recitation*
		W1 *Lab:* "Empirical Formula (online)" (Lab #4) *due Wednesday, October 23 by 11:59 PM*
5	10/21 - 10/25	**EXAM #1** *(Chapters 1-3);* **01, H1***: take in recitation;* **W1***: due 10/25 by 9 AM*
		Due: "Exam Prep I"*;* **01, H1*:*** *due in recitation;* **W1:** *due 10/23 by 11:59 PM*
		01/H1 *Lab:* "Percent Potassium Chlorate (in class)" (Lab #5) *due next week in recitation*
		W1 *Lab:* "Percent Potassium Chlorate (online)" (Lab #5) *due Wednesday, October 30 by 11:59 PM*
6	10/28 - 11/1	*Due:* Problem set #4 *Chapter 4**;*** **01, H1:** *due in recitation;* **W1:** *due 10/30 by 11:59 PM*
		Due: **Class Presentation Rough Draft Paper***;* **01, H1**: *due in recitation;* **W1:** *due 10/30 by 11:59 PM*
		Due: Quiz #4**; 01, H1:** *take in recitation;* **W1:** *due 11/1 by 9 AM*
		01/H1 *Lab:* "Net Ionic Reactions (in class)" (Lab #6) *due next week in recitation*
		W1 *Lab:* "Net Ionic Reactions (online)" (Lab #6) *due by Wednesday, November 6 by 11:59 PM*

7	11/4 - 11/8	*Due:* Problem set #5 *Chapter 3.3, 3.4 & 5*; **01, H1:** *due in recitation;* **W1:** *due 11/6 by 11:59 PM* *Due:* Quiz #5; **01, H1:** *take in recitation;* **W1:** *due 11/8 by 9 AM* **01/H1** *Lab:* "Unknown Chloride (in class)" (Lab #7) **H1:** *due 11/13;* **01:** *due 11/18 (Veterans Day)* **W1** *Lab:* "Unknown Chloride (online)" (Lab #7) *due 11/13 by 11:59 PM*
8	11/11 - 11/15	**Monday, November 11: Veterans Day**, *all classes, office hours canceled,* **01** *schedule delayed 1 week* **EXAM #2** *(Chapters 3-5);* **01:** *take in recitation on 11/18;* **H1***: take in recitation on 11/13;* **W1***: due 11/15 by 9 AM* *Due:* "Exam Prep II"*;* **01:** *due in recitation on 11/18;* **H1***: due in recitation on 11/13;* **W1***: due 11/13 by 11:59 PM* **01/H1** *Lab: "*Calorimetry (in class)" (Lab #8); **01:** *on 11/18, due 11/25;* **H1:** *on 11/13, due 11/20* **W1** *Lab: "*Calorimetry (online)" (Lab #8) *due Wednesday, November 20 by 11:59 PM* **November 15:** *Last day to drop or change grade status this quarter at Mt. Hood Community College*
9	11/18 - 11/22	**CLASS PRESENTATIONS WEEK** **01/H1:** *Class Presentation paper due at time of presentation during recitation.* **01:** *on 11/25;* **H1:** *on 11/20* **W1:** *Class Presentation paper and video due Wednesday, November 20 by 11:59 PM*
10	11/25 - 11/29	*Happy Thanksgiving!* **01** *meets on Monday, nothing due to* **H1** *or* **W1**
11	12/2- 12/6	*Due:* Problem set #6 *Chapter 6*; **01, H1:** *due in recitation;* **W1:** *due 12/4 by 11:59 PM* *Due:* Quiz #6; **01, H1:** *take in recitation;* **W1:** *due 12/6 by 9 AM* **01/H1** *Lab:* "Hydrogen Spectrum (in class)" (Lab #9) *due at the end of the lab period (same day)* **W1** *Lab:* "Hydrogen Spectrum (online)" (Lab #9) *due Wednesday, December 4 by 11:59 PM* *All extra credit closes Friday, December 6 at 9 AM*
12	12/9 - 12/11	**Take Home Lab Final** *released by 9 AM Monday, 12/9 for all CH 221 sections. Sections 01 and H1* ***must*** *print the Take Home Lab Final and turn it in on Wednesday.* *Section 01: Take* **Lecture Final** *tentatively* on Wednesday, December 11 at 8:45 AM in AC 1303. *Due:* **Final Exam Prep** worksheet, **Take Home Lab Final** *Section H1: Take* **Lecture Final** *tentatively* on Wednesday, December 11 at 1:10 PM in AC 2501. *Due:* **Final Exam Prep** worksheet, **Take Home Lab Final** *Section W1: Due:* **Lecture Final** *(available Monday December 9)*, **Final Exam Prep** worksheet and **Take Home Lab Final** *on* Wednesday, December 11 by 11:59 PM

Getting Started in Chemistry 221

Welcome to Chemistry 221! I am glad to have you enrolled in CH 221! Here are some hints on how to get started in the class:

- First, **know that I am here to help you succeed in this class**. If you have any questions, please email me (mike.russell@mhcc.edu) or stop by the Discord server (https://discord.gg/XrumtbY). I try to respond to student inquiries within 24 hours.

- There are **three sections of CH 221** this quarter, namely **section 01** (which meets three times a week on campus), **section H1** (which meets only once a week) and **section W1** (which is completely online). Sections 01 and H1 will have similar schedules, but section W1 will exhibit some differences. Your experience in this class will depend on which section you are in, so email the instructor (mike.russell@mhcc.edu) if you have any questions about anything, ok?

- **Purchase the Chemistry 221 Companion from the MHCC Bookstore**. The MHCC Bookstore (http://mhcc.edu/bookstore) will ship you a printed copy of this necessary information; alternatively, you can print the document (on our website), but I do not recommend it due to the size of the Companion. You will need access to printed materials this quarter!

- The **"What's Due This Week" Schedule for CH 221** located on page 2 of your syllabus **lists all the problem set due dates, assignment deadlines, labs performed, exam/quiz dates, and related information for this term**. You can plan your term by referencing this handout.... follow it closely and you will do well in CH 221!

- **Check your email often during Chemistry 221**. I will be sending weekly reminders as to "what is due this week" in CH 221 closely as well as returning some assignments, etc. If you would prefer that I use a non-saints email address to communicate with you, let me know - this is easy to set up!

- The **Chemistry 221 website** is worth exploring. The Chemistry 221 website has a host of learning opportunities waiting for you. You can download and/or print copies of the syllabus, lecture notes, labs, quiz answers, and more; plus there are opportunities for extra credit available. To get started, send your web browser to

  http://mhchem.org/221

  You should see the CH 221 website on your screen.

- Check out the **Chemistry 221 Chapter Guides** by selecting "**Chapter Guides**" from the upper left hand corner of the CH 221 website. The Chapter Guides offer a detailed approach for studying the course material through a series of lessons. **Read Lesson Zero**, the "Intro to the Chapter Guides System," to understand how they work.

- Start thinking about a **Class Presentation Topic**. Depending on your section, you will be creating a five minute face to face or video presentation this term on an **element**, and you must reserve your element choice with me. A written paper also accompanies the presentation on your element. To reserve your element, email the instructor or send your web browser here:

  http://mhchem.org/cp221

  The "**Class Presentations FAQ**" (available in the syllabus or here: **http://mhchem.org/cp221info**) has more information.

- The **Chemistry 221 Textbook** is **free** and **legal to download** from our website: **http://mhchem.org/text/OpenStaxChem.pdf**

- Section W1 students: I highly recommend you check out the **CamScanner** app (http://camscanner.com) in order to send your work to me as a PDF file over email. CamScanner is free and easy to use.... but there are other options besides CamScanner, use the method best for you. Section 01 and H1 students must submit their work on paper for full credit.

- Many opportunities for **extra credit** exist in this class.... see the Extra Credit Guide for more information: **http://mhchem.org/xc**

- You can **download** the entire **Microsoft Office** suite of programs (Word, Excel, PowerPoint, etc.) **for free**... see this link for information: **https://mhcc.edu/OfficeInstall/**

Again, welcome to Chemistry 221! Let me know if I can make your learning experience better in any way, and I look forward to working with you this term! Peace, Dr. Michael Russell (mike.russell@mhcc.edu, 503.491.7348, AC 2568)

CH 221 CLASS PRESENTATIONS FAQ

FAQ = Frequently Asked Questions

When: **Monday, November 25 (section 01) or Wednesday November 20 (sections H1 & W1)**

What: A chance to share knowledge of the elements with your classmates and the MHCC community

Who: ***Everyone*** **enrolled in CH 221 (All Sections)**

What topic should I pick?

For CH 221, the topic will be **elements**. Pick an element you find interesting and write a report on the topic. Since there are over 100 elements, every student must pick a different element. **Reserve** your element with Dr. Russell using the online form at **http://mhchem.org/cp221**

Once your element has been chosen, begin researching interesting information on the element using the library, internet, etc. You will be preparing a paper on the element and presenting your work to the class in a short (five minute) presentation.

If you need to change your class presentation topic after the fourth week of class for any reason you will be penalized 20 points; hence, it's best to reserve an element early and start researching promptly. Also, if you still have not reserved an element by the end of the sixth week, you will be penalized 20 points for tardiness.

What should I know when writing the paper?

Prepare a paper that is at least **five full pages** of text on your reserved element. Diagrams, pictures, and other graphics are wonderful, but you will need five full pages of writing for complete credit.

The paper should include a separate **cover sheet** with the title of your presentation and your name. The paper must be neat, typed, referenced, and interesting to read; spelling and grammar will count. The paper must use a **"reasonable" font and font size** (Times New Roman, Arial, etc. with size 12 or less); in addition, use **1" margins or less** *(I will measure!)* and **no more than "one and a half" spaced type** (less than double spaced.) If unsure, ask the instructor.

A ***separate*** **page with at least eight references** will be at the end of your paper. References within the paper and at the end should adhere to the "**Class Presentations Citation Guide**" (*found here*: **http://mhchem.org/cg**) For an *example* paper, see: http://mhchem.org/expaper

What is a peer reviewed scientific article?

An important aspect of this assignment is to ensure scientific relevancy. To this end, **find two peer reviewed scientific articles published within the last ten years** that include a reference to your element. **Include the abstracts of these papers with your final Class Presentation paper.**

How do I find my two peer reviewed scientific articles?

A sure-fire way to access **peer-reviewed scientific articles** is through the MHCC library's article databases. Go here (**https://libguides.mhcc.edu/chemistryguide** - you may have to enter your MyMHCC username and password if you are off campus). Select **Articles (**on the left), then select **ScienceDirect College Edition** (under "Chemistry Databases") or **Academic Search Complete** (under "General Databases"), then search for your topic. ***Remember***, your article citation should include the author(s), year of publication, journal title, title of paper, page number(s), volume of journal, etc. and you will need to include the abstract from the peer reviewed article (but not the entire article!) in your report.

Once you conduct a search for your presentation topic, you will likely have a mix of citation/abstract-only and citation/abstract + full-text (whole article) results. You ***only*** need the abstract for your paper - do not include the full article. Here is an **example** of a **peer-reviewed scientific paper with an abstract**: http://mhchem.org/abstr

Tell me about the Class Presentation Rough Draft Paper

In the middle of the term you will be submitting a rough draft of your class presentation paper to the instructor. Ideally this will be the paper in a mostly complete format, but at the very least, two typed pages of text with one peer reviewed abstract and citation should be submitted.

The rough draft should include at least one peer reviewed scientific paper abstract (with its citation) as well as the **Rough Draft Class Presentation form** (http://mhchem.org/rd1) The Class Presentation Rough Draft Paper is worth 20 points (out of the 100 points total.)

What should I know when preparing for the presentation?

You will be creating a five-minute **presentation** on your chosen subject. Sections 01 and H1 will give their presentation during a lab period to their peers; Section W1 will record themselves and upload the video to YouTube for the instructor to view. The presentation must be well prepared and interesting; sloppy preparation shows in the presentation portion. Students can use videos, presentation software (PowerPoint, etc.), posters and chalk to enhance their presentation. Presentation software users will be limited to a maximum of six slides; more invokes a penalty.

Section W1: I encourage students to record themselves on their phone, then upload the video to YouTube, etc. You can send the link (to an **unlisted** video, not private) to me for watching later. I need to see **you** for 90% or more of the presentation for full credit.

Please note that using your paper (or a *copy* of your paper) during the presentation will result in a ten point penalty. This will prevent you from "reading" your presentation to the audience.

How will I be graded?

There are 100 points total for this project. 40 points will reflect the work presented in the paper, 40 points will reflect the work done in the presentation and 20 points will be given for completing the Rough Draft Class Presentation paper.

In addition, failure to turn in the "Class Presentation Reviewer Guide" to the instructor at the end of the day of presentations will result in a ten-point penalty. You will be completing the Reviewer Guide while others are giving their presentations. If curious, you can view this guide on our website (http://mhchem.org/cp221info). *This applies to Sections 01 and H1 only.*

Late class presentations will result in a five-point penalty *per day*. The paper and the presentation must be completed for credit on this assignment. Plagiarism discovered from any source will result in a **total** Class Presentation grade of zero.

A sample **Class Presentation Grading Rubric** is available for viewing on the CH 221 website (http://mhchem.org/cgr1) The rubric will allow you to look at the items deemed most important when grading your Class Presentation.

How do I get started?

Step 1: Reserve your Class Presentation Element

Decide on some elements that interest you, then email the instructor or complete the online web form to reserve your element: **http://mhchem.org/cp221**

You should receive a response from Dr. Russell within 48 hours after the beginning of the second week of class; if you do not, email him directly at mike.russell@mhcc.edu. Be sure to include alternate elements in case your first choice has already been claimed; he can also pick one for you if you are uncertain which element to pick. Reserve your class presentation element by the end of the third week, **October 11** at **9 AM**. You can see which elements are still available here: **http://mhchem.org/221av**

Step 2: Turn in the Class Presentation Rough Draft Paper

The Class Presentation Rough Draft paper should include at least two typed pages and one peer reviewed scientific article and include the handout (http://mhchem.org/rd1) at the beginning of your paper. Deadline: **Mon, Oct. 28 (01)** or **Wed., Oct. 30 (H1, W1)**

Step 3: Give the Class Presentation and turn in your final Class Presentation paper

Section 01 and H1: Bring your final Class Presentation paper and give your presentation during lab. *Section 01:* **Monday, November 25.** *Section H1:* **Wednesday, November 20.**

Section W1: At the time of your Class Presentation, turn in your final Class Presentation paper and a video recording of your presentation (YouTube link, etc.) **by Wednesday, November 20 by 11:59 PM.**

Before you present and submit your paper, check out the reminders contained within the "**Class Presentation 'Last Minute' Checklist**", found here: http://mhchem.org/cpcs Note that you must both present your work and submit your paper to receive any points on this assignment. All presentations over ten minutes in length and all papers with more than 10 pages of writing will suffer a point penalty.

If you have any questions, see this site (http://mhchem.org/cp221info) or contact the instructor.

Name: ____

CH 221 CLASS PRESENTATIONS ROUGH DRAFT PAPER

*Staple this form to the **top** of your Rough Draft Class Presentation Paper for full credit*

Lab Section:

Reserved Element:

Directions:

- This assignment is worth 20 points out of the 100 points assigned to the Class Presentation assignment.
- *Include* at least two typed pages of your Class Presentation report with this form (more is fine!)
- *Include* at least one abstract from a peer reviewed scientific article with a proper citation included (more is fine!)
- This page should be stapled (Sec. 01 and H1) or attached (Sec. W1) to the top of the other pages in this assignment to avoid a five-point penalty

Helpful Resources:

- The CH 221 Class Presentation Frequently Asked Questions handout: **http://mhchem.org/faq1**
- The CH 221 Citation Guide: **http://mhchem.org/cg**

Class Presentation Rough Draft Paper Due Dates:

- **Section 01:** **October 28 at 1:10 PM**
- **Section H1:** **October 30 at 1:10 PM**
- **Section W1:** **October 30 at 11:59 PM**

Section 01: *The final Class Presentation paper is due at the time of your presentation on Monday, November 25.*
Section H1: *The final Class Presentation paper is due at the time of your presentation on Wednesday, November 20.*
Section W1: *The final Class Presentation paper and recorded presentation will be due by 11:59 PM on Wednesday, November 20.*

Staying Connected in Chemistry 221 This Quarter

Success in Chemistry 221 often depends on staying connected with the flow of the course... here are some suggestions on how to be aware of what is happening each week:

- **Discord** is a wonderful medium for keeping students connected while in this class. Our Discord server will offer weekly assignment updates (with links to labs, problem sets, etc.) as well as links to video lectures, tips and hints from the instructor on how to conquer difficult problems, and more! Joining Discord is easy and free; go here (**https://discord.gg/XrumtbY**) and join the CH 221 server to get started.

- I'd be honored if you would subscribe to me on **YouTube**! (**http://youtube.com/marsmars2**) I create videos for more than just chemistry classes... :)

Additional Syllabus Materials for Chemistry 221 Available on the Internet

Some or all of these materials might prove useful to you in our class. All of them are available on the Chemistry 221 website (https://mhchem.org/221/classroom/ci.htm).

To access these materials (and more!), go to our website (**http://mhchem.org/221)** and select "**Getting Started**" then "**Other Class Information**" from the upper left corner. Additional materials include:

- The **Extra-Credit Guide** - a helpful guide containing some of the extra credit options available to you in this course

- **Learning Outcomes for CH 221** - a list of "what you will learn" this quarter

- **MHCC College Information** - key information that you, as a student at Mt. Hood Community College, might wish to know, including the **Student Code of Conduct** and **Internet Privacy Policy**

- A **Printable Periodic Table** – this periodic table from ScienceNotes.org will certainly be useful in this course, and you will be able to use this type of periodic table on exams and quizzes.

- The **Chemistry Smiles Generator –** in case you need a smile :) with a chemistry theme.

In addition, the website has a plethora of other "goodies" which may be of assistance to you throughout this quarter... feel free to browse, and if you have questions, please do not hesitate to contact me.

Have a great quarter! Peace,
Michael A. Russell, Ph.D. (he/him/his)
mike.russell@mhcc.edu
(503) 491-7348, AC 2568 (office on campus)
mhchem.org/221

CH 221 Fall 2024:
“Eight Bottles” ***Lab***
Instructions

Note: **This is the lab for section 01 and H1 of CH 221 only.**

- *If you are taking section W1 of CH 221*, please use this link:
  http://mhchem.org/s/1b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-1-2 through Ia-1-4 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, September 23 (section 01) *or* **Wednesday, September 25 (section H1).** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work and calculations on your own, then **turn it in** (pages Ia-1-3 through Ia-1-4 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, September 30 (section 01)** *or* **Wednesday, October 2 (section H1.)** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Eight Bottles

An Introduction to Scientific Investigations

INTRODUCTION

Problem solving is not restricted to scientific investigations. Indeed, it is a life long process that involves every aspect of human endeavor. The way one solves a problem is related to one's individual learning style. There are, however, some common factors which seem to be part of most scientific investigations. Although, every investigator, being human, approaches each problem with some preconceived ideas, facts are gathered by accurate observation of behavior of the system of interest. Conclusions are based solely on the observed data.

Only by using experimental observations to study the behavior of matter, arranging the results of such studies in an orderly fashions, correlating the observed data and testing these correlations (theories or hypothesis) by further systematic observations can one hope to increase our ability to deal with the physical world around us. This approach is usually referred to as the Scientific Method. There is nothing unique about the order of activity to this method other than it provides a logical way to deduce order and causality for natural phenomena. An inherent part of the scientific method is the element of creativity. This is what makes possible the development of completely new concepts. This experiment is designed to allow you to use some of the elements of scientific investigation mentioned here.

Eight bottles, labeled A through H, containing eight different solutions have been prepared for your examination. When mixed together, in pairs, several of these solutions will undergo a chemical reaction. A reaction can be observed by one of the following changes:

1. A color change will occur.
2. A gas will be evolved (bubbles will be observed)
3. A precipitate (a cloudy mixture) will form.
4. The evolution of heat.

Be carefully observing any changes that occur it should be possible for you obtain enough data to characterize each of these solutions. <u>In this experiment</u> we will use only a <u>color change</u> or a <u>precipitate</u> to detect a chemical reaction. From the results of your study, you should then be able to prepare a concise description of how to identify the contents of an <u>unlabeled</u> bottle assuming the others are available for mixing.

PROCEDURE

Obtain a tray with dropper bottles containing solutions labeled A-H. In each dimple of a spot plate, combine 3 drops each of various combinations of two solutions. Mix with a clean stirring rod, and record your observations (**color change or precipitate formation along with the color of the precipitate**) in the data table provided. When the spot plate is full, simply rinse it off with distilled water into the waste container and continue experimenting.

Obtain two unknown solutions from your instructor and record their ID #s on your data sheet. Experiment with these unknown by mixing with the contents of each of the bottles labeled A-H. Remember to mix only two solutions at a time. Record your observations on the data sheet.

From the data recorded in your data table, determine the identity of your unknowns (one of the solutions A-H)

Eight Bottles

<u>NAME</u>:
Lab Partner(s):
*Include **all** first **and** last names for full credit!*

Solutions	A	B	C	D	E	F	G	H
A								
B								
C								
D								
E								
F								
G								
H								
Unknown Number:								
Unknown Number:								

Conclusion:

Unknown number __________ Identity (letter) __________

Unknown number __________ Identity (letter) __________

Based upon your observations, describe briefly how you identified the unknown numbered solutions (above) containing one of the eight known solutions (A-H).

Postlab Questions:

1. Why do you obtain the maximum useful information about the solutions by mixing only two solutions at a time?

2. How would you detect the evolution of a gas upon mixing the solutions? (and remember, not all gases have a smell.)

3. Which of the A-H "solution(s)" could be distilled water? How do you know?

Please note: The instructor will send you email throughout the term, so *please check your email several times each week!* The instructor will use your @saints.mhcc.edu address by default, but if you wish to use an alternate email address, send an email to mike.russell@mhcc.edu from your alternate email account and it will be changed promptly.

CH 221 Fall 2024:

"Density" *(in class)* ***Lab***

Instructions

Note: **This is the lab for section 01 and H1 of CH 221 only.**

- *If you are taking section W1 of CH 221*, please use this link:
  http://mhchem.org/s/2b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-2-2 through Ia-2-9 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, September 30 (section 01) *or* **Wednesday, October 2 (section H1).** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work and calculations on your own, then **turn it in** (pages Ia-2-5 through Ia-2-9 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, October 7 (section 01)** *or* **Wednesday, October 9 (section H1.)** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Density

Density, like color, odor, melting point, and boiling point, is a physical property of matter. Therefore, density may be used in identifying matter. Every substance (element, compound, alloy, etc.) has a distinct density. **Density** is defined as **mass per unit volume** and is expressed mathematically as **d = m/V** (d is density, m is mass, and V is volume). Density is essentially a measurement of how tightly matter is packed together.

Density is an important concept in a wide range of fields including chemistry, physics, material science, engineering, geology, meteorology, biology and medicine. For example, a bone density test uses X-rays to determine how much calcium and minerals are packed into a segment of bone. You may also be familiar with different types of plastics, including high-density polyethylene (HDPE, #2, used for milk jugs, hula hoops, and breast implants) and low-density polyethylene (LDPE, #4 used mostly for plastic bags.) The difference between these plastics depends on how tightly the polyethylene molecules are packed together during synthesis.

The **density of water** is 1.00000 g/cm^3 at 4 °C and is slightly less at room temperature. In lab today, you will be using the *Handbook of Chemistry and Physics* to determine the exact density of water at a specific temperature. The density of various materials ranges significantly from less than water (styrofoam's density is about 0.1 g/cm^3) to much greater (Osmium has a density of 22.6 g/cm^3.) For example, aluminum has a density of 2.70 g/cm^3 whereas a sample of lead has a density of 11.2 g/cm^3. The same volume of lead will have a mass over four times that of aluminum! That is why lead is used to shield against X-rays whereas aluminum would be ineffective. Aluminum atoms are not only smaller but also packed so that there is more space between atoms.

The SI unit (International System of Units) for density is kg/m^3 and is typically used by physicists and engineers. Because chemists work with much smaller masses and volumes, traditional metric units of **g/cm^3** or **g/mL** are the preferred units of measurement (note that **1 mL = 1 cm^3**). Liquids are usually measured in g/mL while solids are measured in g/cm^3. Gases are much less dense, so their density is measured in **g/L**.

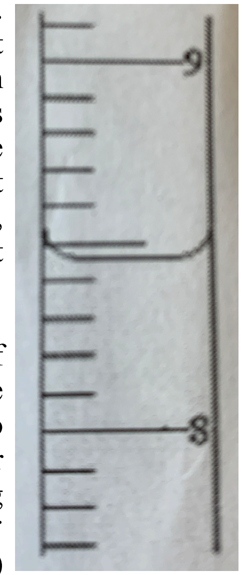
Density is often represented as a relative density or ***specific gravity**,* a dimensionless quantity that expresses density as a multiple of a given standard (such as water or a gas.) For example, gasoline has a density of 0.67 g/cm^3. Its specific gravity, relative to water, is 0.67. Specific gravity is used in many fields from chemical engineers studying concrete to food scientists testing the alcohol content of a microbrew.

To determine density, mass and volume must both be determined. The **mass** can easily be found using a balance. Determining the **volume** is more ambiguous. The volume of a liquid can be determined using a calibrated container such as a graduated cylinder or graduated flask. The volume of a solid sample with a regular geometric shape can be determined by direct measurement. However, most solids have an irregular shape. The volume can be determined by immersing the solid in a known volume of liquid and measuring the volume of liquid displaced.

This is similar to the method utilized in the ancient tale of **Archimedes** to prove that King Hiero II's crown was not real gold. Archimedes is alleged to have come upon the liquid displacement method while bathing and noticing the rise in his bath water. He then ran through the streets shouting "Eureka!" (I found it!), so excited that he forgot his bathrobe. After dressing, he then proved that the king's crown did not displace the same amount of water as a piece of gold of the same mass. This method is called the "**displacement method**" and can be used with a variety of liquids in order to find the density of various materials.

All measurements are approximations. **Significant figures** ("**sig figs**") are those digits that carry meaning which contributes to precision. The uncertainty is in the last digit and determined by the device. For example, when reading a graduated cylinder, the number of sig figs is estimated one digit beyond the gradations. For the *example pictured on the right,* the bottom of the meniscus is between the 8.4 mL and 8.5 mL markings. You can estimate to the hundredth place or 8.45 mL. Reading a meniscus is subjective and takes practice. In this experiment, you will use the mass and density of water to find the volume of a flask more precisely and reduce human bias.

Mass versus Weight: When determining density, you must determine the mass of the sample. The terms *mass* and *weight* are easily confused. The ***mass*** of a substance is how much matter it is composed of. Units of mass are grams and kilograms. The mass of an object is the same on earth or on the moon. ***Weight*** is a measure of the force of gravity acting on the object. Pounds (units = lb) is a unit of weight, a force. The weight of an object is variable depending on the location of the object. If Joe weighs 220 Ibs at the North Pole, he would weigh only 219 Ibs at the equator due to the bulge of the earth. He weighs only 37 Ibs on the moon. In outer space, an astronaut is weightless but never massless. A great blue whale is weightless in space, but it would still cause damage to your spaceship if you bumped into it.



Accuracy and Precision: **Accuracy** is how close a measurement, or average of measurements, come to the actual or accepted value. Accuracy is often compared to hitting the bull's eye on a target. In a chemistry lab, accuracy is how close the final calculated answer is to the accepted book value. When working with an unknown, students are graded on their accuracy, how close their answer is to the actual value. Accuracy is determined by calculating the percent error: **percent error = [(|actual value - experimental value|)/actual value] * 100%** (notice the absolute value in the numerator; percent error should be a positive number.) We will use percent error in an upcoming lab. A "good" percent error varies depending on the experiment, the equipment used, and the technician's experience.

Precision is how close multiple measurements of the same quantity come to each other. Precision is a measure of *consistency in lab technique*; is the data reproducible? One method to determine precision is to calculate **parts per thousand** (see the handout in the "Lab Notes" of your Companion, or ask the instructor.) We will calculate parts per thousand in future labs. The term **precision** also refers to the number of significant digits in a measurement. For example, the balances (scales) that will be used in this lab allow measurements to 1 mg (0.001 g). An analytical balance allows mass.to be determined to 0.1 mg (0.0001 g) and so is more precise. The technique used in this lab for determining volume allows for more significant figures and, hence, is more precise than simply using a graduated cylinder.

Random Error and Systematic Error: **Random errors** originate from uncontrollable variables in an experiment. Momentary fluctuations in air currents can affect balance reading. A student who rushes through the lab and follows directions haphazardly will perform many random errors. Random errors affect the precision of measurements and the overall experiment. **Systematic errors** are controllable or repeated errors in an experiment. A poorly calibrated scale will result in all mass measurements being erroneous by the same factor. A student consistently misreading an instrument is a systematic error. Because a systematic error is consistent throughout the experiment, it does not affect the precision but can significantly affect the accuracy.

In this experiment you will determine densities of an unknown liquid and solid by measuring their mass with a balance and their volume. First, you will determine the exact volume of a flask using water. You will determine the density of a solid by displacement of a known quantity of water.

PROCEDURE:

Part A: Density of a Liquid *All mass measurements should be recorded to the milligram (0.001 g.)*

1. Clean a 10 mL volumetric flask with soap and water. Dry with a small amount of acetone in the hood and by gently blowing compressed air into it. Determine and record the mass (to the nearest mg) of the *clean* and *dry* 10 mL volumetric flask with a stopper in it.
2. Fill this 10 mL volumetric flask with deionized water. Insert stopper so no air remains in flask. Dry the outside of the volumetric flask. Record the mass to the nearest 0.001 grams.
3. Calculate the mass of the water in the flask. Remember to show all calculation steps in your lab report.
4. Determine the temperature of the water to the tenths place. Use the *Handbook of Chemistry and Physics* to find the density of water at this temperature. If the *Handbook* is difficult to read, use this link as an alternative: http://mhchem.org/den
5. Calculate the volume of this volumetric flask. Remember significant digits!
6. Dry the volumetric flask. Obtain an unknown liquid and record the identification number. Fill the volumetric flask with the unknown liquid, stopper and record the mass.
7. Calculate the mass of the unknown liquid added. Calculate the density of the unknown liquid in g/mL to the correct number of significant digits.

Part B: Density of a Solid *All mass measurements should be recorded to the milligram (0.001 g.)*

1. Select an unknown metal and record its identifier.
2. Clean and dry a 50 mL Erlenmeyer flask that will fit your metal sample. Record the mass of the dry flask and stopper. Fill the flask with water. Record the mass.
3. Determine the volume of the Erlenmeyer flask as in part A.
4. Empty and dry the flask thoroughly. Add small chunks of a dry metal sample to the flask until the flask is at least half full. Weigh the flask, with its stopper and the metal, to the nearest milligram. You should have about 50 g of metal in the flask (more is better!)
5. Determine the mass of metal added.
6. Leaving the metal in the flask, fill the flask with water and replace the stopper. Roll the metal around in the flask to make sure that no air is trapped between the metal pieces. Refill the flask if necessary, and then weigh the stoppered flask full of water plus the metal sample.
7. Calculate the mass of water added.
8. Calculate the volume of water added based on its density and mass.
9. Calculate the volume of metal added. Use this value to calculate the density (in g/cm^3) of the metal.
10. Pour the water from the flask. Dry the metal before returning to its container.

Density Lab

YOUR NAME: ______________________
first and last name

LAB PARTNER(s): __________________
first and last names

Data: *Record during lab and use in Calculations section*

Part A

empty flask (g): ______________________

flask + water (g): ______________________

water temperature (°C): ______________________

density of water (g/mL):
Handbook or link ______________________

flask + unknown (g): ______________________

Unknown liquid (number): ______________________

Part B

empty flask (g): ______________________

flask + water (filled) (g): ______________________

water temperature (°C): ______________________

density of water (g/mL):
Handbook or link ______________________

flask + metal (g): ______________________

flask + metal + water (g): ______________________

Unknown solid (letter): ______________________

Notes:

<u>Part A Calculations</u>: ***Determining the Density of an Unknown Liquid***

Show all work, use significant figures and circle the final answer for full credit.

1. Using your data, calculate the mass (g) of water in the flask for Part A.

2. Using the mass of water in the flask (above) and the density of water, calculate the volume (mL) of the water in the flask.

3. Assuming that the water *completely* filled the flask (and it did!), determine the volume (mL) of the flask.

4. Using your data, calculate the mass (g) of the unknown in the flask.

5. Using the mass of the unknown in the flask (g), and assuming the unknown completely filled the flask (it did!), determine the density (g/mL) of the unknown liquid.

<u>Part B Calculations</u>: *Determining the Density of an Unknown Solid*

Show all work, use significant figures and circle the final answer for full credit.

1. Using your data, calculate the mass (g) of water in the flask for Part B.

2. Using the mass of water in the flask (above) and the density of water, calculate the volume (mL) of the water in the flask.

3. Assuming that the water *completely* filled the flask (and it did!), determine the volume (mL) of the flask.

4. Using your data, calculate the mass (g) of the unknown metal in the flask.

5. Using your data, calculate the mass (g) of water in the flask when the metal was present. *This is a different value from step 1 of part B!*

6. Convert the grams of water in the flask when the metal was present (step 5, above) into the volume of water (mL) present. Use the density from step 2, above.

Part B Calculations: *Continued*

7. Find the volume of the metal (cm^3) using the volume of the flask (step 3) and the volume of water present with the metal (step 6.)

8. Using the mass of the metal in the flask (g, step 4) and the volume of the metal (step 7), determine the density (g/cm^3) of the unknown metal.

Postlab Questions:

Show all work, use significant figures and circle the final answer for full credit.

1. In the original Indiana Jones movie, our hero is attempting to claim a precious ancient gold relic from a poor third world country. He estimates the size of his prize and carefully adjusts the *volume* of sand in his bag to equal that of the gold relic. With the dexterity that only Indiana Jones possesses, he swiftly but delicately swaps the sand for the gold. After a moment of delight, our hero realizes he has misjudged and the ancient tomb is not fooled. What went wrong? *You do not have to watch the Indiana Jones movie to answer this question!* ☺

2. Using the techniques covered in this lab, how can the volume of an irregularly shaped object that is less dense than water be found? Assume the object's density is unknown, and "forced submersion" or "weighted submersion" answers will not get credit.

<u>Postlab Questions</u>: ***Continued***

3. While panning for gold, you find a nugget that looks like gold. You find its mass to be 1.25g. You know that the density of pure gold is about 20.0 g/cm^3 and that the density of iron pyrite (fool's gold) is 5.0 g/cm^3. Determine if a cubic nugget about 0.40 cm on each side is fool's gold or pure gold. (Show all work)

4. Dennis obtained a clean, dry stoppered flask. He determined the mass of the flask and stopper to be 32.634 g. He then filled the flask with water and determined the mass of the full stoppered flask to be 59.479 g. Based on the temperature of the water, Dennis found the density of water in the *Handbook of Chemistry and Physics* to be 0.998730 g/cm^3. Calculate the volume of the flask.

5. Dennis emptied the flask from question #4, dried it and filled it with an unknown liquid. The mass of the stoppered flask when completely filled with liquid was 50.376 g. Calculate the density of the unknown liquid.

6. Dennis emptied the flask from question #4 and #5 and dried it again. He added an unknown metal to the flask. He determined the mass of the stoppered flask and metal to be 152.047 g. He then filled the flask with water, stoppered it and obtained a total mass of 165.541 g. Calculate the volume of metal added and the density of the unknown metal.

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CH 221 Fall 2024:
"Chemical Nomenclature"
(in class) ***Lab*** *- Instructions*

Note: **This is the lab for section 01 and H1 of CH 221 only.**

- *If you are taking section W1 of CH 221*, please use this link:
  http://mhchem.org/s/3b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of (at least) pages Ia-3-11 through Ia-3-15 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, October 7 (section 01) *or* **Wednesday, October 9 (section H1).** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work on your own, then **turn it in** (pages Ia-3-11 through Ia-3-15 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, October 14 (section 01)** *or* **Wednesday, October 16 (section H1.)** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Chemical Nomenclature

Chemical nomenclature is the system that chemists use to identify and name compounds. Compounds can have two types of names: *systematic names* (names that identify the chemical composition of a chemical compound) and *common names* (traditional names based on historical discovery or reactivity behavior). For example, N_2O has both a systematic name (dinitrogen monoxide) and a common name (laughing gas).

If every substance were assigned a common name, chemists would be expected to memorize over nine million names! This is why chemists generally prefer systematic names for identifying compounds. The International Union of Pure and Applied Chemistry (IUPAC, see http://www.iupac.com) was founded in 1921 to provide a system of chemical nomenclature for scientists. IUPAC nomenclature rules can provide valuable structural and reactivity information. On the other hand, most people would be hard pressed to call dihydrogen monoxide by any other name but water, so both types of nomenclature have their place.

Nomenclature leads naturally to formula writing. Compounds exist in distinct combinations of elements, and knowing the proper combinations of elements is essential in chemistry. We expect sodium chloride to be NaCl and not Na_2Cl or $NaCl_2$; knowing which combination or combinations exist in nature is crucial.

The following sections will guide you through the rules of *inorganic nomenclature* and formula writing. Later you may experience the nomenclature for organic chemistry or transition metal chemistry, but most of the compounds observed in first year chemistry will fall in this category.

Part A: Nomenclature of Elemental Ions

The first step in learning nomenclature is to learn the names of the elemental ions you might see in compounds. We make a distinction between the following:

- **fixed charge cations** (metal positive ions from groups IA, IIA and Al, Ga, In, Zn, Cd, Ag (“the stairs”)
- **variable charge cations** (positive ions which do not have a fixed charge; this includes ***all*** transition metals not in “the stairs”, lanthanides, actinides, Tl, Pb, Sn, and Bi)
- **anions** (negative ions which are generally nonmetals) have a charge equal to the **group number - 8**

Why two types of cations? Many metals have, for all practical purposes, only one ionic charge observed in nature. Lithium is only observed as Li^+ naturally, and even though gas phase studies of lithium ions have produced Li^{2+} and even Li^- ions, they are not observed in most settings. Many metals (such as iron) have many different *oxidation states* or *ionic charges* associated with them. The ions Fe^{2+}, Fe^{3+} and Fe^{6+} can be observed and manipulated quite readily (even at Mt. Hood Community College!); therefore, we need a method to distinguish between the various ions (namely iron(II), iron(III) and iron(VI), respectively).

Fixed Charge Cations use their elemental name.

Example:	Na^+ is the sodium ion	Cs^+ is the cesium ion
	Mg^{2+} is the magnesium ion	Sr^{2+} is the strontium ion
	Al^{3+} is the aluminum ion	In^{3+} is the indium ion

Variable Charge Cations use their elemental name followed by their ionic charge in parentheses. Use *Roman numerals* to distinguish the charge of the ion.

Example:	Fe^{2+} is the iron(II) ion	Pb^{2+} is the lead(II) ion
	Fe^{3+} is the iron(III) ion	Pb^{4+} is the lead(IV) ion
	Mn^{7+} is the manganese(VII) ion	Co^{9+} is the cobalt(IX) ion
	U^{4+} is the uranium(IV) ion	Ti^{2+} is the titanium(II) ion

Anions use their elemental name with the ending changed to *-ide*. *Notice:* charge = group number - 8

Example:	Cl^{-} is the chloride ion	I^{-} is the iodide ion
	O^{2-} is the oxide ion	Te^{2-} is the telluride ion
	N^{3-} is the nitride ion	As^{3-} is the arsenide ion

Part B: Nomenclature of Polyatomic Ions

Certain combinations of atoms result in stable configurations that are not easily destroyed; these are called *polyatomic ions*. Polyatomic ions can be either positive or negative, but most of them are anions (i.e. they have a negative charge.) Recognizing polyatomic ions in formulas is one of the most difficult concepts to master when learning nomenclature, and it is *very important that you* ***memorize*** *the following list of polyatomic ions*.

A list of polyatomic ions is given below:

nitrate	NO_3^-	**hydroxide**	OH^-	**hypochlorite**	ClO^-
nitrite	NO_2^-	**cyanide**	CN^-	**chlorite**	ClO_2^-
sulfate	SO_4^{2-}	**thiocyanide**	SCN^-	**chlorate**	ClO_3^-
sulfite	SO_3^{2-}	**cyanate**	OCN^-	**perchlorate**	ClO_4^-
phosphate	PO_4^{3-}	**thiosulfate**	$S_2O_3^{2-}$	**hypobromite**	BrO^-
phosphite	PO_3^{3-}	**chromate**	CrO_4^{2-}	**bromite**	BrO_2^-
hydrogen phosphate	HPO_4^{2-}	**dichromate**	$Cr_2O_7^{2-}$	**bromate**	BrO_3^-
dihydrogen phosphate	$H_2PO_4^-$	**permanganate**	MnO_4^-	**perbromate**	BrO_4^-
carbonate	CO_3^{2-}	**acetate**	$C_2H_3O_2^-$	**hypoiodite**	IO^-
hydrogen carbonate	HCO_3^-	**ammonium**	NH_4^+	**iodite**	IO_2^-
hydrogen sulfide	HS^-	**hydrogen**	H^+	**iodate**	IO_3^-
oxalate	$C_2O_4^{2-}$	**hydride**	H^-	**periodate**	IO_4^-

Part C: Nomenclature of Ionic Compounds from Ions

Knowing the nomenclature rules for ions, we can begin the naming of ionic compounds. Ionic compounds involve a *cation* (either *fixed* or *variable charge*) combining with an *anion*. Naming ionic compounds is straightforward; simply combine the ionic names with the cation first followed by the anion.

Example:	**sodium** ion + **chloride** ion give **sodium chloride**
	iron(III) ion + **bromide** ion gives **iron(III) bromide**
	ammonium polyatomic ion + **oxide** ion gives **ammonium oxide**
	aluminum ion + **sulfate** polyatomic ion gives **aluminum sulfate**

Part D: Writing Formulas for Ionic Compounds Using Nomenclature

Another important concept to master is the ability to write a chemical formula using the compound's systematic name. This can be accomplished using the following protocol:

1. Identify the elemental ions and/or polyatomic ions in the compound using the systematic name.
2. Determine the magnitude of the ionic charge on each ion
3. Assume the compound is electrically neutral *unless* the term "ion" appears in the name
4. The sum of the cation charges plus the anion charges must equal zero; combine the ions until this condition is met
5. Write the resulting formula. If more than one polyatomic ion is present, write the polyatomic portion in parentheses with a subscript after it denoting the number of polyatomic ions present.

Example: Write the formula for **sodium chloride**.

1. Sodium chloride has Na^+ and Cl^- ions
2. Sodium has a +1 charge, chloride has a -1 charge
3. Assume sodium chloride is neutral (no "ion" is present in the name)
4. Charge on sodium + charge on chloride = (+1) + (-1) = 0; therefore, **one** sodium ion and **one** chloride ion was required for a neutral compound.
5. 1 Na^+ ion and 1 Cl^- ion gives the formula **NaCl**

Example: Write the formula for **aluminum sulfide**.

1. Aluminum sulfide has Al^{3+} and S^{2-} ions
2. Aluminum has a +3 charge, sulfide has a -2 charge
3. Assume aluminum sulfide is neutral (no "ion" is present in the name)
4. Charge on aluminum + charge on sulfide = (+3) + (-2) = +1; this would indicate that combining one aluminum ion with one sulfide ion would give an *ion*. We want a *neutral* compound (see step 3); this can be accomplished by multiplying aluminum by **2** and sulfide by **3**, which results in: 2(+3) + 3(-2) = 0. Therefore, a neutral compound would result by combing **two** aluminum ions with **three** sulfide ions.
5. 2 Al^{3+} ions and 3 S^{2-} ions give the formula **Al_2S_3.**

Example: Write the formula for **magnesium nitrate**.

1. Magnesium nitrate has Mg^{2+} and NO_3^- ions
2. Magnesium has a +2 charge, nitrate has a -1 charge
3. Assume magnesium nitrate is neutral (no "ion" is present in the name)
4. Charge on magnesium + charge on nitrate = (+2) + (-1) = +1; this would indicate that combining one magnesium ion with one nitrate ion would give an *ion*. We want a *neutral* compound (see step 3); this can be accomplished by multiplying magnesium by **1** and nitrate by **2**, which results in: 1(+2) + 2(-1) = 0. Therefore, a neutral compound would result by combing **one** magnesium ion with **two** nitrate ions.
5. 1 Mg^{2+} ions and 2 NO_3^- ions give the formula **$Mg(NO_3)_2$.** (Note there are *two* nitrate ions, so they are placed in parentheses with a subscript two after it.)

Example: Write the formula for **titanium(IV) oxalate**.

1. Titanium(IV) oxalate has Ti^{4+} and $C_2O_4^{2-}$ ions
2. Titanium(IV) has a +4 charge, oxalate has a -2 charge
3. Assume titanium(IV) oxalate is neutral (no "ion" is present in the name)
4. Charge on titanium(IV) + charge on oxalate = (+4) + (-2) = +2; this would indicate that combining one titanium(IV) ion with one oxalate ion would give an *ion*. We want a *neutral* compound (see step 3); this can be accomplished by multiplying titanium(IV) by **1** and oxalate by **2**, which results in: 1(+4) + 2(-2) = 0. Therefore, a neutral compound would result by combing **one** titanium(IV) ion with **two** oxalate ions.
5. 1 Ti^{4+} ions and 2 $C_2O_4^{2-}$ ions give the formula $\mathbf{Ti(C_2O_4)_2}$**.** (Note there are *two* oxalate ions, so they are placed in parentheses with a subscript two after it.)

Part E: Finding Systematic Names for Ionic Compounds Using Formulas

Determining the systematic name of a compound from its formula is straightforward using these steps:

1. Identify the cation and anion in the formula. *Watch* for polyatomic ions.
2. Assume the compound is electrically neutral *unless* a charge appears in the formula
3. Determine the name of the anion and the charge on the anion
4. If a fixed charge cation is present, determine its name.
5. If a variable charge cation is present, determine its name and use this formula to find the charge on the metal: $\mathbf{charge_{metal} = -\ (\#\ anions)(charge_{anion})\ /\ (\#\ metal\ cations)}$ (where # = "number of")
6. Combine the cation and anion names as per Part C. The cation goes first, followed by the anion; do not forget the Roman numeral charge in parentheses for variable charge cations.

Example: Determine the name for **NaCl**.

1. The cation is Na and the anion is Cl
2. NaCl is neutral (no charges are present in the formula)
3. The anion, the chloride ion, has a -1 charge
4. Na is a fixed charge cation, and its name is the sodium ion
5. There are no variable charge cations in NaCl
6. The name of this compound is **sodium chloride**.

Example: Determine the name for $\mathbf{Sr(NO_3)_2}$**.**

1. The cation is Sr and the anion is NO_3.
2. $Sr(NO_3)_2$ is neutral (no charges are present in the formula)
3. The anion, the nitrate polyatomic ion, has a -1 charge
4. Sr is a fixed charge cation, and its name is the strontium ion
5. There are no variable charge cations in $Sr(NO_3)_2$
6. The name of this compound is **strontium nitrate**.

Example: Determine the name for **$Fe(NO_3)_3$.**

1. The cation is Fe and the anion is NO_3.
2. $Fe(NO_3)_3$ is neutral (no charges are present in the formula)
3. The anion, the nitrate polyatomic ion, has a -1 charge
4. There are no fixed charge cations in $Fe(NO_3)_3$
5. Iron is a variable charge cation; therefore, we must use the formula to calculate the charge on the iron atom. $charge_{Fe}$ = -(# nitrates)($charge_{nitrate+}$) / (# Fe atoms) = - (3)(-1) / (1) = **+3**; therefore, this is the **iron(III) ion**.
6. The name of this compound is **iron(III) nitrate**.

Example: Determine the name for **$Ru_3(PO_4)_2$.**

1. The cation is Ru and the anion is PO_4.
2. $Ru_3(PO_4)_2$ is neutral (no charges are present in the formula)
3. The anion, the phosphate polyatomic ion, has a -3 charge
4. There are no fixed charge cations in $Ru_3(PO_4)_2$
5. Ruthenium is a variable charge cation; therefore, we must use the formula to calculate the charge on the ruthenium atom. $charge_{Ru}$ = -(# phosphates)($charge_{phosphate}$) / (# Ru atoms) = - (2)(-3) / (3) = **+2**; therefore, this is the **ruthenium(II) ion**.
6. The name of this compound is **ruthenium(II) phosphate**.

Part F: Nomenclature for Binary Nonmetal Covalent Molecules

Not all compounds are ionic; indeed, many compounds *share* their electrons over the respective atoms. This class of compound is called *covalent*, and they are formed when two nonmetal elements combine.

The simplest covalent compounds are the elements that exist naturally in pairs; we refer to them as *diatomics*. These are crucial to a successful chemistry experience, and memorization is straightforward using the following acronym:

Name	Compound	Acronym
Hydrogen	H_2	*Have*
Nitrogen	N_2	*No*
Fluorine	F_2	*Fear*
Oxygen	O_2	*Of*
Iodine	I_2	*Ice*
Chlorine	Cl_2	*Clear*
Bromine	Br_2	*Brew*

In addition to the diatomics, several other nonmetals exist naturally in elemental form as combinations of more than one atom. **Phosphorus** exists naturally as P_4, and **sulfur** exists as S_8.

Most nonmetal covalent compounds have more than one type of element. Since there is no ionic charge present in these molecules, we cannot use the system developed above for ionic compounds, and a new method must be used. We will use the **Greek prefixes** for our compounds; they are:

1	*mono*	6	*hexa*
2	*di*	7	*hepta*
3	*tri*	8	*octa*
4	*tetra*	9	*nona*
5	*penta*	10	*deca*

The Greek prefixes refer to the number of atoms present in the molecule. For example, "dinitrogen" implies two nitrogen atoms since the prefix *di* stands for two.

When writing systematic names for binary nonmetal covalent compounds, use the *least electronegative atom first*. The topic of electronegativity will be discussed in Chem 222, but for now, the element listed first (either in the formula or the name) will be the least electronegative element.

Just as with cations in ionic compounds, use the normal element name for the least electronegative element. If more than one exist, use the Greek symbols to represent how many. The *most* electronegative element receives an *-ide* ending (as with anions in ionic compounds) as well as a Greek prefix, *even for single elements.* This is an important distinction between the most and least electronegative elements in binary compounds: the least electronegative element uses Greek symbols only if two or more atoms are present, while the more electronegative element gets an *-ide* ending *and* a Greek prefix *regardless* of the number of atoms present.

Examples:	**NO**	nitrogen monoxide
	N_2O	dinitrogen monoxide
	NO_2	nitrogen dioxide
	P_2O_3	diphosphorus trioxide
	P_2O_5	diphosphorus pentoxide

In addition, there are several common names of binary covalent compounds that you should be familiar with including the following:

Common Name	Formula	Systematic Name
water	H_2O	dihydrogen monoxide
ammonia	NH_3	nitrogen trihydride
laughing gas	N_2O	dinitrogen monoxide
nitric oxide	NO	nitrogen monoxide
phosphine	PH_3	phosphorus trihydride
hydrazine	N_2H_4	dinitrogen tetrahydride
hydrogen sulfide	H_2S	dihydrogen monosulfide

Part G: Nomenclature for Acids and Bases

Acid and base theory shall be discussed in detail during CH 223, but recognizing common acids and bases is important for all chemists. Acids and bases require water to become active; hence, Part G assumes all of the compounds mentioned have been dissolved in water.

Acids contain **H^+**, the **hydrogen ion**. Acids are created when hydrogen ions combine with halogens. If no oxygen atoms are present, add the *hydro-* prefix and an *-ic acid* suffix to find the acid name:

HBr	hydrobromic acid
HI	hydroiodic acid

If oxygen atoms are present in the halogen acid, use the following table:

Prefix and/or Suffix	Name	Formula
hydro-, -ic	hydrochloric acid	HCl
hypo-, -ous	hypochlorous acid	$HClO$
-ous	chlorous acid	$HClO_2$
-ic	chloric acid	$HClO_3$
per-, -ic	perchloric acid	$HClO_4$

Similar rules apply to bromide or iodide, but not fluoride.

Other common names for acids include:

$\mathbf{HNO_3}$	nitric acid	$\mathbf{H_3PO_4}$	phosphoric acid
$\mathbf{HNO_2}$	nitrous acid	$\mathbf{H_3PO_3}$	phosphorous acid
$\mathbf{H_2SO_4}$	sulfuric acid	$\mathbf{HC_2H_3O_2}$	acetic acid
$\mathbf{H_2SO_3}$	sulfurous acid	**HCN**	hydrocyanic acid
$\mathbf{H_2CO_3}$	carbonic acid	**HF**	hydrofluoric acid

More assistance with naming acids can be found in the handout, "**Guide to Common Polyatomic Ions and the Corresponding Acids**" available in the CH 221 Companion or on the CH 221 website.

One final note about acids: *technically*, an acid is only an acid if dissolved in water (i.e. if *aqueous*, with an *aq* state. If not in water, the acidic properties are lost (at least for CH 221!), and the compound should probably be written as either a binary nonmetal covalent molecule (Section F) or, if the acid contains a polyatomic ion, as a fixed charge metal with a nonmetal. Consider the following examples:

HCl(aq)	hydrochloric acid	*This is truly an acid since HCl is dissolved in water*
HCl(g)	hydrogen monochloride	*This is not an acid - no water! - so name this compound as a covalent compound*
$\mathbf{HNO_2(aq)}$	nitrous acid	*This is a true acid, dissolved in water*
$\mathbf{HNO_2(g)}$	hydrogen nitrite	*This is not an acid - no water! - so name this compound as a fixed charge metal + nonmetal due to the polyatomic ion (nitrite) present*

If a designation of state (i.e. aqueous, gas, solid, etc.) is not provided, then the naming system used is up to the observer (i.e. take your pick! ☺)

Bases contain $\mathbf{OH^-}$, the **hydroxide ion**. Bases consist of a metal cation with the hydroxide anion; hence, their nomenclature will be similar to that of Parts C, D and E, above.

Examples:	**NaOH**	sodium hydroxide
	$\mathbf{Fe(OH)_3}$	iron(III) hydroxide
	$\mathbf{NH_4OH}$	ammonium hydroxide

Part H: Final Words

Understanding chemical nomenclature rules and being able to write formulas for compounds can be thought of as learning to read and write a language. At first, the symbols and rules do not make much sense, but as time progresses, you master the language and a moment of euphoric inspiration occurs when "it all falls into place." Regrettably, inspiration only occurs after time has been spent practicing the material. The more you practice, the faster you will master the material.

Remember that there are five general classes of compounds:

Compound Class	Example
Fixed charge cation + anion	Al_2O_3 - aluminum oxide
Variable charge cation + anion	Fe_2O_3 - iron(III) oxide
Nonmetal binary covalent compound	P_2O_3 - diphosphorus trioxide
Acid	HIO_3 - iodic acid
Base	$Al(OH)_3$ - aluminum hydroxide

Each has specific rules to learn and master. Determining the charge of variable charge cations can be difficult at first, but application of the formulas in Part D and Part E should alleviate the distress.

...oh, wait, one more thing: **Waters of Hydration** or **Hydrated Compounds** show up occasionally with a "dot water" after the name of another chemical. If you see one, add the appropriate Greek prefix plus "hydrate." Examples of hydrated compounds:

$MgSO_4.6\ H_2O$ would be **magnesium sulfate hexahydrate**
$Cu(NO_3)_2\ .2\ H_2O$ would be **copper(II) nitrate dihydrate**
$Mn(BrO_3)_3.4\ H_2O$ would be **manganese(III) bromate tetrahydrate**

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Chemical Nomenclature Worksheet **Name:**

Complete the worksheets below and turn in on the due date.

Section One: Ion Names

Complete the chart using the appropriate elemental ion or polyatomic ion name or symbol. The first row has been filled in as an example. A list of polyatomic ions (page I-3-3) might prove helpful.

Ion	Name	Ion	Name
Na^{+}	sodium ion	F^{-1}	fluoride ion
Li^{+}	____________________	____________________	hydride ion
____________________	gold(III) ion	____________________	hydroxide ion
Mo^{3+}	____________________	____________________	cyanide ion
W^{2+}	____________________	SCN^{-1}	____________________
____________________	gold(I) ion	BrO^{-1}	____________________
Mn^{2+}	____________________	____________________	bromite ion
____________________	platinum(IV) ion	____________________	acetate ion
____________________	zirconium(II) ion	CrO_4^{2-}	____________________
Mt^{3+}	____________________	____________________	dichromate ion
Mg^{2+}	____________________	____________________	phosphide ion
____________________	vanadium(II) ion	____________________	phosphate ion
Cr^{3+}	____________________	____________________	phosphite ion
Cr^{2+}	____________________	$S_2O_3^{2-}$	____________________
____________________	tantalum(V) ion	IO_4^{-1}	____________________
Ni^{2+}	____________________	____________________	iodate ion
____________________	silver ion	____________________	hypoiodite ion
____________________	ammonium ion	MnO_4^{-1}	____________________

Section Two: Ions from Formulas

Write the ions that you would expect from the following compounds

Example: NaCl would give: Na^+, Cl^-

Example: $Fe_3(PO_4)_2$ would give: Fe^{2+}, PO_4^{3-}

LiBr would give:

$MgCl_2$ would give:

Na_2O would give:

VCl_2 would give:

$Fe(NO_3)_3$ would give:

$U(ClO_3)_4$ would give:

Section Three: Nomenclature from Ion Names

Complete the chart using the appropriate compound name using the ions given. The first row has been filled in as an example.

Cation	**Anion**	**Compound Name**
potassium	iodide	potassium iodide
magnesium	oxide	____________________
rhodium(III)	chloride	____________________
lead(IV)	chlorate	____________________
gold(I)	cyanide	____________________
cobalt(II)	nitrate	____________________
barium	hydroxide	____________________
ammonium	phosphate	____________________

Section Four: Writing Formulas Using Nomenclature

Complete the chart by providing the correct ion symbols (with the charge) and the correct formula for each compound. The first row has been filled in as an example.

Compound	Cation	Anion	Formula
calcium nitrate	Ca^{2+}	NO_3^{-1}	$Ca(NO_3)_2$
gallium bromide	____________	____________	____________________________
silver nitrate	____________	____________	____________________________
bismuth(III) chloride	____________	____________	____________________________
sodium acetate	____________	____________	____________________________
titanium(II) hypochlorite	____________	____________	____________________________
lithium permanganate	____________	____________	____________________________
iron(III) oxalate	____________	____________	____________________________
cesium chloride	____________	____________	____________________________

Section Five: Chemical Nomenclature Using Formulas

Complete the chart by providing the correct ion symbols (with the charge) and the correct name for each formula. The first row has been filled in as an example.

Formula	Cation	Anion	Name
$Ca(IO_3)_2$	Ca^{2+}	IO_3^-	calcium iodate
ZnS	____________	____________	______________________
$Sr_3(PO_3)_2$	____________	____________	______________________
$Ga_2(SO_4)_3$	____________	____________	______________________
$V(SCN)_5$	____________	____________	______________________
$NaMnO_4$	____________	____________	______________________
$(NH_4)_2S$	____________	____________	______________________
NH_4NO_2	____________	____________	______________________
$CrCl_6$	____________	____________	______________________

Section Six: Nonmetal Binary Covalent Compounds

Complete the chart by providing either the correct formula or name. The first row has been filled in as an example.

Name	Formula	Name	Formula
nitrogen dioxide	NO_2	phosphorus trichloride	PCl_3
______________________	SCl_4	sulfur hexachloride	______________
hydrogen monochloride	______________	______________________	$H_2S(g)$
______________________	PI_3	disulfur dichloride	______________
dinitrogen tetraoxide	______________	______________________	N_2O_3
antimony trichloride	______________	______________________	$SbCl_5$
______________________	SiO	carbon monoxide	______________
______________________	SiO_3	carbon dioxide	______________
phosphorus trihydride	______________	______________________	NO

Section Seven: Acids and Bases

Complete the chart by providing either the correct formula or name. The first entry has been filled in as an example. Use acid and base names only in this section.

Name	Formula	Name	Formula
hydrobromic acid	HBr	phosphoric acid	______________
______________________	HBrO	phosphorous acid	______________
bromous acid	______________	______________________	HCN
______________________	$HBrO_3$	acetic acid	______________
perbromic acid	______________	______________________	NaOH
sulfuric acid	______________	______________________	TiOH
______________________	H_2SO_3	potassium hydroxide	______________
______________________	HNO_3	iron(III) hydroxide	______________
nitrous acid	______________	______________________	$Mg(OH)_2$

Section Eight: Combined Problems: Complete the chart by providing either the correct formula or name.

Name	**Formula**
___________________	HCl(aq)
___________________	HCl(g)
potassium chloride	___________________
___________________	N_2O_4
nitrogen disulfide	___________________
___________________	$LiClO_3$
aluminum dichromate	___________________
___________________	$FeSO_4$
carbonic acid	___________________
___________________	SO_3
___________________	$(NH_4)_2CO_3$
potassium dihydrogen phosphate	___________________
potassium hydrogen phosphate	___________________
___________________	P_4O_{10}
___________________	$TbBr_6$
___________________	$ThBr_3$
___________________	TlBr
___________________	$TiBr_4$
___________________	$TeBr_2$
tetrasulfur decaoxide	___________________
sodium hydrogen carbonate	___________________
___________________	$In(C_2H_3O_2)_3$
___________________	$Mg(ClO_4)_2.6\ H_2O$

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CH 221 Fall 2024:

"Empirical Formula" *(in class)* ***Lab*** *- Instructions*

Note: **This is the lab for section 01 and H1 of CH 221 only.**

- *If you are taking section W1 of CH 221*, please use this link:

http://mhchem.org/s/4b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-4-2 through Ia-4-11 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, October 14 (section 01) *or* **Wednesday, October 16 (section H1).** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work and calculations on your own, then **turn it in** (pages Ia-4-7 through Ia-4-11 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, October 21 (section 01)** *or* **Wednesday, October 23 (section H1.)** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Empirical Formula

One of the fundamental statements of the atomic theory is that elements combine in simple whole number ratios. This observation gives support to the theory of atoms, since one would expect whole atoms to combine. Furthermore, it is observed the combining ratio for a given compound is constant regardless of the origin of the pure substance. This is known as the **Law of constant composition**. The mass contribution of each atom in a compound is a function of the number of atoms in the simplest formula and the relative mass of each atom. The mass contribution is usually referred to as the **percent composition** of a compound.

The **empirical formula** represents the smallest whole number ratio of atoms in a compound. The **molecular formula** represents the actual number of atoms in a compound. The molecular formula may be the same as the empirical formula or it may be a multiple of the empirical formula. For example, hydrogen peroxide has a molecular formula of H_2O_2 and an empirical formula of HO, while water has a molecular formula of H_2O and an empirical formula of H_2O. *No fractions should be present in empirical or molecular formulas!*

Because atoms combine in a definite ratio, the mass composition or **percent by mass** of a compound is fixed. You can determine the mass contribution of each element in a compound by the number of atoms in the simplest formula and the relative mass of each atom. This is typically referred to as the **percent composition** by mass of a compound. For example, $CuCO_3$ is always **38.847%** oxygen by mass:

$$\%\text{ Oxygen} = \left(\frac{3\text{ oxygen atoms x }15.999\text{ amu per oxygen}}{123.554\text{ amu total molar mass for }CuCO_3}\right)\text{ x }100\%$$

In this experiment you will determine the empirical formula of a hydrated compound of copper and chloride. You will first remove the water from the compound by heating the sample. Any mass lost is water. To determine the mass of copper in the compound, a simple exchange reaction with zinc is performed. Zinc is referred to as an active metal; in contact with a solution containing copper ions, the zinc metal will react to convert the copper ions into copper metal (zinc is transformed into zinc chloride.) As long as excess zinc is added, all of the copper ion should be removed as copper metal, which is easily massed. Any undetermined, remaining mass is chloride.

Once the mass of each component of the compound is determined (water, copper, chloride), you can calculate their corresponding moles and determine the empirical formula.

Because you are novice chemists, your experimental data will not be perfect. By following significant digit rules and using reasonable analytical deduction, you should be able to determine the formula of the hydrate. You can discuss your rounding and sources of error in your lab report.

In addition, you will be able to calculate the **theoretical yield** of copper based on the amount of zinc reacted. Comparing the **actual yield** of copper recovered to the theoretical yield, you can determine a **percent yield** of copper.

Operating a Bunsen Burner

Bunsen burners rely on the combustion of natural gas. An optimal mixture of gas (methane, CH_4) and air (oxygen) will produce a flame with an obvious blue (oxidizing) cone. **To set up your Bunsen burner:**

1. Attach one end of the rubber tubing to the sidearm at the base of the Bunsen burner and attach the other end to a gas outlet. (Be certain that the outlet is labeled "gas".)
2. Adjust the air (oxygen) intake to halfway (see picture below).
3. Adjust your ring stand to the correct height. The ideal flame should be only 2-3 cm above the burner. The ring and stand are metal and will be too hot to adjust once you begin heating. If you are unsure how to properly set up your ring stand, politely ask your instructor for assistance.
4. Make sure you can obtain a spark from the striker before proceeding.

To light your Bunsen burner you must do two things <u>simultaneously</u>:

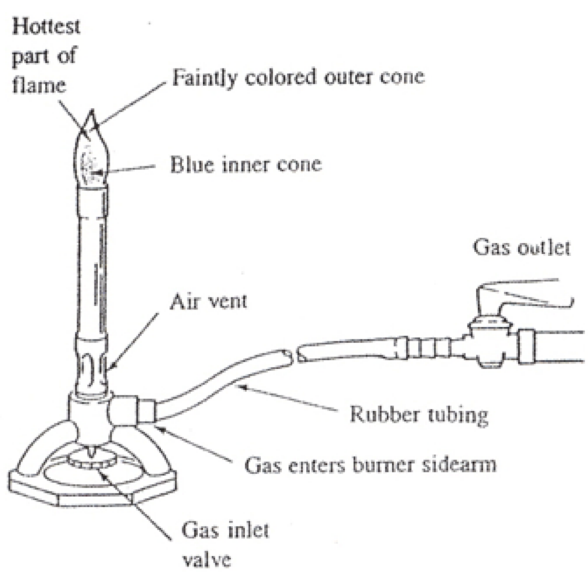
1. Open the gas inlet valve on the burner about halfway (see picture below).
2. As you open the gas outlet, light the burner by bringing the striker from the side to the top.

Do not leave the gas outlet open without the burner lit. It is unsafe to allow natural gas to enter the lab. Always check that the gas outlet is turned off when you are not using the burner. If you are unsure how to light the burner, please patiently ask your instructor for assistance.

To adjust your Bunsen burner:

Adjust the air (oxygen) intake until the flame becomes two concentric cones about 2-3 cm above the burner. The outer cone will be only faintly (dark blue) colored, but the inner cone will be a light blue color. The hottest part of the flame is at the tip of the inner light blue cone. If the flame is luminous and yellow or orange, but not blue, the air vent is not well adjusted. To adjust the height of the flame, adjust the gas flow at the gas outlet or at the gas inlet at the bottom of the Bunsen burner. Proper burner adjustment is crucial for good results. If you are unsure how to adjust the flame, please graciously ask your instructor for help.

Metal and ceramics hold heat well. Be careful when heating metal and ceramics as they will stay hot for quite a while after you turn the burner off. If unsure, remember that patience is a virtue, and wait longer for it to cool.



<u>PROCEDURE</u>: Part A: Dehydration of the copper compound

1. Record the mass of a clean, dry small crucible (no lid) in the "Data" section, below. ***All* measurements obtained in this lab should be to the nearest milligram (0.001g).**

2. Place approximately one gram (1.0 to 1.2 g) of unknown hydrated copper chloride in the crucible. Break up any sizable crystals with your spatula by pressing against the side of the crucible. Record the mass of the crucible (no lid) and sample to the nearest milligram.

3. Place the uncovered crucible on a clay triangle supported by a ring stand. Light your Bunsen burner away from the crucible, adjust the flame as described in the introduction, and ***gently*** heat the crucible as you move the burner back and forth judiciously. Flame should be small to avoid "popping" of sample. *If you overheat your sample*, it will turn into a black nasty liquid, and you must start over... so **heat gently!!!**

4. You should notice the crystals change color as they are heated. Record observations. Why do the crystals change color when heated? Continue to slowly heat the sample until all the crystals are brown. After about 5 minutes, carefully stir sample with a glass stir rod to check color. Once the entire sample is brown, gently heat for two additional minutes.

5. Turn off the burner. Cover the crucible as it cools to prevent the re-absorption of water vapor. Cool the crucible for about 10 minutes. *Caution:* The crucible is ceramic and retains heat. The crucible can severely burn you if you try to touch it before it is cool. Patience is a virtue!

6. After 10 minutes, remove the cover and slowly roll the brown crystals around the crucible. If there is any evidence of green crystals, repeat the heating and cooling process... but if all the crystals appear brown and the crucible is cool, **record the mass of the crucible (no lid) and dehydrated sample**.

Part B: Exchange reaction between the copper compound and zinc

1. Transfer the brown crystals to a small 250 mL beaker. Rinse the crucible with two 5-7 mL portions of deionized H_2O, adding each rinse to the beaker. Swirl the beaker gently to dissolve the crystals and record your observations. Why did the color change?

2. Obtain a piece of clean zinc and record its mass to the nearest 0.001 g. You need at least 0.5-0.8g of zinc; more is fine. Gently slide the piece of zinc into the beaker so that it is submerged in the copper chloride solution. Be careful not to splash. Add ~5 mL of water if your volume is too low.

3. Stir the solution with a glass rod so that as copper forms, it does not adhere to the zinc. Record observations. Allow the reaction to continue until all blue and green color has disappeared from the solution. The solution might have an unattractive grey hue, but no tint of green should remain.

4. Add 10 drops of 10% HCl to the solution and stir thoroughly. This will dissolve any insoluble zinc salts formed and clear up the solution if cloudy.

5. Carefully remove the unreacted zinc metal from the solution using tongs. Inspect the zinc for any adhering copper. Use a wash bottle of deionized water and a rubber policeman to scrape and clean the copper off the zinc into the beaker. Dry the remaining zinc on a paper towel. Record the mass of the dry zinc. (Note: this

must be less than your starting mass, right?) Place the zinc in the waste container when this step is complete.

Part C: Cleaning the copper:

1. Set up a Buchner funnel suction filtration apparatus with a moistened piece of filter paper. Attach the rubber hose to the “VAC” outlet, and only turn the vacuum to a 45 degree angle initially to prevent losing copper.

2. With light suction, carefully decant (pour off) the solution over the copper into the funnel. It is okay if some of the copper is transferred to the funnel.

3. Wash the copper solid in the beaker with about 10 mL of deionized water. Stir thoroughly, allow the copper to settle and carefully decant the wash water into the funnel. Break up any large chunks of copper with your glass stir rod. Repeat with a second 10 mL portion of deionized water.

4. Transfer the copper to the funnel using a small amount of deionized water. Use your wash bottle and rubber policeman to facilitate the transfer all of the copper to the funnel. Rinse any copper adhering to the rubber policeman into the funnel. All of the copper must be transferred to the funnel.

5. Turn off the suction. Add 10 mL of methanol to the funnel. After one minute, turn on the suction (slow at first, then to a roughly 45 degree angle.) Methanol evaporates faster than water and will enhance the drying process.

6. Draw air through the funnel for about 3-5 minutes. Meanwhile, record the mass of a clean, dry watch glass.

7. Transfer the dry copper to the massed watch glass. The transfer must be quantitative; scrape any copper that adheres to the paper on to the watch glass with your spatula or rubber policeman. If the copper is still damp, dry under a heat lamp for 5 minutes or press with a dry piece of filter paper. Allow the sample and watch glass to cool. Record the mass of the copper to the nearest 0.001 g. If you have more than 0.50 grams of copper, your sample is probably wet. It is recommended that you dry it under a heat lamp and take a new measurement.

8. Clean up! Dispose of the liquid methanol waste from the suction filtration apparatus into the appropriate waste bottle. Discard the copper in the garbage can unless directed otherwise by the instructor.

9. Complete the worksheets below using the data obtained in lab.

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Empirical Formula Worksheet | **Name:**

Complete the worksheets below and turn in on the due date | **Lab Partner(s)**:

All mass measurements should be made to the nearest milligram (0.001 g).

Mass of clean, dry, small crucible (g): ______________________

Mass of crucible and copper chloride sample (g): ______________________

Mass of unheated copper chloride sample (g): ______________________

Color of copper chloride sample before heating: ______________________

Mass of crucible and dehydrated copper chloride sample (g): ______________________

Mass of dehydrated copper chloride sample (g): ______________________

Mass of water lost upon heating (g): ______________________

Color of dehydrated copper chloride: ______________________

Color of dehydrated copper chloride sample when dissolved in water: ______________________

Mass of zinc sample before reaction (g): ______________________

Observations as zinc is added to the copper solution:

Mass of dry zinc after reacting with copper (g): ______________________

Mass of zinc that reacted with copper (g): ______________________

Mass of clean, dry watch glass (g): ______________________

Mass of dried copper and watch glass (g): ______________________

Mass of dried copper (g): ______________________

Color of dried copper: ______________________

Mass of chloride (g): *(dehydrated copper chloride - dried copper)* ______________________

Calculations Worksheet *for the* **Empirical Formula** *Lab*
Do not wait until the last minute to start these calculations!

1. Calculate the **mass (grams) of water** lost upon heating the copper chloride sample. Calculate the **moles of water lost** upon heating the hydrated copper compound

2. Calculate the **mass (grams) of copper** collected at the very end of lab; this is your **actual yield** of copper (g). Determine the **mass (grams) of chloride lost** in your sample (mass of dehydrated copper sample – mass of copper collected after filtration = mass of chloride.) Convert mass (grams) of copper into **moles of copper**; also convert mass (grams) of chloride into **moles of chloride.**

3. Use the moles of water, moles of copper and moles of chloride to **find the empirical formula of the hydrated copper chloride**. Round to whole numbers when determining the empirical formula.

4. Use the masses of water, copper, chloride, and the original hydrated copper chloride sample to **find the percent copper, percent chloride and percent of water in the original hydrated copper chloride sample**.

5. Show how to calculate the **mass (g) of zinc reacted** (i.e. the initial weight of Zn minus the final weight of Zn after the reaction was complete). Calculate the **theoretical yield of copper** using the formula:

   Theoretical yield (grams) of Cu = (grams of zinc reacted in the reaction) * 0.9720

6. Calculate the **percent yield** of copper. %yield = (actual yield / theoretical yield) x100%. **Comment** on why the percent yield might be greater than 100% in this lab.

POSTLAB QUESTIONS:

1. The *limiting reactant* (also known as the limiting reagent) is defined as the starting substance which is totally consumed in a reaction; the *excess reactant* is a starting material which is still present at the end of a reaction. Which of the reactants was the limiting reactant and which was the excess reactant? (the reactants in this reaction were zinc metal and the unknown copper chloride.) *Briefly* explain your answer.

2. Explain the color changes in part A (from blue to brown and back to blue) *and* in part B (from blue to clear).

   Part A:

   Part B:

3. **Explain** the effect each of the following would have on the experimentally determined %Cu. Use the terms **increase**, **decrease** or have **no effect** to describe the effect on the %Cu.
   a. Some solution splashed onto the bench when the zinc was plopped into the beaker.

   b. The student removed the zinc before the blue color disappeared from the solution.

   c. The student did not completely dry the Cu before the final weighing

4. Your final mass of zinc was less than your initial. What happened to the zinc? What did it become?

5. Determine the %Cl by mass value if the sample was pure anhydrous copper(II) chloride. *Hint:* do not use your data for this question; use the formula... what *is* the formula for copper(II) chloride?

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CH 221 Fall 2024:
“Percent Potassium Chlorate”
(in class) ***Lab*** *- Instructions*

Note: **This is the lab for section 01 and H1 of CH 221 only.**

- *If you are taking section W1 of CH 221*, please use this link:
  http://mhchem.org/s/5b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-5-2 through Ia-5-6 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, October 21 (section 01) *or* **Wednesday, October 23 (section H1).** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work and calculations on your own, then **turn it in** (pages Ia-5-3 through Ia-5-6 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, October 28 (section 01)** *or* **Wednesday, October 30 (section H1.)** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Percent Potassium Chlorate in a Mixture

Potassium chlorate ($KClO_3$) decomposes on heating to produce potassium chloride and oxygen. The Law of Conservation of Mass states that the mass of the reactants (potassium chlorate) will equal the mass of the products (potassium chloride and oxygen). Since oxygen is a gas, the mass of the final solid will be less than the starting weight. The mass loss is equal to the mass of oxygen.

In this experiment, you will begin with a sample that is a mixture of potassium chlorate and potassium chloride. Your objective is to determine the percentage by mass of potassium chlorate in the original mixture. Upon heating, only the potassium chlorate will decompose. Using the balanced equation and the fact that all the mass that is lost is oxygen gas, you can use stoichiometry to calculate the mass of potassium chlorate in the original mixture.

A catalyst, manganese(IV) oxide, is added to the reaction mixture in order to speed up the reaction. Like all catalysts, the same amount of catalyst is present at the end of the reaction as in the beginning. Therefore, we will include the mass of the catalyst in with the mass of the crucible.

To ensure that the decomposition is complete, the product must be heated to a constant weight. After the first heating, cooling and weighing, the sample must be heated again, cooled and reweighed. This process is continued until two successive weights are within 5 mg of each other (up to four heating cycles.)

PROCEDURE: *You **must** wear safety goggles while performing this lab! **All** mass measurements should be recorded to the milligram (0.001 g.)*

1. Set up a ring stand with a triangle as demonstrated by your teacher. The small white crucible should fit inside the triangle.

2. Place about 0.5 g of manganese(IV) oxide into a clean, dry small white crucible. Heat the crucible and catalyst with a Bunsen burner for about 3 minutes to drive off any moisture that may be in the catalyst or crucible. Wear safety glasses at all times if a Bunsen burner is operational at your lab bench!

3. When the crucible is cool enough to touch, record the entire mass to the nearest 0.001g.

4. Add between 2.0 to 2.5 grams of the unknown mixture to the crucible. Mix the contents to obtain a somewhat uniform mixture. Record the mass of the crucible plus catalyst plus mixture to the nearest 0.001 g. Be sure to also record your unknown number!

5. Begin heating the crucible gently at first followed by a more aggressive treatment for a total of 10 minutes. Be aware that the sample may begin to bubble and spurt; if this happens, turn the flame down a bit.

6. Allow the sample to cool to room temperature. Record the mass to the nearest 0.001 g.

7. Reheat your sample for 5 minutes. Cool and record the mass. If your mass is within 0.005 g of the mass after the previous heating with the unknown sample, congratulations, you can move on to calculations. If not, you should reheat, cool, and weigh until you have two successive masses within 0.005 g of each other. Clean up and put away your equipment (all waste in this lab can be washed down the drain with water.)

8. Clean up! Work on the remainder of the lab.

PERCENT POTASSIUM CHLORATE

Name:
Lab Partner(s):

DATA:

Unknown Number __________

1. Mass of crucible + catalyst
   (after drying moisture & before adding unknown ___________

2. Mass of crucible + catalyst + unknown
   (before heating) ___________

3. Mass after first heating ___________

4. Mass after second heating ___________

   Mass after third heating (if necessary) ___________

   Mass after fourth heating (if necessary) ___________

Lab Notes and Observations (if any):

CALCULATIONS: *Clearly **show all work** in the area provided, **watch significant figures** and circle final answers.*

1. Write a balanced equation for the decomposition of potassium chlorate into potassium chloride and oxygen gas.

2. Using your data, determine the **mass of the $KClO_3$ mixture** used in this experiment before heating (no MnO_2!).

3. Using your data, determine the **mass of oxygen lost** upon heating the mixture. This answer will be the **α** (below) in the equation.

4. Determine the **molar mass** of **oxygen** (**O_2**) to 0.01 g/mol. This answer will be the **β** (below) in the equation.

5. Determine the **molar mass** of **potassium chlorate** (**$KClO_3$**) to 0.01 g/mol. This answer will be the **δ** (below) in the equation.

6. Use the balanced equation and your values of **α** (the mass of oxygen lost), **β** (the molar mass of oxygen) and **δ** (the molar mass of potassium chlorate) to **determine the mass of potassium chlorate present in the original mixture** (this is the $KClO_3$ that decomposed in this experiment and is represented by **λ**, below, in the equation.) *Show your work! This is the "grams - moles - moles - grams' application we will be talking a lot about soon!* The equation to use:

$$\lambda \text{ g } \mathrm{KClO_3} = (\alpha \text{ g } \mathrm{O_2} \text{ lost}) * \left(\frac{1 \text{ mol } \mathrm{O_2}}{\beta \text{ g } \mathrm{O_2}}\right) * \left(\frac{2 \text{ mol } \mathrm{KClO_3}}{3 \text{ mol } \mathrm{O_2}}\right) * \left(\frac{\delta \text{ mol } \mathrm{KClO_3}}{1 \text{ mol } \mathrm{KClO_3}}\right)$$

7. **Determine the percentage of potassium chlorate in your unknown** using your answers from step 6 (the pure KClO3) and step 2 (the mass of the original mixture.)

POSTLAB QUESTIONS:

1. A white powder is a mixture of magnesium carbonate and magnesium oxide. Upon heating, the magnesium carbonate decomposes into magnesium oxide and carbon dioxide. If you have 1.897 g of the mixture and after heating are left with 1.494 g of magnesium oxide, calculate the weight percent of magnesium carbonate in the original mixture. *Hint:* Start by writing a balanced reaction, and remember the 1.897 g value is not pure!

2. Calculate the % oxygen by mass for the following (show calculations): a) $LiNO_3$ b) $NaHCO_3$ *Hint:* first find the molar mass (to 0.01 g/mol) of the compound!

3. If we had doubled the mass of the original mixture and completed the lab as written, would the calculated %$KClO_3$ have changed? Explain.

CH 221 Fall 2024:

"Net Ionic Reactions" *(in class)* ***Lab*** *- Instructions*

Note: **This is the lab for section 01 and H1 of CH 221 only.**

- *If you are taking section W1 of CH 221*, please use this link:
  http://mhchem.org/s/6b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-6-2 through Ia-6-12 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, October 28 (section 01) *or* **Wednesday, October 30 (section H1).** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work and calculations on your own, then **turn it in** (pages Ia-6-9 through Ia-6-12 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, November 4 (section 01)** *or* **Wednesday, November 6 (section H1.)** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Net Ionic Reactions in Aqueous Solutions

Double replacements are among the most common of the simple chemical reactions. Consider the hypothetical reaction:

$$\mathbf{AB + CD \rightarrow AD + CB}$$

where AB exists as A^+ and B^- ions in solution and CD exists as C^+ and D^- ions in solution. As the ions come in contact with each other, there are six possible combinations that might conceivably cause a chemical reaction. Two of these combinations are the meeting of ions of like charge; that is, A^+ with C^+ and B^- with D^-. Since particles with like electrical charges repel each other, no reaction will occur. Two other possible combinations are those of the original two compounds; that is A^+ with B^- and C^+ with D^-. This combination would lead to no change. Thus the only possibilities for chemical reaction are the combination of each of the positive ions with the negative ion of the other compound; that is, A^+ with D^- and C^+ with B^-.

Example 1: When solutions of sodium chloride and silver nitrate are mixed, the combination of silver cations and chloride anions form silver chloride, which precipitates and settles to the bottom of the container. Note that the states of matter are included: (aq) substance is soluble in water; (s) substance is insoluble in water (solid precipitate)

$$\mathbf{NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)}$$

This combination – the **molecular equation** - of chemicals is referred to as a **precipitation reaction** since an insoluble solid, AgCl, is present as a product.

Example 2: When solutions of potassium chloride and sodium nitrate are mixed, the equation for the hypothetical double replacement reaction is:

$$\mathbf{KCl(aq) + NaNO_3(aq) \rightarrow KNO_3 + NaCl}$$

But has there been a reaction? Double replacement reactions occur when one of the following is formed as a product of the reaction:

a. an **insoluble solid** (precipitate) - check the solubility table in this lab report. If a solid has formed, this is called a **precipitation reaction**.
b. a **gas** - for example, CO_2 (from H_2CO_3), SO_2 (from H_2SO_3), or NH_3 (from NH_4OH). If a gas has formed, this is called a **gas forming reaction**.
c. **water** from an acid (source of H^+) and a base (source of OH^{-1}). If water forms from an acid and a base (along with an ionic "salt"), this is called an **acid-base reaction**.

Using the solubility table (see below) we find both KNO_3 and NaCl are water-soluble (aqueous, or *aq* for short) products. There is no precipitate, gas or water from this combination. Thus in Example 2, we conclude that even though we can write an equation for a double replacement reaction, no reaction occurs. We simply end up with a solution containing four kinds of ions - Na^+, K^+, Cl^-, and NO_3^-.
Thus the **molecular equation** is more properly written:

$$\mathbf{KCl(aq) + NaNO_3(aq) \rightarrow KNO_3(aq) + NaCl(aq)}$$

but in terms of "if something happens", we should write:

$$\mathbf{KCl(aq) + NaNO_3(aq) \rightarrow No\ Reaction}$$

Aqueous solutions of sodium chloride and silver nitrate will undergo double replacement reaction to produce a white precipitate of silver chloride and aqueous sodium nitrate. What would happen if we just mixed solid silver nitrate and solid sodium chloride together? No apparent reaction occurs. Thus the water performs some necessary function that allows the reaction to proceed. When ionic compounds are dissolved in water, the ions separate and become surrounded by water molecules. This frees the ions from the crystal lattice, allowing them to move throughout the solution and react with appropriate ions of opposite charge.

To clarify what reaction occurs between ions in electrolyte solutions, we write **total ionic equations**. In this type of equation, compounds are written in the form in which they are predominately present in water. Most notably, soluble compounds (aq) are written as ions in solution. Others (s, l, g) are written in their molecular form.

For example, if we write the <u>total ionic equation</u> for the double replacement precipitation reaction (See Example 1) we get the following:

<u>Total Ionic Equation:</u> $Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + AgCl(s)$

Note that during the course of reaction, there has been no change in the Na^+ and NO_3^- ions. These unreacted ions (**spectator ions**) can be left out of the total ionic equation to yield the **net ionic equation**. <u>Net ionic equations</u> tell us only what is actually changing during reaction.

<u>Net Ionic Equation</u>: $Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$

Another example is illustrated below for the reaction of nitric acid and a dilute aqueous solution of barium hydroxide (an **acid-base reaction**):

<u>Molecular Equation</u>: $2\ HNO_3(aq) + Ba(OH)_2(aq) \rightarrow 2\ H_2O(l) + Ba(NO_3)_2(aq)$

<u>Total Ionic Equation</u>: $2\ H^+(aq) + 2\ NO_3^-(aq) + Ba^{2+}(aq) + 2\ OH^-(aq) \rightarrow 2\ H_2O(l) + Ba^{2+}(aq) + 2\ NO_3^-(aq)$

<u>Net Ionic Equation:</u> $2\ H^+(aq) + 2\ OH^-(aq) \rightarrow 2\ H_2O(l)$
or, reduced to simplest form: $\mathbf{H^+(aq) + OH^-(aq) \rightarrow H_2O(l)}$
This is an example of an **acid-base reaction**.

We will use the following solubility table in CH 221:

CH 221 Solubility Table for Ionic Compounds

SOLUBLE COMPOUNDS	EXCEPTIONS
Almost all salts of Na^+, K^+, NH_4^+	
Salts of nitrate, NO_3^- chlorate, ClO_3^- perchlorate, ClO_4^- acetate, $CH_3CO_2^-$	
Almost all salts of Cl^-, Br^-, I^-	Halides of Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds containing F^-	Fluorides of Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
Salts of sulfate, SO_4^{2-}	Sulfates of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

INSOLUBLE COMPOUNDS	EXCEPTIONS
Most salts of carbonate, CO_3^{2-} phosphate, PO_4^{3-} oxalate, $C_2O_4^{2-}$ chromate, CrO_4^{2-}	Salts of NH_4^+ and the alkali metal cations
Most metal sulfides, S^{2-}	
Most metal hydroxides and oxides	

Note: **Use this table for *all* CH 221 solubility questions!**

The following examples should help you understand how the solubility table works and also how to complete the written part of this assignment.

Example: Is $PbSO_4$ soluble in water? What species are present in a water solution?

> *Answer:* To solve, notice how $PbSO_4$ has a sulfate ion (SO_4^{2-}). Most salts of sulfate are soluble (i.e. they dissolve in water, or are (aq)), but salts of sulfate with a Pb^{2+} ion are an exception to this rule. Hence, $PbSO_4$ is insoluble (it does not dissolve in water), and we would write $PbSO_4$ in water as $PbSO_4(s)$. Indeed, $PbSO_4(s)$ is the only species present in a water solution.

Example: Is Li_2S soluble in water? What species are present in a water solution?

> *Answer:* To solve, notice how Li_2S has a sulfide ion (S^{2-}). Most salts of sulfide are insoluble (i.e. they do not dissolve in water, or are (s)), but salts of alkali metals (which includes Li^+) are an exception to this rule. Hence, Li_2S is soluble (it *does* dissolve in water), and we would write Li_2S in water as $Li_2S(aq)$.
>
> Further, because Li_2S dissolves in water, we should really write it as the dissociation ions – i.e. the molecular Li_2S dissociates into lithium and sulfide ions. The species which are present in a water solution of $Li_2S(aq)$ are 2 $Li^+(aq)$ and $S^{-2}(aq)$ (molecular Li_2S does not exist in water.)

Example: Describe $CaCO_3$, $RbMnO_4$ and $TiCrO_4$ in terms of their solubility in water.

> *Answer:* Use the solubility table to answer these types of questions:
>
> - **$CaCO_3$** has a carbonate ion, and calcium is not an alkali metal or ammonium, so CaCO3 is insoluble in water. We would write it as $CaCO_3(s)$.
> - **$RbMnO_4$** has an alkali metal ion (rubidium), so by default, all alkali metals are water soluble (no exceptions, at least in CH 221!) So we would write this species as $RbMnO_4(aq)$ or, as dissolve ions, we would write it as $Rb^+(aq)$ and $MnO_4^-(aq)$.
> - **$TiCrO_4$** has a chromate ion, and titanium is not an alkali metal or ammonium, so $TiCrO_4$ is insoluble in water. We would write it as $TiCrO_4(s)$.

Example: Write the balanced molecular equation and net ionic reaction that occurs between potassium nitrate and calcium chloride in water. Classify this reaction type.

> *Answer:* First, we need the chemical equations for potassium nitrate and lithium chloride. They are KNO_3 and $CaCl_2$. Notice the ionic charges on the cations and anions: K^+, NO_3^{-1}, Ca^{2+}, Cl^{-1}.
>
> All of the reactions in this lab are "double displacement" – the reactant cations will switch places, forming new products. *The ionic charges will not change upon going from reactant to product*. Potassium and chloride will come together as KCl (only one Cl^{-1} for every one K^{+1}), and calcium and nitrate will come together as $Ca(NO_3)_2$ (two nitrates being needed for every calcium +2 ion.) Initially, the equation looks like this:
>
> $$KNO_3 + CaCl_2 \rightarrow KCl + Ca(NO_3)_2$$

Notice the parentheses used for more than one polyatomic ion ($Ca(NO_3)_2$) but parentheses are not used when only one polyatomic ion is used (KNO_3).

We need to balance this reaction and add states of matter. Every compound with potassium (an alkali metal) or nitrate will dissolve in water; $CaCl_2$ is also soluble in water (Ca is not Ag, Pb or Hg), leading to:

$$\mathbf{2\ KNO_3(aq) + CaCl_2(aq) \rightarrow 2\ KCl(aq) + Ca(NO_3)_2(aq)}$$

To *classify* this reaction, our options include: precipitate, acid-base, gas forming, or no reaction. Since no solids have formed, it is not a precipitation reaction. Water has not formed from an acid or base, so this is excluded; and H_2CO_3 and NH_4OH have not formed (see next example), so gas forming is excluded. Indeed, all the reactants and products are (aq), so nothing really happens; classify this reaction as "**no reaction**." Nothing needs to be written for a net ionic reaction because nothing happens!

Example: Write the balanced molecular equation and net ionic reaction that occurs between potassium carbonate and hydrobromic acid in water. Classify this reaction type.

Answer: First, we need the chemical equations for the reactants. They are K_2CO_3 and HBr, and they make K^+, CO_3^{-2}, H^{1+}, Br^{-1}. Performing a "double displacement" on these reactants – the reactant cations will switch places – we get: $K_2CO_3 + HBr \rightarrow KBr + H_2CO_3$

Balancing this reaction and adding states of matter, we get:

$$\mathbf{K_2CO_3(aq) + 2\ HBr(aq) \rightarrow 2\ KBr(aq) + H_2CO_3(aq)}$$

At first, it looks like this is a "no reaction" classification – all states are aqueous – but **make sure you check for H_2CO_3 and NH_4OH** – these two species are the hallmarks of the **gas forming reaction** since both are unstable compounds and further decompose to new products. ***Be watchful for H_2CO_3 and NH_4OH!***

Carbonic acid (H_2CO_3 breaks down into water and carbon dioxide, so really you should write:

$$\mathbf{H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)}$$

and ammonium hydroxide (NH_4OH), which breaks down into water and ammonia, should be written:

$$\mathbf{NH_4OH(aq) \rightarrow H_2O(l) + NH_3(g)}$$

So in this **gas forming** example,

$$\mathbf{K_2CO_3(aq) + 2\ HBr(aq) \rightarrow 2\ KBr(aq) + H_2CO_3(aq)}$$

should be written as a net ionic equation in the following manner:

$$\mathbf{CO_3^{2-}(aq) + 2\ H^+(aq) \rightarrow H_2O(l) + CO_2(g)}$$

PROCEDURE and LAB REPORT:

Use the attached sheets to complete this week's lab. For each reaction,

- **Mix** 1.0 mL (20 drops) of each of the two indicated solutions (below) in a clean (but not necessarily dry) **small** test tube and record observations that might indicate a chemical change has occurred **(color, precipitate, bubbles of a gas, or heat released.)**

- Write the **balanced molecular equation** (double displacement or exchange reaction) for each reaction. Show **states of matter** (**use the solubility table in this lab report for your answers**) and **ionic charges (on ions, *not* molecules)** for all species.

- Write the **total ionic equation** and the **net ionic equation** for each reaction. Be sure to include all states of matter and ionic charges. If all the products are aqueous, no reaction has occurred, and you should write **no reaction** in place of the net ionic equation. Note that even if no reaction occurs, you will still be required to write a balanced molecular equation and the total ionic equation.

- Finally, **classify each reaction** as **precipitation, acid-base, gas forming or** (if nothing happened) **no reaction**. Remember that gas forming reactions often create unstable precursors (such as $\mathbf{H_2CO_3}$ (which creates CO_2(g) and H_2O(l)) and $\mathbf{NH_4OH}$ (which creates NH_3(g) and H_2O(l)).)

The reactions:

1. Barium Nitrate + Magnesium Sulfate
2. Barium Nitrate + Hydrochloric Acid
3. Barium Nitrate + Sodium Carbonate
4. Iron(III) Chloride + Sodium Hydroxide
5. Iron(III) Chloride + Sodium Phosphate
6. Iron(III) Chloride + Magnesium Sulfate

7. Magnesium Sulfate + Sodium Hydroxide
8. Magnesium Sulfate + Sodium Carbonate
9. Ammonium Oxalate + Barium Nitrate
10. Hydrochloric Acid + Sodium Hydroxide
11. Hydrochloric Acid + Sodium Carbonate
12. Silver Nitrate + Potassium Chromate

13. Silver Nitrate + Iron(III) Chloride
14. Sodium Hydroxide + Ammonium Chloride
15. Sodium Hydroxide + Sulfuric Acid
16. Copper(II) Sulfate + Iron(III) Chloride
17. Copper(II) Sulfate + Sodium Phosphate
18. Acetic Acid + Sodium Carbonate

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Net Ionic Reactions ***Worksheet*** **Name:**

Lab Partner(s):

Complete the following worksheet using the instructions provided. Remember to show states of matter and charges where appropriate. **M** = **Molecular Equation**, **T** = **Total Ionic Equation**, and **N** = **Net Ionic Equation**.

---

1. Barium Nitrate + Magnesium Sulfate Observations: ____________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

---

2. Barium Nitrate + Hydrochloric Acid Observations: ____________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

---

3. Barium Nitrate + Sodium Carbonate Observations: ____________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

---

4. Iron(III) Chloride + Sodium Hydroxide Observations: ____________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

5. Iron(III) Chloride + Sodium Phosphate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

6. Iron(III) Chloride + Magnesium Sulfate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

7. Magnesium Sulfate + Sodium Hydroxide Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

8. Magnesium Sulfate + Sodium Carbonate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

9. Ammonium Oxalate + Barium Nitrate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

10. Hydrochloric Acid + Sodium Hydroxide Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

11. Hydrochloric Acid + Sodium Carbonate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

12. Silver Nitrate + Potassium Chromate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

13. Silver Nitrate + Iron(III) Chloride Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

14. Sodium Hydroxide + Ammonium Chloride Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

15. Sodium Hydroxide + Sulfuric Acid Observations: ____________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

16. Copper(II) Sulfate + Iron(III) Chloride Observations: ______________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

17. Copper(II) Sulfate + Sodium Phosphate Observations: ______________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

18. Acetic Acid + Sodium Carbonate Observations: ______________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

Bonus! On a *separate piece of paper*, add an original poem or short story of at least 50 words in length for extra credit… content will not be criticized, but the poem must be original, and short haikus written at the bottom of this page will not count (although the instructor will find them fun to read! ☺) Original music will also count!

CH 221 Fall 2024:

"Unknown Chloride" *(in class)* ***Lab*** *- Instructions*

Note: **This is the lab for section 01 and H1 of CH 221 only.**

- *If you are taking section W1 of CH 221*, please use this link:
  http://mhchem.org/s/7b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-7-2 through Ia-7-10 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, November 4 (section 01) *or* **Wednesday, November 6 (section H1).** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work and calculations on your own, then **turn it in** (pages Ia-7-5 through Ia-7-10 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, November 18 (section 01** *due to Veterans Day*) *or* **Wednesday, November 13 (section H1.)** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Determination of an Unknown Chloride

The determination of a soluble chloride salt concentration is a classic titrametric analysis. A titration involves delivering a measured amount of a solution whose concentration is known accurately (the **titrant**) into a solution whose concentration is not known (the **titrate**). The purpose of the titration is to determine the number of moles of titrate present. When the reaction is complete, some physical change is observed, indicating the **endpoint** of the titration. The endpoint of a titration occurs when stoichiometric ratios of reactants are present and must be determined accurately.

In the titration in this lab, a dilute solution of silver nitrate with a known concentration acts as the titrant. It is added to a salt solution with an unknown amount of chloride, i.e. the titrate. Silver chloride, a white insoluble solid will precipitate from the solution. In order to detect when all the AgCl has been precipitated, another reagent is used as an **indicator**. The indicator in this lab, potassium chromate, is yellow and reacts with silver ions to form a bright orange silver chromate precipitate. This solid is slightly more soluble than the silver chloride so it does not form until essentially all the chloride has precipitated from the solution.

$$\mathbf{Ag^{+}(aq) + Cl^{-}(aq) \rightleftharpoons AgCl(s)} \text{ at } \textbf{low} \text{ silver ion concentration}$$

$$\mathbf{2\,Ag^{+}(aq) + CrO_4^{2-}(aq) \rightleftharpoons Ag_2CrO_4(s)} \text{ at } \textbf{higher} \text{ silver ion concentration}$$

All standard solutions must first be standardized using a **primary standard** because of potential evaporation. A primary standard is a solid that is stable and does not pick up water. The primary standard in this experiment is purified sodium chloride.

In this lab you will perform six titrations. In the first three titrations you will use a known amount of a pure NaCl sample to determine the exact concentration of an approximate 0.05 M silver nitrate solution. The endpoint is the first permanent orange-red color of Ag_2CrO_4. From this information one can determine the concentration of the $AgNO_3$. The last three titrations will allow you to find the percentage of chloride in your salt when used in conjunction with the average silver nitrate concentration.

Note: Silver is a heavy metal toxin and should never be flushed down the drain. Dispose of all silver waste (silver nitrate and silver chloride) in the waste bottles provided.

PROCEDURE:

Part A: ***Standardizing the Silver Nitrate Solution***

1. Clean a 50 mL buret with soap and water, then rinse well with water.

2. Fill your buret with silver nitrate from an amber bottle. To prevent contamination, ***never*** add anything to the amber bottle. Fill the buret to the 0.00 mL mark with the $AgNO_3$ solution. Drain 5 mL from the buret into a beaker (to remove air bubbles) and fill to 0.00 mL again. *Note:* $AgNO_3$ is the only solution that will be placed in your buret! Also, do not dispose of $AgNO_3$ in the sink – place this heavy metal in a waste container.

3. Use an analytical balance to weigh three 0.1000 - 0.1200 gram samples of purified NaCl. Record exact mass.

4. Add about 50 mL of distilled water to each sample in a 125 mL Erlenmeyer flask (or larger) to dissolve the NaCl sample. Add about three drops of indicator (K_2CrO_4).

5. Titrate with 0.05 M $AgNO_3$ solution as you continually swirl the flask to a lovely peach end point. As you add the silver nitrate solution initially in short bursts you will see the orange-red color form and disappear as the solution is swirled. As you approach the end point (which should be between 20-40 mL) the color should begin to persist. At this point you should be adding the solution dropwise. Read the buret to the nearest 0.01 mL. Stop when the sample has a permanent faint peach color.

6. Repeat the titration with the second and third samples.

Part B: ***Determination of Percent Chloride***

1. Obtain an unknown chloride salt and record the ID number in your lab notes. Use an analytical balance to weigh **three** 0.1000 - 0.1200 gram samples.

2. To **each** sample add 50 mL of distilled water and 3 drops of K_2CrO_4 indicator solution in an Erlenmeyer flask. Titrate each sample with the standardized silver nitrate solution as in part A.

3. When done, place excess $AgNO_3$ in the waste container and rinse the buret with water before leaving the lab.

CALCULATIONS:

For Part A, calculate the molarity of the silver nitrate solution ([$AgNO_3$]) for each titration. Calculate the Parts Per Thousand (PPT) for [$AgNO_3$] using the "Parts Per Thousand" handout in the “Lab Notes” of the Companion. If your PPT is greater than 30 for the three trials, consider omitting a deviant molarity value to improve your PPT.

For Part B, calculate the percent chloride. (*Note:* Use the average molarity of $AgNO_3$ as determined in part A.) Average your three percent chloride values and find the PPT for the %Cl values. As in Part A, if one trial is quite different from the other two, report data from all three trials, but only average two trials.

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Determination of an Unknown Chloride

All masses should be recorded to 0.0001 g and volumes to 0.01 mL

YOUR NAME: _______________________ ***LAB PARTNER(s):*** _________________

DATA: ***Part A***

NaCl sample #1 (g): _________________

Volume buret (initial, mL): _________________

Volume buret (final, mL): _________________

Volume $AgNO_3$ used for sample #1 (mL): _____________

NaCl sample #2 (g): _________________

Volume buret (initial, mL): _________________

Volume buret (final, mL): _________________

Volume $AgNO_3$ used for sample #2 (mL): _____________

NaCl sample #3 (g): _________________

Volume buret (initial, mL): _________________

Volume buret (final, mL): _________________

Volume $AgNO_3$ used for sample #3 (mL): _____________

Lab Notes and Observations:

DATA: *Part B*

Unknown Sample ID = ______________

Unknown sample #1 (g): _______________

Volume buret (initial, mL): _______________

Volume buret (final, mL): _______________

Volume $AgNO_3$ used for sample #1 (mL): ___________

Unknown sample #2 (g): _______________

Volume buret (initial, mL): _______________

Volume buret (final, mL): _______________

Volume $AgNO_3$ used for sample #2 (mL): ___________

Unknown sample #3 (g): _______________

Volume buret (initial, mL): _______________

Volume buret (final, mL): _______________

Volume $AgNO_3$ used for sample #3 (mL): ___________

Lab Notes and Observations:

<u>Part A Calculations</u>: *Determination of an Unknown Chloride*

Show all work, use significant figures and circle the final answer for full credit.

Use the space below to **calculate three values for the molarity of the $AgNO_3$ solution**. The lab video (http://mhchem.org/w/7.htm) offers help with this calculation, but use your own data, not the data provided in the video. *Show all work! Watch significant figures!*

Using the three calculated values of [$AgNO_3$] from above, calculate the **average concentration** of $AgNO_3$, the **average deviation** of your molarity calculations and the precision in **parts per thousand**. (The parts per thousand handout can be found here: **http://mhchem.org/ppt**)

Part B Calculations: ***Determination of an Unknown Chloride***

Show all work, use significant figures and circle the final answer for full credit.

Use the space below to **calculate three values for the %Cl of your unknown sample**. The lab video (http://mhchem.org/w/7.htm) offers help with this calculation, but use your own data, not the data provided in the video. *Show all work! Watch significant figures!*

Using the three calculated values of %Cl from above, calculate the **average %Cl**, the **average deviation** of your %Cl calculations and the precision in **parts per thousand**. (ppt handout: http://mhchem.org/ppt)

<u>Part C Postlab Questions</u>: ***Determination of an Unknown Chloride***

Show all work, use significant figures and circle the final answer for full credit.

1. 35.46 mL of a silver nitrate solution was used to reach the chromate end point with a 50 mL solution containing 0.1165 g of pure NaCl. What is the molarity of the $AgNO_3$ solution?

2. How many mL of the silver nitrate solution used in question 1 above will react with 0.2595 g of $BaCl_2$ dissolved in 50 mL of water?

3. A solid chloride sample weighing 0.09969 g required 18.25 mL of 0.05205 M $AgNO_3$ to reach the chromate end point. What is the % chloride in this sample?

<u>Part C Postlab Questions</u>: ***Determination of an Unknown Chloride - continued***

Show all work, use significant figures and circle the final answer for full credit.

4. How would the following hypothetical errors affect the calculated % chloride (increase, decrease or no change)? Explain.

   a. The pure sodium chloride was left open in the scale room and absorbed moisture.

      This will **increase** **decrease** **not change** the percent chloride. ***(circle one)***

      Explain your answer:

   b. The calculated molarity of the silver nitrate solution was 5% too high.

      This will **increase** **decrease** **not change** the percent chloride. ***(circle one)***

      Explain your answer:

   c. Two mL of $AgNO_3$ are added beyond the chromate end in titrating the unknown chloride.

      This will **increase** **decrease** **not change** the percent chloride. ***(circle one)***

      Explain your answer:

CH 221 Fall 2024:
"Calorimetry" *(in class)*
Lab - Instructions

Note: **This is the lab for section 01 and H1 of CH 221 only.**

- *If you are taking section W1 of CH 221*, please use this link:
  http://mhchem.org/s/8b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-8-2 through Ia-8-14 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, November 18 (section 01) *or* **Wednesday, November 13 (section H1).** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work and calculations on your own, then **turn it in** (pages Ia-8-7 through Ia-8-14 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, November 25 (section 01)** *or* **Wednesday, November 20 (section H1.)** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Calorimetry

Thermal energy, usually called heat, is one of the most familiar forms of energy. We observe this form of energy when it is passed from an object of higher temperature to one at a lower temperature when the objects are in contact. Heat flow can be measured using a device called a **calorimeter**. A calorimeter is a device that is insulated from the surroundings so that essentially no heat can flow in or out of the device. Within the calorimeter, heat can be transferred from one system to another and this transfer causes a temperature change.

Specific Heat of an Unknown Metal

When heat flows into a substance, the temperature of the substance increases. The quantity of heat, **q**, required to cause a temperature change **ΔT** of any substance varies directly with the mass, **m**, of the substance such that:

$$\mathbf{q = mC\Delta T}$$

The **specific heat, C,** can be considered the amount of heat required to raise the temperature of one gram of a pure substance by one degree Celsius. The calorie is a unit of heat as well as the joule. The **specific heat of water** is 1.000 calorie/g ˚C which equals **4.184 J/g ˚C**. The joule (J) is directly related to mechanical work and is the S.I. unit of energy; hence, we shall use the joule almost exclusively.

The specific heat of a metal can be easily measured with a calorimeter. A weighed amount of metal is heated to a specific temperature and is quickly added to a measured amount of water at a known temperature in a calorimeter. The metal loses heat to the water and eventually the metal and water equilibrate at a temperature above the original temperature of the water but below that of the hot metal. If we assume no heat loss to the surroundings or the walls of the container, the heat lost by the hot metal should equal the heat gain of the water. For heat flow *q*,

$$q_{\text{water}} = -q_{\text{metal}} \quad \text{(the negative sign indicates the opposite flow of the heat)}$$

If we now express the heat flow in terms of specific heat (C) for both the metal as well as the water we get:

$$(C_{\text{water}})(m_{\text{water}})(\Delta T_{\text{water}}) = -(C_{\text{metal}})(m_{\text{metal}})(\Delta T_{\text{metal}}) \qquad \textit{equation A}$$

Knowing the specific heat of water we can use this equation to solve for the specific heat of the metal. The specific heat of a metal is related to its molar mass by a simple relationship. Dulong and Petit discovered that 25 joules is required to raise the temperature of one mole of many metals by 1 ˚C. This relationship, shown below, is known as the **Law of Dulong and Petit**:

$$\textbf{Molar Mass (g / mol)} = \textbf{25 / } C_{\text{metal}} \textbf{ (J/g ˚C)} \qquad \textit{equation B}$$

In part A of this lab you will determine the specific heat and molar mass of an unknown metal.

Heat of Reaction and Hess's Law

When a physical or chemical change occurs, it is usually accompanied by a change in the heat content (**enthalpy**) of the material in question. Enthalpy (**H**) is defined as the heat content of a given set of conditions, called a **state**. Since there is only one value of enthalpy for any given state, the enthalpy is one of a number of thermodynamic variables called **state functions**. Because many factors internally contribute to the enthalpy of a substance there is no way to measure the enthalpy of a pure substance. Instead we can determine the change in the enthalpy (**ΔH**) when a chemical or physical change occurs. When a chemical reaction occurs in water solutions, the situation is similar to that which is present when a hot metal sample is put into water. As in the specific heat experiment the heat flow for the reaction mixture is equal in magnitude but opposite in sign to that for the water.

$$q_{\text{reaction}} = -q_{\text{water}} \qquad \textit{and} \qquad \Delta H_{\text{reaction}} = q_{\text{reaction}} / \text{mol}$$

By measuring the mass of the water used as solvent, and by observing the temperature change that the water undergoes, we can find q_{water} and therefore $\Delta H_{\text{reaction}}$. An increase in water temperature indicates that heat is given off by the reaction; the reaction is **exothermic**, and $\Delta H_{\text{reaction}}$ is negative. Conversely, if the temperature decreases, heat is absorbed by the reaction from the surroundings, the reaction is **endothermic**, and $\Delta H_{\text{reaction}}$ is positive.

Hess's law further states that when two or more chemical equations are *combined* to produce a balanced chemical equation, the enthalpy changes combined in the same manner will yield the enthalpy change of the new reaction. This will enable us to determine the enthalpy change for a reaction that may not be easily performed in the laboratory, i.e. the enthalpy of formation of acetylene gas (C_2H_2).

The reaction we are trying to determine is:	$2\ C(s) + H_2(g) \rightarrow C_2H_2(g)$	$\Delta H = ?$
By taking 2 x the heat of formation of CO_2:	$2\ C(s) + 2\ O_2(g) \rightarrow 2\ CO_2(g)$	$\Delta H = -787.0$ kJ
$^1/_2$ x the heat of formation of H_2O:	$H_2(g) + {}^1/_2\ O_2(g) \rightarrow H_2O(l)$	$\Delta H = -285.8$ kJ
Reverse the heat of combustion of C_2H_2:	$2\ CO_2(g) + H_2O(l) \rightarrow C_2H_2(g) + {}^5/_2$	$\Delta H = +846.1$ kJ

The **sum** of these enthalpies is **-226.7 kJ**, which is the enthalpy of formation of acetylene.

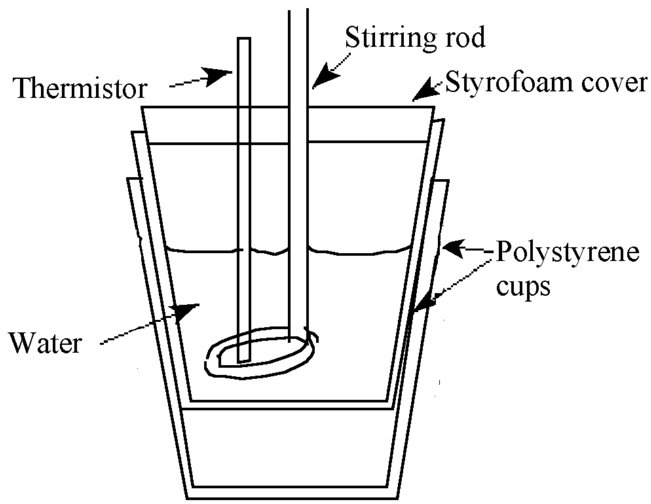
In this experiment we will measure the enthalpy change for the reaction of a metal, zinc, with acid to produce a zinc salt. We will then measure the enthalpy change for zinc oxide reacting with the same acid. From these two reactions along with the value for the reaction of hydrogen with oxygen, one can determine the *heat of combustion of zinc metal* (or the **heat of formation for zinc oxide**):

$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$	ΔH_1 (1)
$ZnO(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$	ΔH_2 (2)
$H_2(g) + {}^1/_2\ O_2(g) \rightarrow H_2O(l)$	$\Delta H_3 = -285.8$ kJ
$Zn(s) + {}^1/_2\ O_2(g) \rightarrow ZnO(s)$	$\Delta H_4 = ?$

As before in the last experiment, you can use a calorimeter to determine the heats (***q***) of reaction and enthalpies for reactions ***1*** and ***2***, above. Combining these enthalpy values with the enthalpy of formation for water (equation ***3***, above), one can use Hess's Law to calculate the heat of formation for Zinc Oxide (equation ***4***.)

PROCEDURE: Part A: SPECIFIC HEAT of an UNKNOWN METAL

1. Assemble the calorimeter as diagrammed. The calorimeter consists of two nested polystyrene coffee cups with a Styrofoam cover. There are two holes in the cover, one for the **thermistor** (which records temperature when connected to a **Vernier LabQuest** apparatus) and one for the glass stirrer provided for this experiment. Weigh the dry calorimeter to 0.001 g. Add about 40 mL of tap water and reweigh the calorimeter and water.



2. Fill a 600 mL beaker 2/3 full with tap water and heat it to boiling. While waiting for the water to boil, weigh a sample of dry metal to the nearest 0.001 g. Return the metal to the dry test tube and clamp the test tube in the boiling water bath so that the metal is below the water line.

3. Record the temperature of the boiling water bath using the LabQuest thermistor probe. Remove the thermistor from the boiling water bath and wipe off all the hot water before placing it in the calorimeter. Record the temperature of the water in the calorimeter.

4. Remove the test tube from the boiling water bath, quickly wipe excess water off the outside of the test tube. Pour the hot metal into the calorimeter without causing the water to splash (tilt the calorimeter). While stirring the water in the calorimeter, monitor the temperature until it remains steady or begins to fall. Record the temperature when it has stabilized.

5. Dry the metal sample, return it to the large test tube and heat it again in the boiling water. **Repeat** the experiment.

Part B: HEAT of REACTION: HESS'S LAW – *all waste should go in a waste container!*

Zinc Reaction

1) Using a graduated cylinder, add about 75.0 mL of 6.00 M HCl in the dry calorimeter. Determine the mass of the HCl solution in the calorimeter to 0.001 g, then record the temperature.
2) Weigh about 0.65 g of Zn to the nearest 0.001 g.
3) Add the metal to the calorimeter, stir and record the highest temperature (when it stabilizes.)

Zinc Oxide Reaction

4) Perform a similar experiment using about 75.0 mL of 6.00 M HCl and 1.2 g zinc oxide.

CALCULATIONS for Part A:

1. For each trial, calculate the specific heat of the metal. Use "equation A" on the front page of this lab.
2. Determine the **average** specific heat and deviation in **parts per thousand**.
3. **Estimate the molar mass** using the law of Dulong and Petit (equation B, front page). What is the identity of your metal?

CALCULATIONS for Part B:

1. Calculate the heat change (q) for the Zn and ZnO reactions.
   Example for q_{Zn}: $\mathbf{q_{Zn}}$ = **-(Heat capacity of HCl)(g HCl *solution*)(ΔT)** *notice the negative sign!*
   ***Heat capacity** for HCl is 3.86 J/g·˚C. Use the mass of the HCl solution for the "g of HCl".

2. Calculate the heat of reaction (ΔH) for the Zn and ZnO reactions. Watch the sign of your value!
   Example for Zn: **ΔH = $\mathbf{q_{Zn}}$ / mol Zn**

3. Write balanced equations for the two reactions performed in lab, including your experimentally determined ΔH. *Hint:* See the second page of this lab, towards the bottom.

4. Use Hess's Law to determine the heat of formation for zinc oxide: $Zn_{(s)} + \frac{1}{2} O_{2\,(g)} \rightarrow ZnO_{(s)}$ (***Hint**:* See the second page of this lab, towards the bottom! You will need the heat of formation for water to calculate the heat of formation for zinc oxide.)

5. Look up the value for the heat of formation of $ZnO_{(s)}$ in your text. Calculate your **percent error**.
   Percent error = absolute value{ (actual - experimental) / actual }*100%. Remember to explain (in your conclusion) any discrepancies.

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Calorimetry Lab

YOUR NAME: ______________________ ***LAB PARTNER(s):*** ________________

DATA: *Record during lab and use in Calculations section. Temperatures should be read to 0.1 °C and masses to 0.001 g for this lab.*

mass of empty calorimeter (g): ______________________ Unknown letter of metal in Part A: ____________

PART A:

Trial 1

calorimeter + water (g): _____________________

water (g): _____________________

metal sample (g): _____________________

initial metal temperature (°C): _______________

initial water temperature (°C): _______________

final temperature (metal + water, °C): _________

Trial 2

calorimeter + water (g): ________________

water (g): ________________

metal sample (g): ________________

initial metal temperature (°C): _______________

initial water temperature (°C): _______________

final temperature (metal + water, °C): _________

PART B:

Zn + HCl

calorimeter + HCl(aq) (g): _____________________

HCl(aq) (g): _____________________

zinc (g): _____________________

initial solution temperature (°C): _______________

final solution temperature (°C): ________________

ZnO + HCl

calorimeter + HCl(aq) (g): ________________

HCl(aq) (g): _____________________

zinc oxide (g): ____________________

initial solution temperature (°C): _______________

final solution temperature (°C): ________________

Notes:

<u>Part A Calculations</u>: *Determining the Specific Heat of an Unknown Metal*

Show all work, use significant figures and circle the final answer for full credit.

- Use the data from Part A to calculate the **specific heat of the metal for both trial #1 and trial #2** (two separate calculations.) *Hint:* Use **equation A** on page I-8-2 to solve for C_{metal}. Use this page and the following page to show ***all*** calculations; attach additional pages of work if necessary. Be sure to include the **unknown letter** of your metal.
- Find the **average specific heat** of your metal and the **parts per thousand** (http://mhchem.org/ppt).
- Use the **Law of Dulong and Petit** (**equation B** on page I-8-2) to estimate the molar mass of the unknown metal and then predict it's identity using the periodic table. ***<u>Discuss briefly</u>*** if you agree with the identity of the metal based on the Law of Dulong and Petit.

Part A Calculations (continued, if needed):

Unknown letter of metal: ___________________________

Trial #1 Specific Heat of Unknown Metal (J g^{-1} K^{-1}): ___________________________

Trial #2 Specific Heat of Unknown Metal (J g^{-1} K^{-1}): ___________________________

Average Specific Heat of Unknown Metal (J g^{-1} K^{-1}): ___________________________

Parts Per Thousand: ____________________________________

Molar Mass of Unknown Metal using the Law of Dulong and Petit (g/mol): ___________

Probable identity of the Unknown Metal (name and symbol): _________________________

Explain briefly if you agree with the probable identity of the metal

Part B Calculations: *Heat of a Reaction / Hess's Law*

For the Zn reaction:

1. Calculate the mass (g) of the HCl solution for the Zn reaction. *Hint:* this will be a number larger than twenty grams!

2. Calculate the change in temperature (ΔT) for the zinc reaction.

3. Using a *heat capacity* of 3.86 J g^{-1} K^{-1} , find q_{Zn} in **Joules (J)** using the equation:

$$q_{Zn} = -(\textit{heat capacity} \text{ of HCl})(\text{g HCl solution})(\Delta T)$$

4. Calculate the moles of Zn (**mol Zn**) used using the grams of Zinc used and the molar mass of Zn.

5. Find the heat of reaction for the Zn reaction (**ΔH_{Zn}**) ***in kJ/mol*** using the equation:

$$\Delta H_{Zn} = q_{Zn} / \text{mol Zn}$$

For the ZnO reaction:

1. Calculate the mass (g) of the HCl solution for the ZnO reaction. *Hint:* this will be a number larger than twenty grams!

2. Calculate the change in temperature (ΔT) for the zinc oxide reaction.

3. Using a *heat capacity* of 3.86 J g^{-1} K^{-1} , find q_{ZnO} in **Joules (J)** using the equation:

$$\boldsymbol{q_{ZnO} = -(\textit{heat capacity} \text{ of HCl})(\text{g HCl solution})(\Delta T)}$$

4. Calculate the molar mass for zinc oxide (g/mol).

5. Calculate the moles of ZnO (**mol ZnO**) used using the grams of zinc oxide used and the molar mass of zinc oxide.

6. Find the heat of reaction for the ZnO reaction (**ΔH_{ZnO}**) ***in kJ/mol*** using the equation:

$$\boldsymbol{\Delta H_{ZnO} = q_{ZnO} / \text{mol ZnO}}$$

Hess's Law:

1. Use your previously calculated values of $\mathbf{\Delta \mathit{H}_{Zn}}$ and $\mathbf{\Delta \mathit{H}_{ZnO}}$ to complete the missing values in the equations below:

$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ ΔH_{Zn} (kJ/mol) = ___________________

$ZnO(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$ ΔH_{ZnO} (kJ/mol) = ___________________

$H_2(g) + {}^1/_2\ O_2(g) \rightarrow H_2O(l)$ ΔH_3 (kJ/mol) = -285.8 kJ

2. Use Hess's Law and the three equations above (ΔH_{Zn}, ΔH_{ZnO}, ΔH_3) to find ΔH_f for the following equation:

$Zn(s) + {}^1/_2\ O_2(g) \rightarrow ZnO(s)$ ΔH_f (kJ/mol) = ___________________

3. Look up the value of ΔH_f for ZnO in your textbook (or here: **https://mhchem.org/thermo**). Calculate the percent error (% error) using the *actual* value (from the textbook or table) and your *experimental* value (answer #2 above.) Why are the two values not equal? Explain briefly.

Percent error = absolute value{ (actual - experimental)/ actual }*100%

Postlab Questions:

Show all work, use significant figures and circle the final answer for full credit.

1. A metal sample weighing 45.23 g is heated to 100.00 ˚C. It is placed in 38.64 g of water in a calorimeter at 25.25 ˚C. At equilibrium, the temperature of the water and metal is 33.05 ˚C. Calculate the specific heat of the metal.

2. When 0.5000 g of Zn are added to 50.00 mL of 6.00 M HCl, the temperature of the solution rises from 20.00 ˚C to 32.56 ˚C. Calculate $\boldsymbol{q_{Zn}}$ in **Joules (J)** and $\Delta \boldsymbol{H_{Zn}}$ in **kilojoules/mol (kJ/mol)** for the reaction. *Some helpful constants:*(C(HCl) = 3.86 J g^{-1} K^{-1}, *d*(HCl) = 1.00 g/mL

<u>Postlab Questions</u>: ***Continued***

3. Calculate the enthalpy change for the allotropic transformation of graphite into diamond using the following data:

   C (graphite) $+ O_2(g) \rightarrow CO_2(g)$ $\Delta H = -390.5$ kJ

   C (diamond) $+ O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393.5$ kJ

4. Using the following equations:

   $C(s) + 2\,Cl_2(g) \rightarrow CCl_4(g)$ $\Delta H^0 = -135.4$ kJ
   $H_2(g) + Cl_2(g) \rightarrow 2\,HCl(g)$ $\Delta H^0 = -184.6$ kJ
   $2\,H_2(g) + C(s) \rightarrow CH_4(g)$ $\Delta H^0 = -74.8$ kJ

   Calculate the standard enthalpy of reaction for the process:

   $CH_4(g) + 4\,Cl_2(g) \rightarrow CCl_4(g) + 4\,HCl(g)$ $\Delta H^0 = ?$

5. Using the Law of Dulong and Petit, calculate the heat capacity of pure gold.

CH 221 Fall 2024:
"Hydrogen Spectrum" *(in class)* ***Lab - Instructions***

Note: **This is the lab for section 01 and H1 of CH 221 only.**

- *If you are taking section W1 of CH 221*, please use this link:
  http://mhchem.org/s/9b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-9-2 through Ia-9-7 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, December 2 (section 01) *or* **Wednesday, December 4 (section H1).** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work and calculations, then **turn it in (**pages Ia-9-5 through Ia-9-7 *only* to avoid a point penalty**) at the end of lab to the instructor** (i.e. the same day as you complete the lab.)

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

The Atomic Spectrum of Hydrogen

When atoms are excited, either in an electric discharge or with heat, they tend to give off light. The light is emitted only at certain wavelengths that are characteristic of the atoms in the sample. These wavelengths constitute what is called the atomic spectrum of the excited element and reveal much of the detailed information we have regarding the electronic structure of atoms.

Atomic spectra are interpreted in terms of quantum theory, which states that atoms can exist only at certain states that correspond to fixed energy levels. When an atom changes its state, it must absorb or emit an amount of energy that is just equal to the difference between the energies of the initial and final states. This energy may be absorbed or emitted in the form of light. The emission spectrum of an atom is obtained when excited atoms fall from a higher to a lower energy level. Since there are many such levels, the atomic spectra of most elements are very complex.

Light is absorbed or emitted by atoms in the form of photons, each with a specific amount of **energy**, **E**. This energy is related to the **frequency (ν)** and **wavelength** (λ) of light by the following equation:

$$E_{photon} = h\nu = \frac{hc}{\lambda}$$

where **h** = **Planck's constant** = 6.62608 x 10^{-34} J·s *and* **c** = **speed of light** = 2.997925 x 10^{8} m / s

The law of conservation of energy states that total energy is conserved. Thus, the change in energy of the atom must equal the energy change of the photon emitted. The energy change of the atom is equal to the energy of the upper energy level minus the energy of the lower level.

$$\Delta E_{atom} = (E_{final} - E_{initial}) = E_{photon} = \frac{hc}{\lambda}$$

The amount of energy in a photon given off when an atom changes from one level to another is very small, of the order of 10^{-19} joules. To avoid working with such small numbers, we will work with one mole of atoms. The above equation is multiplied by **Avogadro's number**, **N**:

$$\Delta E\ (\text{kJ/mol}) = \frac{Nhc}{\lambda} = \frac{(6.02214 \times 10^{23})(6.62608 \times 10^{-34}\,\text{J}\cdot\text{sec})(2.997925 \times 10^{8}\,\text{m/sec})}{\lambda} = \frac{1.19627 \times 10^{5}}{\lambda\ (\text{nm})}$$

The above equation is useful in the interpretation of atomic spectra. For example, in the study of the atomic spectrum of sodium, a strong yellow line of wavelength 589.16 nm is observed. The above equation can be used to determine the change in energy. This in turn is corresponded to the transition in energy levels.

The simplest of all atomic spectra is that of the hydrogen atom. In 1886 Balmer showed that the lines in the spectrum of the hydrogen atom had wavelengths that could be expressed by a rather simple equation. In 1913, Bohr explained the spectrum on a theoretical basis with his model of the hydrogen atom. According to Bohr's theory, the energies allowed to a hydrogen atom are given by the so-called **Bohr's Equation**:

$$E_n = \frac{-B}{n^2} = \frac{-1312.04}{n^2}$$

where **B** = a constant (1312.04 kJ/mol) and **n** = the quantum number (1, 2, 3, ...)
Bohr's equation allows you to calculate quite accurately the energy levels for hydrogen. Transitions between these levels give rise to the wavelengths in the atomic spectrum of hydrogen. These wavelengths are also known very accurately.

Given both the energy levels and the wavelengths, it is possible to determine the actual levels associated with each wavelength. In this experiment, your task will be to measure the wavelengths of the hydrogen spectrum and then determine the transition in energy levels associated with each wavelength. This lab consists of a worksheet that needs to be completed for credit; no formal typed laboratory report is due this week.

PROCEDURE: Part A: Visual Observation of a Hydrogen Discharge Tube Using a Spectroscope

The instructor will set up a spectroscope with a hydrogen discharge tube. Note the color of the emitted light without using the spectroscope.

Now view the hydrogen tube using the spectroscope. How many lines do you see? What color does each line have?

Now we shall use the Vernier system and an emission spectrometer to determine the wavelengths of the hydrogen lines you viewed in the spectroscope.

Part B: The Emission Spectrum of Hydrogen Using the Vernier LabQuest 2

Assemble the emission spectrometer and Lab Quest 2 per your instructor's instructions. In the Lab Quest program, you should see "USB: Intensity rel" if everything is connected correctly.

Place the close to (but not touching!) the middle of a hydrogen discharge tube using a LabJack. Start the data collection by pushing the green "start" button (a green triangle) in the lower left of the Vernier LabQuest 2. **You should see at least three distinct peaks to perform the analysis**: the main central peak (656.3 nm) and at least two to the left of the main 656.3 nm peak (at 486.1 and 434.1 nm.) Ignore any peaks to the right (i.e. larger than 700 nm) – these are due to impurities in the hydrogen tube Once you have at least three peaks visible, **stop** your experiment and turn off the hydrogen discharge tube.

Use the LabQuest 2 tools to find the wavelengths of the four emission peaks for hydrogen. Once the experiment is stopped, go "Graph – Graph Options", then change the following: **Left** to 400, **Right** to 700 and **Top** to 0.400, then press **OK**. Your data points will be visible, and a fourth point should be apparent at about 410 nm.

Determine the four experimental emission wavelengths using the "left" and "right" buttons (which control the cursor; alternatively, use the pointer to select the exact point.) **Record** these values and **compare** them to the theoretical wavelengths for hydrogen to calculate the **percent error** for each line on the next page.

Part C: Calculations for the Energy Levels of Hydrogen Atom

Next, you can use the hydrogen wavelengths to calculate the **energy change** for each line in the observed hydrogen spectrum. Using Bohr's equation, calculate the **energy levels (ε_n)** in kJ/mole for each of the eight lowest allowed levels of the hydrogen atom starting with n=1 to n=8. Note that all the energies are negative, so that the lowest energy will have the largest allowed negative value.

The energy levels will allow you to determine the **energy transition (ΔE)** that corresponds to the observed wavelengths. Determine the **quantum numbers** for the initial (n_{hi}) and final (n_{low}) states for these transitions.

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The Atomic Spectrum of Hydrogen: *Worksheet*

Name:
Lab Partner(s):

Complete the following worksheet using the instructions provided. *Use this form or suffer a point penalty!*

Part A: ***Visual Observations***

What **color** was the hydrogen discharge tube when you looked directly at it? ____________________

What **four colors** did you observe through the spectroscope? ________________________________

Part B: ***The Emission Spectrum of Hydrogen*** – Use the Vernier system to find the hydrogen wavelengths.

Color	λ (Vernier, nm)	λ (theoretical, nm)	Percent Error
Red	____________	656.3	___________
Blue	____________	486.1	___________
Violet	____________	434.1	___________
Violet	____________	410.2	___________

Recall:

$$\text{Percent Error} = \frac{\text{absolute value (theoretical value - experimental value)}}{\text{theoretical value}} * 100\%$$

Part C: Calculations for the Energy Levels of Hydrogen Atom

Find the energy level ε_n (in kJ/mol) for each quantum number from 1 through 8 using the following equation:

$$\varepsilon_n = \frac{-1312.04\ \text{(kJ/mol)}}{n^2}$$

Example: The n=1 energy level can be calculated as follows: $\varepsilon_n = (-1312.04/1^2)$ = **-1312.04 kJ/mol**
Example: The n=2 energy level can be calculated as follows: $\varepsilon_n = (-1312.04/2^2)$ = **-328.010 kJ/mol**

n value	ε_n (kJ/mol)	n value	ε_n (kJ/mol)
1	-1312.04	5	______
2	-328.010	6	______
3	______	7	______
4	______	8	______

Now complete the table below. The first row has been partially completed for you.

Color	λ (actual, nm)	ν (s^{-1})	ΔE (J/photon)	ΔE (kJ/mol)	n_{hi}	n_{low}
red	**656.3**	**4.568×10^{14}**	**3.027×10^{-19}**	**182.3**	_____	_____
blue	**486.1**	________	________	______________	_____	_____
violet	**434.1**	________	________	______________	_____	_____
violet	**410.2**	________	________	______________	_____	_____

Convert the wavelength values into ΔE (in kJ/mol) using the following equation:

$$\Delta E \text{ (kJ/mole)} = \frac{1.19627 \times 10^{5}}{\lambda \text{ (nm)}}$$

The calculated ΔE values correspond to a transition between the various energy levels, ε_n, calculated previously. Determine which transition they correspond to by finding the change in energy (i.e. ΔE) between levels.

Example: Find the change in energy in a transition of hydrogen between the n=2 and n=1 energy levels.

The energy level, ε_n, for n=2 is **–328.010** kJ/mol, and the energy level, ε_n, for n=1 is **–1312.04** kJ/mol. A change in energy, **ΔE**, corresponds to the final energy state minus the initial energy state, or:

$\Delta E = \varepsilon_{final} - \varepsilon_{initial} = \varepsilon_1 - \varepsilon_2 = -1312.04 - (-328.010) =$ **-984.03 kJ/mol**

If your calculated value of ΔE is about –984.03 kJ/mol, then your n_{hi} would be **2** (the higher value of n) and your n_{low} value would be **1** (the lower value of n).

Show the frequency calculation for the red line below:

Show the ΔE (J/photon) calculation for the red line below:

Show the ΔE (kJ/mol) calculation for the red line below:

Post Lab Questions:

1. When Balmer found his famous series for hydrogen in 1886, he was limited experimentally to wavelengths in the visible and near ultraviolet regions from 250 nm to 700 nm, as in your experiment. What common characteristic do the lines in the Balmer series have?

2. In the hydrogen atom, the electron is in its lowest energy state, n=1. The maximum electron energy that a hydrogen atom can have is 0 kJ/mole, at which point the electron would essentially be removed from the atom and it would become a H^+ ion. How much energy does it take to ionize one hydrogen atom in **kilojoules per mole** *and* in **Joules per atom**? (*Hint:* calculate ΔE where ε_{final} is zero and $\varepsilon_{initial}$ is –1312.04 kJ/mol.)

Questions #3 through #5 will use the equation below for the helium ion. The **helium ion**, He^+, has energy levels similar to those of the hydrogen atom, since both species have only one electron. The energy levels of the helium ion are given by the following equation:

$$E_n = \frac{-5248.16}{n^2} \text{ kJ/mol where } n = 1, 2, 3...$$

3. Calculate the energies in kJ/mole for the four lowest energy levels of the helium ion using the equation above.

ε_1 __________ ε_3 __________

ε_2 __________ ε_4 __________

4. One of the most important transitions for the helium ion involves a jump from the n = 2 to the n = 1 level. Calculate the **change in energy** in kJ/mole for this transition. (*Hint:* $\Delta E = \varepsilon_1 - \varepsilon_2$). Use the equation found in Part C of the worksheet to calculate the **wavelength** (in nm) of this transition.

5. Three of the strongest lines in the helium ion spectrum are observed at the following wavelengths. Find the quantum numbers of the initial and final energy states for the transitions that give rise to these three lines:

λ	ΔE (kJ/mol)	n_{hi}	n_{low}
121.57 nm			
164.12 nm			
468.90 nm			

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"Introduce Yourself" Lab for Chemistry 221 Section W1

***Create a video, sign the form below** and turn in via email to mike.russell@mhcc.edu by **9 AM, Friday, September 27**.*
Remember to turn in the video link to me as well!

Note: **This is the lab for section W1 of CH 221 only.**

- *If you are taking section 01 or section H1 of CH 221*, please use this link: http://mhchem.org/s/1a.htm

Welcome to Chemistry 221! I am glad you enrolled in CH 221 this quarter, and I look forward to an exciting term with you!

This class will be quite different from previous Chemistry 221 classes taught at Mt. Hood Community College... it will be taught exclusively online; there will be no "face to face" lectures, labs, office hours, exams, etc. So…. let's make the best of it, ok? :)

The goal of this "lab" is to "meet you." I want to know a bit about you, so **I want you to make a short (about 3 minutes or so) video (preferably on YouTube or a similar platform) about yourself**. ***Show yourself talking*** (no pictures, etc. - just show you!) and tell me a bit about yourself. Maybe you could tell me about your college goals - why are you taking CH 221? Or maybe you could tell me about a cool movie you watched, or a book you read, or a music band you're enjoying…. it's totally up to you, but it will help me get a better idea as to "who you are", and this is important to me! **Email me a link to the video**.

I also want you to read the "Memorandum" page (which is found below), then **initial, sign and return the "Memorandum" page to me electronically** (i.e. email to mike.russell@mhcc.edu).

I have suggestions for completing both assignments on the next several pages.

How to Create a Video for this Assignment:

Making a video should not be a difficult assignment for you. I do not expect a "Hollywood quality" video; instead, I just want to see YOU and hear some of your stories. *This **MUST** be an original and current video - do not re-use a video made from a previous class.*

To create the video, I recommend creating a video on your phone, then using the YouTube app to upload the video. Connect your gmail/Google account in YouTube, select the "plus" symbol (which is at the bottom middle of the screen) to start uploading the video (and make sure you set the video to "**unlisted**", *not* "private".) Once ready, email me a link (use the "Share" - "Copy Link" function)... and then you're done!

You do not have to edit your video - it can be pretty rough! And if you don't like your video, record a new version and start again.

How to Fill Ou the "Memorandum" for this Assignment:

The final page of this assignment has the "Memorandum" which I also want you to submit to me via email. **All assignments must be submitted to the instructor via email** (mike.russell@mhcc.edu) as a **PDF file**, **written by hand** (no typed assignments except for the Class Presentation materials) and **only in one file** (i.e. if the assignment is five pages, submit all five pages as one file and not five individual files.)

How you do this depends on you… here are some suggestions:

1 - **If you have a printer and wish to complete the work on regular paper.**

• print the assignment and fill out as usual. You cannot print at MHCC currently (hopefully this will change soon) so this must be done on your own.
• On your **phone** (Android or iPhone), use a free program like **CamScammer** to make pdf scans and combine into one file. Alternatively you can use **CombinePDF** (https://combinepdf.com) to automatically convert multiple picture (.jpg or .png) files into a single PDF file. These services should be free - do not pay for any upgrades or extras!
• Email the PDF to the instructor! done! (and again, I *really* like CamScanner!)

2 - If you have a tablet (iPad, Surface, etc.) and can write directly on the screen:

• Download the PDF file (to the desktop, Google Drive, etc.)
• Use a program which allows you to import the PDF and write directly on the tablet. Examples include (but are not limited to): GoodNotes (my current favorite), Notability, Apple Notes, Evernote, Google Keep, Typora or Microsoft OneNote Some of these programs might have a cost associated with them.
• Email the completed PDF assignment to the instructor... you're done!

Note that as a MHCC student, you can **access Microsoft Office for free** *on both Windows and Mac platforms. More info:* **https://mhcc.edu/OfficeInstall/**

You **pick which of these methods works well for you, and use it complete all assignments in CH 221 this quarter.**

And if you have questions on anything, please email me (**mike.russell@mhcc.edu**) - I'm happy to help!

Good luck, and I look forward to having you in my classroom this quarter! Peace, Michael

p.s. Want to know more about me? https://mhchem.org/221/russellm/index.htm

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Memorandum for Chemistry 221 Section W1

Create a video, sign the form below *and turn in everything via email to mike.russell@mhcc.edu by* ***9 AM, Friday, September 27****. Also remember to turn in the* ***video link!***

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* Please note: **if you plan on taking Organic Chemistry in the future, you really should sign up for Section 01 or Section H1 of CH 221**. You will be missing out on essential lab skills and procedures which will be missed if you take Organic Chemistry. Email me if you have questions, I might be able to switch you to Section 01 or Section H1 this term.

* **I agree to turn in all assignments via email as PDF files**. All assignments will be **submitted as a single file** (do not submit one assignment as multiple files) **and hand written** (ie do not type the assignments.)

* **I understand that problem sets, labs, quizzes and most assignments are due on Wednesdays by 11:59 PM *or* Fridays by 9 AM via email** (check the syllabus for exact due dates.) **Late assignments (even due to technology reasons) will incur a point penalty. Quizzes and exams will be worth zero points if turned in late.** Email assignments early if you worry about the quality of your internet connection.

* **I understand that all assignments will be returned via email to your MHCC @saints account**. This address will be used to discus items related to our class during the term.

* **If you cannot complete the midterm or final exam** *(i.e. vacation, etc.)*, **you will need to complete a make up exam in person on the main MHCC campus in Gresham**.... so for the duration of the term, try to focus on the class and not be away from your computer and phone, ok?

* **I will try to have a sense of humor** as the instructor frantically tries to keep up with the changes of this class :). **I will email the instructor** if I have any questions!

* **I have read this document and will stay informed with the class through the instructor's email messages and class syllabus.**

____

Signature

____

Printed name

____

Date

CH 221 Fall 2024:
"Density" *(online)* ***Lab*** *Instructions*

Note: **This is the lab for section W1 of CH 221 only.**

- *If you are taking section 01 or section H1 of CH 221*, please use this link: http://mhchem.org/s/2a.htm

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Step One:

Watch the lab video for the "Density" lab, found here: **http://mhchem.org/w/2.htm**
Record the data found at the ***end*** of the lab video on page Ib-2-5.

Step Two:

Complete pages Ib-2-5 through Ib-2-9 using the "Density" video and the actual lab instructions on pages Ib-2-2 through Ib-2-4. Include your name on page Ib-2-5!

Step Three:

Submit your lab (pages Ib-2-5 through Ib-2-9 *only* to avoid a point penalty**) as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, October 9 by 11:59 PM.** I recommend a free program (ex: CamScanner, https://camscanner.com) or a website (ex: CombinePDF, https://combinepdf.com) to convert your work to a PDF file.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Determining the Density of Liquids & Solids

Density, like color, odor, melting point, and boiling point, is a physical property of matter. Therefore, density may be used in identifying matter. Every substance (element, compound, alloy, etc.) has a distinct density. **Density** is defined as **mass per unit volume** and is expressed mathematically as **d = m/V** (d is density, m is mass, and V is volume). Density is essentially a measurement of how tightly matter is packed together.

Density is an important concept in a wide range of fields including chemistry, physics, material science, engineering, geology, meteorology, biology and medicine. For example, a bone density test uses X-rays to determine how much calcium and minerals are packed into a segment of bone. You may also be familiar with different types of plastics, including high-density polyethylene (HDPE, #2, used for milk jugs, hula hoops, and breast implants) and low-density polyethylene (LDPE, #4 used mostly for plastic bags.) The difference between these plastics depends on how tightly the polyethylene molecules are packed together during synthesis.

The **density of water** is 1.00000 g/cm^3 at 4 °C and is slightly less at room temperature. In lab today, you will be using the *Handbook of Chemistry and Physics* to determine the exact density of water at a specific temperature. The density of various materials ranges significantly from less than water (styrofoam's density is about 0.1 g/cm^3) to much greater (Osmium has a density of 22.6 g/cm^3.) For example, aluminum has a density of 2.70 g/cm^3 whereas a sample of lead has a density of 11.2 g/cm^3. The same volume of lead will have a mass over four times that of aluminum! That is why lead is used to shield against X-rays whereas aluminum would be ineffective. Aluminum atoms are not only smaller but also packed so that there is more space between atoms.

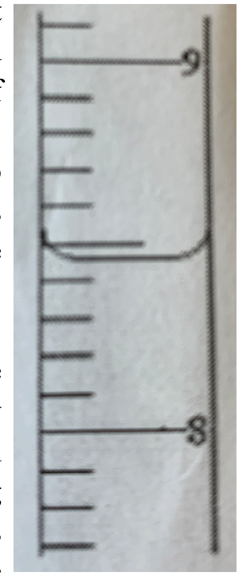
The SI unit (International System of Units) for density is kg/m^3 and is typically used by physicists and engineers. Because chemists work with much smaller masses and volumes, traditional metric units of **g/cm^3** or **g/mL** are the preferred units of measurement (note that **1 mL = 1 cm^3**). Liquids are usually measured in g/mL while solids are measured in g/cm^3. Gases are much less dense, so their density is measured in **g/L**.

Density is often represented as a relative density or ***specific gravity***, a dimensionless quantity that expresses density as a multiple of a given standard (such as water or a gas.) For example, gasoline has a density of 0.67 g/cm^3. Its specific gravity, relative to water, is 0.67. Specific gravity is used in many fields from chemical engineers studying concrete to food scientists testing the alcohol content of a microbrew.

To determine density, mass and volume must both be determined. The **mass** can easily be found using a balance. Determining the **volume** is more ambiguous. The volume of a liquid can be determined using a calibrated container such as a graduated cylinder or graduated flask. The volume of a solid sample with a regular geometric shape can be determined by direct measurement. However, most solids have an irregular shape. The volume can be determined by immersing the solid in a known volume of liquid and measuring the volume of liquid displaced.

This is similar to the method utilized in the ancient tale of **Archimedes** to prove that King Hiero II's crown was not real gold. Archimedes is alleged to have come upon the liquid displacement method while bathing and noticing the rise in his bath water. He then ran through the streets shouting "Eureka!" (I found it!), so excited that he forgot his bathrobe. After dressing, he then proved that the king's crown did not displace the same amount of water as a piece of gold of the same mass. This method is called the "**displacement method**" and can be used with a variety of liquids in order to find the density of various materials.

All measurements are approximations. **Significant figures** ("**sig figs**") are those digits that carry meaning which contributes to precision. The uncertainty is in the last digit and determined by the device. For example, when reading a graduated cylinder, the number of sig figs is estimated one digit beyond the gradations. For the *example pictured on the right,* the bottom of the meniscus is between the 8.4 mL and 8.5 mL markings. You can estimate to the hundredth place or 8.45 mL. Reading a meniscus is subjective and takes practice. In this experiment, you will use the mass and density of water to find the volume of a flask more precisely and reduce human bias.



Mass versus Weight: When determining density, you must determine the mass of the sample. The terms *mass* and *weight* are easily confused. The ***mass*** of a substance is how much matter it is composed of. Units of mass are grams and kilograms. The mass of an object is the same on earth or on the moon. ***Weight*** is a measure of the force of gravity acting on the object. Pounds (units = lb) is a unit of weight, a force. The weight of an object is variable depending on the location of the object. If Joe weighs 220 Ibs at the North Pole, he would weigh only 219 Ibs at the equator due to the bulge of the earth. He weighs only 37 Ibs on the moon. In outer space, an astronaut is weightless but never massless. A great blue whale is weightless in space, but it would still cause damage to your spaceship if you bumped into it.

Accuracy and Precision: Accuracy is how close a measurement, or average of measurements, come to the actual or accepted value. Accuracy is often compared to hitting the bull's eye on a target. In a chemistry lab, accuracy is how close the final calculated answer is to the accepted book value. When working with an unknown, students are graded on their accuracy, how close their answer is to the actual value. Accuracy is determined by calculating the percent error: **percent error = [(|actual value - experimental value|)/actual value] * 100%** (notice the absolute value in the numerator; percent error should be a positive number.) We will use percent error in an upcoming lab. A "good" percent error varies depending on the experiment, the equipment used, and the technician's experience.

Precision is how close multiple measurements of the same quantity come to each other. Precision is a measure of *consistency in lab technique*; is the data reproducible? One method to determine precision is to calculate **parts per thousand** (see the handout in the "Lab Notes" of your Companion, or ask the instructor.) We will calculate parts per thousand in future labs. The term **precision** also refers to the number of significant digits in a measurement. For example, the balances (scales) that will be used in this lab allow measurements to 1 mg (0.001 g). An analytical balance allows mass.to be determined to 0.1 mg (0.0001 g) and so is more precise. The technique used in this lab for determining volume allows for more significant figures and, hence, is more precise than simply using a graduated cylinder.

Random Error and Systematic Error: Random errors originate from uncontrollable variables in an experiment. Momentary fluctuations in air currents can affect balance reading. A student who rushes through the lab and follows directions haphazardly will perform many random errors. Random errors affect the precision of measurements and the overall experiment. **Systematic errors** are controllable or repeated errors in an experiment. A poorly calibrated scale will result in all mass measurements being erroneous by the same factor. A student consistently misreading an instrument is a systematic error. Because a systematic error is consistent throughout the experiment, it does not affect the precision but can significantly affect the accuracy.

In this experiment you will determine densities of an unknown liquid and solid by measuring their mass with a balance and their volume. First, you will determine the exact volume of a flask using water. You will determine the density of a solid by displacement of a known quantity of water.

***PROCEDURE*:** (this is what we would have done in the lab room under "normal" circumstances)

Part A: Density of a Liquid *All mass measurements should be recorded to the milligram (0.001 g.)*

1. Clean a 10 mL volumetric flask with soap and water. Dry with a small amount of acetone in the hood and by gently blowing compressed air into it. Determine and record the mass (to the nearest mg) of the *clean* and *dry* 10 mL volumetric flask with a stopper in it.
2. Fill this 10 mL volumetric flask with deionized water. Insert stopper so no air remains in flask. Dry the outside of the volumetric flask. Record the mass to the nearest 0.001 grams.
3. Calculate the mass of the water in the flask. Remember to show all calculation steps in your lab report.
4. Determine the temperature of the water to the tenths place. Use the *Handbook of Chemistry and Physics* to find the density of water at this temperature. If the *Handbook* is difficult to read, use this link as an alternative: http://mhchem.org/den
5. Calculate the volume of this volumetric flask. Remember significant digits!
6. Dry the volumetric flask. Obtain an unknown liquid and record the identification number. Fill the volumetric flask with the unknown liquid, stopper and record the mass.
7. Calculate the mass of the unknown liquid added. Calculate the density of the unknown liquid in g/mL to the correct number of significant digits.

Part B: Density of a Solid *All mass measurements should be recorded to the milligram (0.001 g.)*

1. Select an unknown metal and record its identifier.
2. Clean and dry a 50 mL Erlenmeyer flask that will fit your metal sample. Record the mass of the dry flask and stopper. Fill the flask with water. Record the mass.
3. Determine the volume of the Erlenmeyer flask as in part A.
4. Empty and dry the flask thoroughly. Add small chunks of a dry metal sample to the flask until the flask is at least half full. Weigh the flask, with its stopper and the metal, to the nearest milligram. You should have about 50 g of metal in the flask (more is better!)
5. Determine the mass of metal added.
6. Leaving the metal in the flask, fill the flask with water and replace the stopper. Roll the metal around in the flask to make sure that no air is trapped between the metal pieces. Refill the flask if necessary, and then weigh the stoppered flask full of water plus the metal sample.
7. Calculate the mass of water added.
8. Calculate the volume of water added based on its density and mass.
9. Calculate the volume of metal added. Use this value to calculate the density (in g/cm^3) of the metal.
10. Pour the water from the flask. Dry the metal before returning to its container.

Density

Watch this lab video first: **http://mhchem.org/w/2.htm**

YOUR NAME: ____

DATA: *Watch the video (***http://mhchem.org/w/2.htm***) to acquire these values using the data at the very end:*

Part A

empty flask (g): ____________________

flask + water (g): ____________________

water temperature (°C): ____________________

density of water (g/mL): ____________________

flask + unknown (g): ____________________

Part B

empty flask (g): ____________________

flask + water (filled) (g): ____________________

water temperature (°C): ____________________

density of water (g/mL): ____________________

flask + metal (g): ____________________

flask + metal + water (g): ____________________

<u>Part A Calculations</u>: ***Determining the Density of an Unknown Liquid***

Show all work, use significant figures and circle the final answer for full credit.

1. Using the data from the video, calculate the mass (g) of water in the flask for Part A.

2. Using the mass of water in the flask (above) and the given density from the video, calculate the volume (mL) of the water in the flask.

3. Assuming that the water *completely* filled the flask (and it did!), determine the volume (mL) of the flask.

4. Using the data from the video, calculate the mass (g) of the unknown in the flask.

5. Using the mass of the unknown in the flask (g), and assuming the unknown completely filled the flask (it did!), determine the density (g/mL) of the unknown liquid.

Part B Calculations: ***Determining the Density of an Unknown Solid***

Show all work, use significant figures and circle the final answer for full credit.

1. Using the data from the video, calculate the mass (g) of water in the flask for Part B.

2. Using the mass of water in the flask (above) and the given density from the video, calculate the volume (mL) of the water in the flask.

3. Assuming that the water *completely* filled the flask (and it did!), determine the volume (mL) of the flask.

4. Using the data from the video, calculate the mass (g) of the unknown metal in the flask.

5. Using the data from the video, calculate the mass (g) of water in the flask when the metal was present. *This is a different value from step 1 of part B!*

6. Convert the grams of water in the flask when the metal was present (step 5, above) into the volume of water (mL) present. Use the density from step 2, above.

Part B Calculations: ***Continued***

7. Find the volume of the metal (cm^3) using the volume of the flask (step 3) and the volume of water present with the metal (step 6.)

8. Using the mass of the metal in the flask (g, step 4) and the volume of the metal (step 7), determine the density (g/cm^3) of the unknown metal.

Postlab Questions:

Show all work, use significant figures and circle the final answer for full credit.

1. In the original Indiana Jones movie, our hero is attempting to claim a precious ancient gold relic from a poor third world country. He estimates the size of his prize and carefully adjusts the *volume* of sand in his bag to equal that of the gold relic. With the dexterity that only Indiana Jones possesses, he swiftly but delicately swaps the sand for the gold. After a moment of delight, our hero realizes he has misjudged and the ancient tomb is not fooled. What went wrong? *You do not have to watch the Indiana Jones movie to answer this question!* ☺

2. Using the techniques covered in this lab, how can the volume of an irregularly shaped object that is less dense than water be found? Assume the object's density is unknown, and "forced submersion" or "weighted submersion" answers will not get credit.

Postlab Questions: ***Continued***

3. While panning for gold, you find a nugget that looks like gold. You find its mass to be 1.25g. You know that the density of pure gold is about 20.0 g/cm^3 and that the density of iron pyrite (fool's gold) is 5.0 g/cm^3. Determine if a cubic nugget about 0.40 cm on each side is fool's gold or pure gold. (Show all work)

4. Dennis obtained a clean, dry stoppered flask. He determined the mass of the flask and stopper to be 32.634 g. He then filled the flask with water and determined the mass of the full stoppered flask to be 59.479 g. Based on the temperature of the water, Dennis found the density of water in the *Handbook of Chemistry and Physics* to be 0.998730 g/cm^3. Calculate the volume of the flask.

5. Dennis emptied the flask from question #4, dried it and filled it with an unknown liquid. The mass of the stoppered flask when completely filled with liquid was 50.376 g. Calculate the density of the unknown liquid.

6. Dennis emptied the flask from question #4 and #5 and dried it again. He added an unknown metal to the flask. He determined the mass of the stoppered flask and metal to be 152.047 g. He then filled the flask with water, stoppered it and obtained a total mass of 165.541 g. Calculate the volume of metal added and the density of the unknown metal.

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CH 221 Fall 2024:

“Chemical Nomenclature”

(online) ***Lab*** *- Instructions*

Note: **This is the lab for section W1 of CH 221 only.**

- *If you are taking section 01 or section H1 of CH 221*, please use this link:
  http://mhchem.org/s/3a.htm

Step One:

Watch the lab video for the “Nomenclature” lab, found here:
http://mhchem.org/w/3.htm
There is no data to record from the video in this lab.

Step Two:

Complete pages Ib-3-11 through Ib-3-15 using the “Nomenclature” video and the actual lab instructions on pages Ib-3-2 through Ib-3-10. Include your name on page Ib-3-11!

Step Three:

Submit your lab (pages Ib-3-11 through Ib-3-15 *only* to avoid a point penalty**) as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, October 16 by 11:59 PM.** I recommend a free program like CamScanner (https://camscanner.com) to convert your work to a PDF file.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Chemical Nomenclature

Chemical nomenclature is the system that chemists use to identify and name compounds. Compounds can have two types of names: *systematic names* (names that identify the chemical composition of a chemical compound) and *common names* (traditional names based on historical discovery or reactivity behavior). For example, N_2O has both a systematic name (dinitrogen monoxide) and a common name (laughing gas).

If every substance were assigned a common name, chemists would be expected to memorize over nine million names! This is why chemists generally prefer systematic names for identifying compounds. The International Union of Pure and Applied Chemistry (IUPAC, see http://www.iupac.com) was founded in 1921 to provide a system of chemical nomenclature for scientists. IUPAC nomenclature rules can provide valuable structural and reactivity information. On the other hand, most people would be hard pressed to call dihydrogen monoxide by any other name but water, so both types of nomenclature have their place.

Nomenclature leads naturally to formula writing. Compounds exist in distinct combinations of elements, and knowing the proper combinations of elements is essential in chemistry. We expect sodium chloride to be NaCl and not Na_2Cl or $NaCl_2$; knowing which combination or combinations exist in nature is crucial.

The following sections will guide you through the rules of *inorganic nomenclature* and formula writing. Later you may experience the nomenclature for organic chemistry or transition metal chemistry, but most of the compounds observed in first year chemistry will fall in this category.

Part A: Nomenclature of Elemental Ions

The first step in learning nomenclature is to learn the names of the elemental ions you might see in compounds. We make a distinction between the following:

- **fixed charge cations** (metal positive ions from groups IA, IIA and Al, Ga, In, Zn, Cd, Ag ("the stairs")
- **variable charge cations** (positive ions which do not have a fixed charge; this includes ***all*** transition metals not in "the stairs", lanthanides, actinides, Tl, Pb, Sn, and Bi)
- **anions** (negative ions which are generally nonmetals) have a charge equal to the **group number - 8**

Why two types of cations? Many metals have, for all practical purposes, only one ionic charge observed in nature. Lithium is only observed as Li^+ naturally, and even though gas phase studies of lithium ions have produced Li^{2+} and even Li^- ions, they are not observed in most settings. Many metals (such as iron) have many different *oxidation states* or *ionic charges* associated with them. The ions Fe^{2+}, Fe^{3+} and Fe^{6+} can be observed and manipulated quite readily (even at Mt. Hood Community College!); therefore, we need a method to distinguish between the various ions (namely iron(II), iron(III) and iron(VI), respectively).

Fixed Charge Cations use their elemental name.

Example:	Na^+ is the sodium ion	Cs^+ is the cesium ion
	Mg^{2+} is the magnesium ion	Sr^{2+} is the strontium ion
	Al^{3+} is the aluminum ion	In^{3+} is the indium ion

Variable Charge Cations use their elemental name followed by their ionic charge in parentheses. Use *Roman numerals* to distinguish the charge of the ion.

Example:	Fe^{2+} is the iron(II) ion	Pb^{2+} is the lead(II) ion
	Fe^{3+} is the iron(III) ion	Pb^{4+} is the lead(IV) ion
	Mn^{7+} is the manganese(VII) ion	Co^{9+} is the cobalt(IX) ion
	U^{4+} is the uranium(IV) ion	Ti^{2+} is the titanium(II) ion

Anions use their elemental name with the ending changed to *-ide*. *Notice:* charge = group number - 8

Example:	Cl^- is the chloride ion	I^- is the iodide ion
	O^{2-} is the oxide ion	Te^{2-} is the telluride ion
	N^{3-} is the nitride ion	As^{3-} is the arsenide ion

Part B: Nomenclature of Polyatomic Ions

Certain combinations of atoms result in stable configurations that are not easily destroyed; these are called *polyatomic ions*. Polyatomic ions can be either positive or negative, but most of them are anions (i.e. they have a negative charge.) Recognizing polyatomic ions in formulas is one of the most difficult concepts to master when learning nomenclature, and it is *very important that you* ***memorize*** *the following list of polyatomic ions*.

A list of polyatomic ions is given below:

nitrate	NO_3^-	**hydroxide**	OH^-	**hypochlorite**	ClO^-
nitrite	NO_2^-	**cyanide**	CN^-	**chlorite**	ClO_2^-
sulfate	SO_4^{2-}	**thiocyanide**	SCN^-	**chlorate**	ClO_3^-
sulfite	SO_3^{2-}	**cyanate**	OCN^-	**perchlorate**	ClO_4^-
phosphate	PO_4^{3-}	**thiosulfate**	$S_2O_3^{2-}$	**hypobromite**	BrO^-
phosphite	PO_3^{3-}	**chromate**	CrO_4^{2-}	**bromite**	BrO_2^-
hydrogen phosphate	HPO_4^{2-}	**dichromate**	$Cr_2O_7^{2-}$	**bromate**	BrO_3^-
dihydrogen phosphate	$H_2PO_4^-$	**permanganate**	MnO_4^-	**perbromate**	BrO_4^-
carbonate	CO_3^{2-}	**acetate**	$C_2H_3O_2^-$	**hypoiodite**	IO^-
hydrogen carbonate	HCO_3^-	**ammonium**	NH_4^+	**iodite**	IO_2^-
hydrogen sulfide	HS^-	**hydrogen**	H^+	**iodate**	IO_3^-
oxalate	$C_2O_4^{2-}$	**hydride**	H^-	**periodate**	IO_4^-

Part C: Nomenclature of Ionic Compounds from Ions

Knowing the nomenclature rules for ions, we can begin the naming of ionic compounds. Ionic compounds involve a *cation* (either *fixed* or *variable charge*) combining with an *anion*. Naming ionic compounds is straightforward; simply combine the ionic names with the cation first followed by the anion.

Example: **sodium** ion + **chloride** ion give **sodium chloride**
iron(III) ion + **bromide** ion gives **iron(III) bromide**
ammonium polyatomic ion + **oxide** ion gives **ammonium oxide**
aluminum ion + **sulfate** polyatomic ion gives **aluminum sulfate**

Part D: Writing Formulas for Ionic Compounds Using Nomenclature

Another important concept to master is the ability to write a chemical formula using the compound's systematic name. This can be accomplished using the following protocol:

1. Identify the elemental ions and/or polyatomic ions in the compound using the systematic name.
2. Determine the magnitude of the ionic charge on each ion
3. Assume the compound is electrically neutral *unless* the term "ion" appears in the name
4. The sum of the cation charges plus the anion charges must equal zero; combine the ions until this condition is met
5. Write the resulting formula. If more than one polyatomic ion is present, write the polyatomic portion in parentheses with a subscript after it denoting the number of polyatomic ions present.

Example: Write the formula for **sodium chloride**.

1. Sodium chloride has Na^+ and Cl^- ions
2. Sodium has a +1 charge, chloride has a -1 charge
3. Assume sodium chloride is neutral (no "ion" is present in the name)
4. Charge on sodium + charge on chloride = (+1) + (-1) = 0; therefore, **one** sodium ion and **one** chloride ion was required for a neutral compound.
5. 1 Na^+ ion and 1 Cl^- ion gives the formula **NaCl**

Example: Write the formula for **aluminum sulfide**.

1. Aluminum sulfide has Al^{3+} and S^{2-} ions
2. Aluminum has a +3 charge, sulfide has a -2 charge
3. Assume aluminum sulfide is neutral (no "ion" is present in the name)
4. Charge on aluminum + charge on sulfide = (+3) + (-2) = +1; this would indicate that combining one aluminum ion with one sulfide ion would give an *ion*. We want a *neutral* compound (see step 3); this can be accomplished by multiplying aluminum by **2** and sulfide by **3**, which results in: 2(+3) + 3(-2) = 0. Therefore, a neutral compound would result by combing **two** aluminum ions with **three** sulfide ions.
5. 2 Al^{3+} ions and 3 S^{2-} ions give the formula **Al_2S_3.**

Example: Write the formula for **magnesium nitrate**.

1. Magnesium nitrate has Mg^{2+} and NO_3^- ions
2. Magnesium has a +2 charge, nitrate has a -1 charge
3. Assume magnesium nitrate is neutral (no "ion" is present in the name)
4. Charge on magnesium + charge on nitrate = (+2) + (-1) = +1; this would indicate that combining one magnesium ion with one nitrate ion would give an *ion*. We want a *neutral* compound (see step 3); this can be accomplished by multiplying magnesium by **1** and nitrate by **2**, which results in: 1(+2) + 2(-1) = 0. Therefore, a neutral compound would result by combing **one** magnesium ion with **two** nitrate ions.
5. 1 Mg^{2+} ions and 2 NO_3^- ions give the formula **$Mg(NO_3)_2$.** (Note there are *two* nitrate ions, so they are placed in parentheses with a subscript two after it.)

Example: Write the formula for **titanium(IV) oxalate**.

1. Titanium(IV) oxalate has Ti^{4+} and $C_2O_4^{2-}$ ions
2. Titanium(IV) has a +4 charge, oxalate has a -2 charge
3. Assume titanium(IV) oxalate is neutral (no "ion" is present in the name)
4. Charge on titanium(IV) + charge on oxalate = (+4) + (-2) = +2; this would indicate that combining one titanium(IV) ion with one oxalate ion would give an *ion*. We want a *neutral* compound (see step 3); this can be accomplished by multiplying titanium(IV) by **1** and oxalate by **2**, which results in: 1(+4) + 2(-2) = 0. Therefore, a neutral compound would result by combing **one** titanium(IV) ion with **two** oxalate ions.
5. 1 Ti^{4+} ions and 2 $C_2O_4^{2-}$ ions give the formula $\mathbf{Ti(C_2O_4)_2}$**.** (Note there are *two* oxalate ions, so they are placed in parentheses with a subscript two after it.)

Part E: Finding Systematic Names for Ionic Compounds Using Formulas

Determining the systematic name of a compound from its formula is straightforward using these steps:

1. Identify the cation and anion in the formula. *Watch* for polyatomic ions.
2. Assume the compound is electrically neutral *unless* a charge appears in the formula
3. Determine the name of the anion and the charge on the anion
4. If a fixed charge cation is present, determine its name.
5. If a variable charge cation is present, determine its name and use this formula to find the charge on the metal: $\mathbf{charge_{metal} = -\ (\#\ anions)(charge_{anion})\ /\ (\#\ metal\ cations)}$ (where # = "number of")
6. Combine the cation and anion names as per Part C. The cation goes first, followed by the anion; do not forget the Roman numeral charge in parentheses for variable charge cations.

Example: Determine the name for **NaCl**.

1. The cation is Na and the anion is Cl
2. NaCl is neutral (no charges are present in the formula)
3. The anion, the chloride ion, has a -1 charge
4. Na is a fixed charge cation, and its name is the sodium ion
5. There are no variable charge cations in NaCl
6. The name of this compound is **sodium chloride**.

Example: Determine the name for $\mathbf{Sr(NO_3)_2}$**.**

1. The cation is Sr and the anion is NO_3.
2. $Sr(NO_3)_2$ is neutral (no charges are present in the formula)
3. The anion, the nitrate polyatomic ion, has a -1 charge
4. Sr is a fixed charge cation, and its name is the strontium ion
5. There are no variable charge cations in $Sr(NO_3)_2$
6. The name of this compound is **strontium nitrate**.

Example: Determine the name for **$Fe(NO_3)_3$.**

1. The cation is Fe and the anion is NO_3.
2. $Fe(NO_3)_3$ is neutral (no charges are present in the formula)
3. The anion, the nitrate polyatomic ion, has a -1 charge
4. There are no fixed charge cations in $Fe(NO_3)_3$
5. Iron is a variable charge cation; therefore, we must use the formula to calculate the charge on the iron atom. $charge_{Fe}$ = -(# nitrates)($charge_{nitrate+}$) / (# Fe atoms) = - (3)(-1) / (1) = **+3**; therefore, this is the **iron(III) ion**.
6. The name of this compound is **iron(III) nitrate**.

Example: Determine the name for **$Ru_3(PO_4)_2$.**

1. The cation is Ru and the anion is PO_4.
2. $Ru_3(PO_4)_2$ is neutral (no charges are present in the formula)
3. The anion, the phosphate polyatomic ion, has a -3 charge
4. There are no fixed charge cations in $Ru_3(PO_4)_2$
5. Ruthenium is a variable charge cation; therefore, we must use the formula to calculate the charge on the ruthenium atom. $charge_{Ru}$ = -(# phosphates)($charge_{phosphate}$) / (# Ru atoms) = - (2)(-3) / (3) = **+2**; therefore, this is the **ruthenium(II) ion**.
6. The name of this compound is **ruthenium(II) phosphate**.

Part F: Nomenclature for Binary Nonmetal Covalent Molecules

Not all compounds are ionic; indeed, many compounds *share* their electrons over the respective atoms. This class of compound is called *covalent*, and they are formed when two nonmetal elements combine.

The simplest covalent compounds are the elements that exist naturally in pairs; we refer to them as *diatomics*. These are crucial to a successful chemistry experience, and memorization is straightforward using the following acronym:

Name	Compound	Acronym
Hydrogen	H_2	*Have*
Nitrogen	N_2	*No*
Fluorine	F_2	*Fear*
Oxygen	O_2	*Of*
Iodine	I_2	*Ice*
Chlorine	Cl_2	*Clear*
Bromine	Br_2	*Brew*

In addition to the diatomics, several other nonmetals exist naturally in elemental form as combinations of more than one atom. **Phosphorus** exists naturally as P_4, and **sulfur** exists as S_8.

Most nonmetal covalent compounds have more than one type of element. Since there is no ionic charge present in these molecules, we cannot use the system developed above for ionic compounds, and a new method must be used. We will use the **Greek prefixes** for our compounds; they are:

1	*mono*	6	*hexa*
2	*di*	7	*hepta*
3	*tri*	8	*octa*
4	*tetra*	9	*nona*
5	*penta*	10	*deca*

The Greek prefixes refer to the number of atoms present in the molecule. For example, "dinitrogen" implies two nitrogen atoms since the prefix *di* stands for two.

When writing systematic names for binary nonmetal covalent compounds, use the *least electronegative atom first*. The topic of electronegativity will be discussed in Chem 222, but for now, the element listed first (either in the formula or the name) will be the least electronegative element.

Just as with cations in ionic compounds, use the normal element name for the least electronegative element. If more than one exist, use the Greek symbols to represent how many. The *most* electronegative element receives an *-ide* ending (as with anions in ionic compounds) as well as a Greek prefix, *even for single elements.* This is an important distinction between the most and least electronegative elements in binary compounds: the least electronegative element uses Greek symbols only if two or more atoms are present, while the more electronegative element gets an *-ide* ending *and* a Greek prefix *regardless* of the number of atoms present.

Examples:	**NO**	nitrogen monoxide
	N_2O	dinitrogen monoxide
	NO_2	nitrogen dioxide
	P_2O_3	diphosphorus trioxide
	P_2O_5	diphosphorus pentoxide

In addition, there are several common names of binary covalent compounds that you should be familiar with including the following:

Common Name	Formula	Systematic Name
water	H_2O	dihydrogen monoxide
ammonia	NH_3	nitrogen trihydride
laughing gas	N_2O	dinitrogen monoxide
nitric oxide	NO	nitrogen monoxide
phosphine	PH_3	phosphorus trihydride
hydrazine	N_2H_4	dinitrogen tetrahydride
hydrogen sulfide	H_2S	dihydrogen monosulfide

Part G: Nomenclature for Acids and Bases

Acid and base theory shall be discussed in detail during CH 223, but recognizing common acids and bases is important for all chemists. Acids and bases require water to become active; hence, Part G assumes all of the compounds mentioned have been dissolved in water.

Acids contain **H^+**, the **hydrogen ion**. Acids are created when hydrogen ions combine with halogens. If no oxygen atoms are present, add the *hydro-* prefix and an *-ic acid* suffix to find the acid name.

Examples:	**HCl**	hydrochloric acid
	HBr	hydrobromic acid
	HI	hydroiodic acid

If oxygen atoms are present in the halogen acid, use the following table:

Prefix and/or Suffix	Name	Formula
hydro-, -ic	hydrochloric acid	HCl
hypo-, -ous	hypochlorous acid	$HClO$
-ous	chlorous acid	$HClO_2$
-ic	chloric acid	$HClO_3$
per-, -ic	perchloric acid	$HClO_4$

Similar rules apply to bromide or iodide, but not fluoride.

Other common names for acids include:

$\mathbf{HNO_3}$	nitric acid	$\mathbf{H_3PO_4}$	phosphoric acid
$\mathbf{HNO_2}$	nitrous acid	$\mathbf{H_3PO_3}$	phosphorous acid
$\mathbf{H_2SO_4}$	sulfuric acid	$\mathbf{HC_2H_3O_2}$	acetic acid
$\mathbf{H_2SO_3}$	sulfurous acid	**HCN**	hydrocyanic acid
$\mathbf{H_2CO_3}$	carbonic acid	**HF**	hydrofluoric acid

More assistance with naming acids can be found in the handout, "**Guide to Common Polyatomic Ions and the Corresponding Acids**" available in the CH 221 Companion or on the CH 221 website.

One final note about acids: *technically*, an acid is only an acid if dissolved in water (i.e. if *aqueous*, with an *aq* state. If not in water, the acidic properties are lost (at least for CH 221!), and the compound should probably be written as either a binary nonmetal covalent molecule (Section F) or, if the acid contains a polyatomic ion, as a fixed charge metal with a nonmetal. Consider the following examples:

HCl(aq)	hydrochloric acid	*This is truly an acid since HCl is dissolved in water*
HCl(g)	hydrogen monochloride	*This is not an acid - no water! - so name this compound as a covalent compound*
$\mathbf{HNO_2(aq)}$	nitrous acid	*This is a true acid, dissolved in water*
$\mathbf{HNO_2(g)}$	hydrogen nitrite	*This is not an acid - no water! - so name this compound as a fixed charge metal + nonmetal due to the polyatomic ion (nitrite) present*

If a designation of state (i.e. aqueous, gas, solid, etc.) is not provided, then the naming system used is up to the observer (i.e. take your pick! ☺)

Bases contain $\mathbf{OH^-}$, the **hydroxide ion**. Bases consist of a metal cation with the hydroxide anion; hence, their nomenclature will be similar to that of Parts C, D and E, above.

Examples:	**NaOH**	sodium hydroxide
	$\mathbf{Fe(OH)_3}$	iron(III) hydroxide
	$\mathbf{NH_4OH}$	ammonium hydroxide

Part H: Final Words

Understanding chemical nomenclature rules and being able to write formulas for compounds can be thought of as learning to read and write a language. At first, the symbols and rules do not make much sense, but as time progresses, you master the language and a moment of euphoric inspiration occurs when "it all falls into place." Regrettably, inspiration only occurs after time has been spent practicing the material. The more you practice, the faster you will master the material.

Remember that there are five general classes of compounds:

Compound Class	Example
Fixed charge cation + anion	Al_2O_3 - aluminum oxide
Variable charge cation + anion	Fe_2O_3 - iron(III) oxide
Nonmetal binary covalent compound	P_2O_3 - diphosphorus trioxide
Acid	HIO_3 - iodic acid
Base	$Al(OH)_3$ - aluminum hydroxide

Each has specific rules to learn and master. Determining the charge of variable charge cations can be difficult at first, but application of the formulas in Part D and Part E should alleviate the distress.

...oh, wait, one more thing: **<u>Waters of Hydration</u>** or **<u>Hydrated Compounds</u>** show up occasionally with a "dot water" after the name of another chemical. If you see one, add the appropriate Greek prefix plus "hydrate." Examples of hydrated compounds:

$MgSO_4.6\ H_2O$ would be **magnesium sulfate hexahydrate**
$Cu(NO_3)_2\ .2\ H_2O$ would be **copper(II) nitrate dihydrate**
$Mn(BrO_3)_3.4\ H_2O$ would be **manganese(III) bromate tetrahydrate**

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Chemical Nomenclature Worksheet **Name:**

Complete the worksheets below and turn in on the due date.

Section One: Ion Names

Complete the chart using the appropriate elemental ion or polyatomic ion name or symbol. The first row has been filled in as an example. A list of polyatomic ions (page I-3-3) might prove helpful.

Ion	Name	Ion	Name
Na^{+}	sodium ion	F^{-1}	fluoride ion
Li^{+}	___________________	___________________	hydride ion
___________________	gold(III) ion	___________________	hydroxide ion
Mo^{3+}	___________________	___________________	cyanide ion
W^{2+}	___________________	SCN^{-1}	___________________
___________________	gold(I) ion	BrO^{-1}	___________________
Mn^{2+}	___________________	___________________	bromite ion
___________________	platinum(IV) ion	___________________	acetate ion
___________________	zirconium(II) ion	CrO_4^{2-}	___________________
Mt^{3+}	___________________	___________________	dichromate ion
Mg^{2+}	___________________	___________________	phosphide ion
___________________	vanadium(II) ion	___________________	phosphate ion
Cr^{3+}	___________________	___________________	phosphite ion
Cr^{2+}	___________________	$S_2O_3^{2-}$	___________________
___________________	tantalum(V) ion	IO_4^{-1}	___________________
Ni^{2+}	___________________	___________________	iodate ion
___________________	silver ion	___________________	hypoiodite ion
___________________	ammonium ion	MnO_4^{-1}	___________________

Section Two: Ions from Formulas

Write the ions that you would expect from the following compounds

Example: NaCl would give: Na^+, Cl^-

Example: $Fe_3(PO_4)_2$ would give: Fe^{2+}, PO_4^{3-}

LiBr would give:

$MgCl_2$ would give:

Na_2O would give:

VCl_2 would give:

$Fe(NO_3)_3$ would give:

$U(ClO_3)_4$ would give:

Section Three: Nomenclature from Ion Names

Complete the chart using the appropriate compound name using the ions given. The first row has been filled in as an example.

Cation	Anion	Compound Name
potassium	iodide	potassium iodide
magnesium	oxide	___________________
rhodium(III)	chloride	___________________
lead(IV)	chlorate	___________________
gold(I)	cyanide	___________________
cobalt(II)	nitrate	___________________
barium	hydroxide	___________________
ammonium	phosphate	___________________

Section Four: Writing Formulas Using Nomenclature

Complete the chart by providing the correct ion symbols (with the charge) and the correct formula for each compound. The first row has been filled in as an example.

Compound	Cation	Anion	Formula
calcium nitrate	Ca^{2+}	NO_3^{-1}	$Ca(NO_3)_2$
gallium bromide	______________	______________	________________________________
silver nitrate	______________	______________	________________________________
bismuth(III) chloride	______________	______________	________________________________
sodium acetate	______________	______________	________________________________
titanium(II) hypochlorite	______________	______________	________________________________
lithium permanganate	______________	______________	________________________________
iron(III) oxalate	______________	______________	________________________________
cesium chloride	______________	______________	________________________________

Section Five: Chemical Nomenclature Using Formulas

Complete the chart by providing the correct ion symbols (with the charge) and the correct name for each formula. The first row has been filled in as an example.

Formula	Cation	Anion	Name
$Ca(IO_3)_2$	Ca^{2+}	IO_3^-	calcium iodate
ZnS	______________	______________	__________________________
$Sr_3(PO_3)_2$	______________	______________	__________________________
$Ga_2(SO_4)_3$	______________	______________	__________________________
$V(SCN)_5$	______________	______________	__________________________
$NaMnO_4$	______________	______________	__________________________
$(NH_4)_2S$	______________	______________	__________________________
NH_4NO_2	______________	______________	__________________________
$CrCl_6$	______________	______________	__________________________

Section Six: Nonmetal Binary Covalent Compounds

Complete the chart by providing either the correct formula or name. The first row has been filled in as an example.

Name	Formula	Name	Formula
nitrogen dioxide	NO_2	phosphorus trichloride	PCl_3
	SCl_4	sulfur hexachloride	
hydrogen monochloride			$H_2S(g)$
	PI_3	disulfur dichloride	
dinitrogen tetraoxide			N_2O_3
antimony trichloride			$SbCl_5$
	SiO	carbon monoxide	
	SiO_3	carbon dioxide	
phosphorus trihydride			NO

Section Seven: Acids and Bases

Complete the chart by providing either the correct formula or name. The first entry has been filled in as an example. Use acid and base names only in this section.

Name	Formula	Name	Formula
hydrobromic acid	HBr	phosphoric acid	
	HBrO	phosphorous acid	
bromous acid			HCN
	$HBrO_3$	acetic acid	
perbromic acid			NaOH
sulfuric acid			TiOH
	H_2SO_3	potassium hydroxide	
	HNO_3	iron(III) hydroxide	
nitrous acid			$Mg(OH)_2$

Section Eight: Combined Problems: Complete the chart by providing either the correct formula or name.

Name	**Formula**
____________________	$HCl(aq)$
____________________	$HCl(g)$
potassium chloride	____________________
____________________	N_2O_4
nitrogen disulfide	____________________
____________________	$LiClO_3$
aluminum dichromate	____________________
____________________	$FeSO_4$
carbonic acid	____________________
____________________	SO_3
____________________	$(NH_4)_2CO_3$
potassium dihydrogen phosphate	____________________
potassium hydrogen phosphate	____________________
____________________	P_4O_{10}
____________________	$TbBr_6$
____________________	$ThBr_3$
____________________	$TlBr$
____________________	$TiBr_4$
____________________	$TeBr_2$
tetrasulfur decaoxide	____________________
sodium hydrogen carbonate	____________________
____________________	$In(C_2H_3O_2)_3$
____________________	$Mg(ClO_4)_2.6\ H_2O$

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CH 221 Fall 2024:
"Empirical Formula"
(online) ***Lab*** *- Instructions*

Note: **This is the lab for section W1 of CH 221 only.**

- *If you are taking section 01 or section H1 of CH 221*, please use this link:
  http://mhchem.org/s/4a.htm

Step One:

Watch the lab video for the "Empirical Formula" lab, found here:
http://mhchem.org/w/4.htm
Record the data found at the ***end*** of the lab video on page Ib-4-7.

Step Two:

Complete pages Ib-4-7 through Ib-4-11 using the "Empirical Formula" video and your lecture notes. Include your name on page Ib-4-7!

Step Three:

Submit your lab (pages Ib-4-7 through Ib-4-11 *only* to avoid a point penalty) **as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, October 23 by 11:59 PM.** I recommend a free program (ex: CamScanner, https://camscanner.com) or a website (ex: CombinePDF, https://combinepdf.com) to convert your work to a PDF file.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Empirical Formula

One of the fundamental statements of the atomic theory is that elements combine in simple whole number ratios. This observation gives support to the theory of atoms, since one would expect whole atoms to combine. Furthermore, it is observed the combining ratio for a given compound is constant regardless of the origin of the pure substance. This is known as the **Law of constant composition**. The mass contribution of each atom in a compound is a function of the number of atoms in the simplest formula and the relative mass of each atom. The mass contribution is usually referred to as the **percent composition** of a compound.

The **empirical formula** represents the smallest whole number ratio of atoms in a compound. The **molecular formula** represents the actual number of atoms in a compound. The molecular formula may be the same as the empirical formula or it may be a multiple of the empirical formula. For example, hydrogen peroxide has a molecular formula of H_2O_2 and an empirical formula of HO, while water has a molecular formula of H_2O and an empirical formula of H_2O. *No fractions should be present in empirical or molecular formulas!*

Because atoms combine in a definite ratio, the mass composition or **percent by mass** of a compound is fixed. You can determine the mass contribution of each element in a compound by the number of atoms in the simplest formula and the relative mass of each atom. This is typically referred to as the **percent composition** by mass of a compound. For example, $CuCO_3$ is always **38.847%** oxygen by mass:

$$\%\ \text{Oxygen} = \left(\frac{\text{3 oxygen atoms x 15.999 amu per oxygen}}{\text{123.554 amu total molar mass for CuCO}_3}\right) \text{x } 100\%$$

In this experiment you will determine the empirical formula of a hydrated compound of copper and chloride. You will first remove the water from the compound by heating the sample. Any mass lost is water. To determine the mass of copper in the compound, a simple exchange reaction with zinc is performed. Zinc is referred to as an active metal; in contact with a solution containing copper ions, the zinc metal will react to convert the copper ions into copper metal (zinc is transformed into zinc chloride.) As long as excess zinc is added, all of the copper ion should be removed as copper metal, which is easily massed. Any undetermined, remaining mass is chloride.

Once the mass of each component of the compound is determined (water, copper, chloride), you can calculate their corresponding moles and determine the empirical formula.

Because you are novice chemists, your experimental data will not be perfect. By following significant digit rules and using reasonable analytical deduction, you should be able to determine the formula of the hydrate. You can discuss your rounding and sources of error in your lab report.

In addition, you will be able to calculate the **theoretical yield** of copper based on the amount of zinc reacted. Comparing the **actual yield** of copper recovered to the theoretical yield, you can determine a **percent yield** of copper.

Operating a Bunsen Burner

Bunsen burners rely on the combustion of natural gas. An optimal mixture of gas (methane, CH_4) and air (oxygen) will produce a flame with an obvious blue (oxidizing) cone. **To set up your Bunsen burner:**

1. Attach one end of the rubber tubing to the sidearm at the base of the Bunsen burner and attach the other end to a gas outlet. (Be certain that the outlet is labeled "gas".)
2. Adjust the air (oxygen) intake to halfway (see picture below).
3. Adjust your ring stand to the correct height. The ideal flame should be only 2-3 cm above the burner. The ring and stand are metal and will be too hot to adjust once you begin heating. If you are unsure how to properly set up your ring stand, politely ask your instructor for assistance.
4. Make sure you can obtain a spark from the striker before proceeding.

To light your Bunsen burner you must do two things simultaneously:

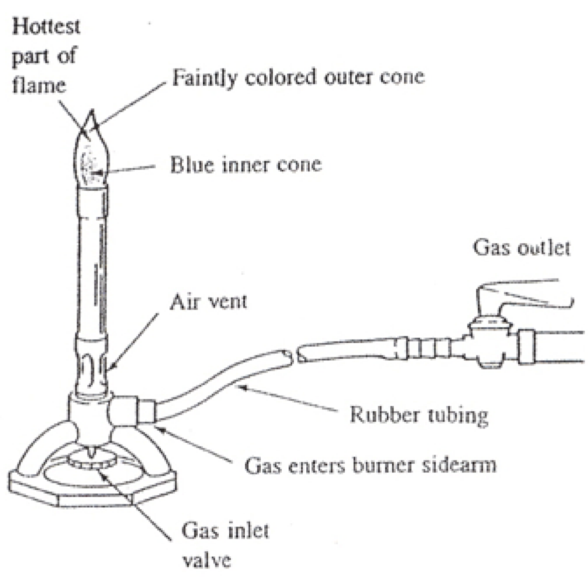
1. Open the gas inlet valve on the burner about halfway (see picture below).
2. As you open the gas outlet, light the burner by bringing the striker from the side to the top.

Do not leave the gas outlet open without the burner lit. It is unsafe to allow natural gas to enter the lab. Always check that the gas outlet is turned off when you are not using the burner. If you are unsure how to light the burner, please patiently ask your instructor for assistance.

To adjust your Bunsen burner:

Adjust the air (oxygen) intake until the flame becomes two concentric cones about 2-3 cm above the burner. The outer cone will be only faintly (dark blue) colored, but the inner cone will be a light blue color. The hottest part of the flame is at the tip of the inner light blue cone. If the flame is luminous and yellow or orange, but not blue, the air vent is not well adjusted. To adjust the height of the flame, adjust the gas flow at the gas outlet or at the gas inlet at the bottom of the Bunsen burner. Proper burner adjustment is crucial for good results. If you are unsure how to adjust the flame, please graciously ask your instructor for help.

Metal and ceramics hold heat well. Be careful when heating metal and ceramics as they will stay hot for quite a while after you turn the burner off. If unsure, remember that patience is a virtue, and wait longer for it to cool.



<u>PROCEDURE</u>: Part A: Dehydration of the copper compound

1. Record the mass of a clean, dry small crucible (no lid) in the "Data" section, below. ***All* measurements obtained in this lab should be to the nearest milligram (0.001g).**

2. Place approximately one gram (1.0 to 1.2 g) of unknown hydrated copper chloride in the crucible. Break up any sizable crystals with your spatula by pressing against the side of the crucible. Record the mass of the crucible (no lid) and sample to the nearest milligram.

3. Place the uncovered crucible on a clay triangle supported by a ring stand. Light your Bunsen burner away from the crucible, adjust the flame as described in the introduction, and ***gently*** heat the crucible as you move the burner back and forth judiciously. Flame should be small to avoid "popping" of sample. *If you overheat your sample*, it will turn into a black nasty liquid, and you must start over... so **heat gently!!!**

4. You should notice the crystals change color as they are heated. Record observations. Why do the crystals change color when heated? Continue to slowly heat the sample until all the crystals are brown. After about 5 minutes, carefully stir sample with a glass stir rod to check color. Once the entire sample is brown, gently heat for two additional minutes.

5. Turn off the burner. Cover the crucible as it cools to prevent the re-absorption of water vapor. Cool the crucible for about 10 minutes. *Caution:* The crucible is ceramic and retains heat. The crucible can severely burn you if you try to touch it before it is cool. Patience is a virtue!

6. After 10 minutes, remove the cover and slowly roll the brown crystals around the crucible. If there is any evidence of green crystals, repeat the heating and cooling process... but if all the crystals appear brown and the crucible is cool, **record the mass of the crucible (no lid) and dehydrated sample**.

Part B: Exchange reaction between the copper compound and zinc

1. Transfer the brown crystals to a small 250 mL beaker. Rinse the crucible with two 5-7 mL portions of deionized H_2O, adding each rinse to the beaker. Swirl the beaker gently to dissolve the crystals and record your observations. Why did the color change?

2. Obtain a piece of clean zinc and record its mass to the nearest 0.001 g. You need at least 0.5-0.8g of zinc; more is fine. Gently slide the piece of zinc into the beaker so that it is submerged in the copper chloride solution. Be careful not to splash. Add ~5 mL of water if your volume is too low.

3. Stir the solution with a glass rod so that as copper forms, it does not adhere to the zinc. Record observations. Allow the reaction to continue until all blue and green color has disappeared from the solution. The solution might have an unattractive grey hue, but no tint of green should remain.

4. Add 10 drops of 10% HCl to the solution and stir thoroughly. This will dissolve any insoluble zinc salts formed and clear up the solution if cloudy.

5. Carefully remove the unreacted zinc metal from the solution using tongs. Inspect the zinc for any adhering copper. Use a wash bottle of deionized water and a rubber policeman to scrape and clean the copper off the zinc into the beaker. Dry the remaining zinc on a paper towel. Record the mass of the dry zinc. (Note: this

must be less than your starting mass, right?) Place the zinc in the waste container when this step is complete.

Part C: Cleaning the copper:

1. Set up a Buchner funnel suction filtration apparatus with a moistened piece of filter paper. Attach the rubber hose to the “VAC” outlet, and only turn the vacuum to a 45 degree angle initially to prevent losing copper.

2. With light suction, carefully decant (pour off) the solution over the copper into the funnel. It is okay if some of the copper is transferred to the funnel.

3. Wash the copper solid in the beaker with about 10 mL of deionized water. Stir thoroughly, allow the copper to settle and carefully decant the wash water into the funnel. Break up any large chunks of copper with your glass stir rod. Repeat with a second 10 mL portion of deionized water.

4. Transfer the copper to the funnel using a small amount of deionized water. Use your wash bottle and rubber policeman to facilitate the transfer all of the copper to the funnel. Rinse any copper adhering to the rubber policeman into the funnel. All of the copper must be transferred to the funnel.

5. Turn off the suction. Add 10 mL of methanol to the funnel. After one minute, turn on the suction (slow at first, then to a roughly 45 degree angle.) Methanol evaporates faster than water and will enhance the drying process.

6. Draw air through the funnel for about 3-5 minutes. Meanwhile, record the mass of a clean, dry watch glass.

7. Transfer the dry copper to the massed watch glass. The transfer must be quantitative; scrape any copper that adheres to the paper on to the watch glass with your spatula or rubber policeman. If the copper is still damp, dry under a heat lamp for 5 minutes or press with a dry piece of filter paper. Allow the sample and watch glass to cool. Record the mass of the copper to the nearest 0.001 g. If you have more than 0.50 grams of copper, your sample is probably wet. It is recommended that you dry it under a heat lamp and take a new measurement.

8. Clean up! Dispose of the liquid methanol waste from the suction filtration apparatus into the appropriate waste bottle. Discard the copper in the garbage can unless directed otherwise by the instructor.

9. Complete the worksheets below using the data obtained in lab.

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Empirical Formula Worksheet

Name:

Complete the worksheets below and turn in on the due date

*Watch the video (***http://mhchem.org/w/4.htm***) to acquire these values:*

Mass of clean, dry, small crucible (g, from video): ______________________

Mass of crucible and copper chloride sample (g, from video): ______________________

Mass of unheated copper chloride sample (g): ______________________

Color of copper chloride sample before heating *(see video)*: ______________________

Mass of crucible and dehydrated copper chloride sample (g, from video): ______________________

Mass of dehydrated copper chloride sample (g): ______________________

Mass of water lost upon heating (g): ______________________

Color of dehydrated copper chloride *(see video)*: ______________________

Color of dehydrated copper chloride sample when dissolved in water *(see video)*: ______________________

Mass of zinc sample before reaction (g, from video): ______________________

Observations as zinc is added to the copper solution *(see video)*:

Mass of dry zinc after reacting with copper (g, from video): ______________________

Mass of zinc that reacted with copper (g): ______________________

Mass of clean, dry watch glass (g, from video): ______________________

Mass of dried copper and watch glass (g, from video): ______________________

Mass of dried copper (g): ______________________

Color of dried copper *(see video)*: ______________________

Mass of chloride (g): *(dehydrated copper chloride - dried copper)* ______________________

Calculations Worksheet *for the* **Empirical Formula** *Lab*
Do not wait until the last minute to start these calculations!

1. Calculate the **mass (grams) of water** lost upon heating the copper chloride sample. Calculate the **moles of water lost** upon heating the hydrated copper compound

2. Calculate the **mass (grams) of copper** collected at the very end of lab; this is your **actual yield** of copper (g). Determine the **mass (grams) of chloride lost** in your sample (mass of dehydrated copper sample – mass of copper collected after filtration = mass of chloride.) Convert mass (grams) of copper into **moles of copper**; also convert mass (grams) of chloride into **moles of chloride.**

3. Use the moles of water, moles of copper and moles of chloride to **find the empirical formula of the hydrated copper chloride**. Round to whole numbers when determining the empirical formula. What is the **name** of your compound?

4. Use the masses of water, copper, chloride, and the original hydrated copper chloride sample to **find the percent copper, percent chloride and percent of water in the original hydrated copper chloride sample**.

5. Show how to calculate the **mass (g) of zinc reacted** (i.e. the initial weight of Zn minus the final weight of Zn after the reaction was complete). Calculate the **theoretical yield of copper** using the formula:

   Theoretical yield (grams) of Cu = (grams of zinc reacted in the reaction) * 0.9720

6. Calculate the **percent yield** of copper. %yield = (actual yield / theoretical yield) x100%. **Comment** on why the percent yield might be greater than 100% in this lab.

<u>POSTLAB QUESTIONS</u>:

1. The *limiting reactant* (also known as the limiting reagent) is defined as the starting substance which is totally consumed in a reaction; the *excess reactant* is a starting material which is still present at the end of a reaction. Which of the reactants was the limiting reactant and which was the excess reactant? (the reactants in this reaction were zinc metal and the unknown copper chloride.) *Briefly* explain your answer.

2. Explain the color changes in part A (from blue to brown and back to blue) *and* in part B (from blue to clear).

   Part A:

   Part B:

3. **Explain** the effect each of the following would have on the experimentally determined %Cu. Use the terms **increase**, **decrease** or have **no effect** to describe the effect on the %Cu.
   a. Some solution splashed onto the bench when the zinc was plopped into the beaker.

   b. The student removed the zinc before the blue color disappeared from the solution.

   c. The student did not completely dry the Cu before the final weighing

4. Your final mass of zinc was less than your initial. What happened to the zinc? What did it become?

5. Determine the %Cl by mass value if the sample was pure anhydrous copper(II) chloride. *Hint:* do not use your data for this question; use the formula... what *is* the formula for copper(II) chloride?

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CH 221 Fall 2024:
"Percent Potassium Chlorate"
(online) ***Lab*** *- Instructions*

Note: **This is the lab for section W1 of CH 221 only.**

- *If you are taking section 01 or section H1 of CH 221*, please use this link:
  http://mhchem.org/s/5a.htm

Step One:

Watch the lab video for the "%$KClO_3$" lab, found here: **http://mhchem.org/w/5.htm**
Record the data found at the ***end*** of the lab video on page Ib-5-3.

Step Two:

Complete pages Ib-5-3 through Ib-5-6 using the "%$KClO_3$" video and the actual lab instructions on pages Ib-5-2. Include your name on page Ib-5-3!

Step Three:

Submit your lab (pages Ib-5-3 through Ib-5-6 *only* to avoid a point penalty**) as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, October 30 by 11:59 PM.** I recommend a free program (ex: CamScanner, https://camscanner.com) or a website (ex: CombinePDF, https://combinepdf.com) to convert your work to a PDF file.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Percent Potassium Chlorate in a Mixture

Potassium chlorate ($KClO_3$) decomposes on heating to produce potassium chloride and oxygen. The Law of Conservation of Mass states that the mass of the reactants (potassium chlorate) will equal the mass of the products (potassium chloride and oxygen). Since oxygen is a gas, the mass of the final solid will be less than the starting weight. The mass loss is equal to the mass of oxygen.

In this experiment, you will begin with a sample that is a mixture of potassium chlorate and potassium chloride. Your objective is to determine the percentage by mass of potassium chlorate in the original mixture. Upon heating, only the potassium chlorate will decompose. Using the balanced equation and the fact that all the mass that is lost is oxygen gas, you can use stoichiometry to calculate the mass of potassium chlorate in the original mixture. A catalyst, manganese(IV) oxide, is added to the reaction mixture in order to speed up the reaction. Like all catalysts, the same amount of catalyst is present at the end of the reaction as in the beginning. Therefore, we will include the mass of the catalyst in with the mass of the crucible.

To ensure that the decomposition is complete, the product must be heated to a constant weight. After the first heating, cooling and weighing, the sample must be heated again, cooled and reweighed. This process is continued until two successive weights are within 5 mg of each other (up to four heating cycles.)

PROCEDURE (this is what we would have done in the lab room under "normal" circumstances)**:** - *You must wear safety goggles while performing this lab!* ***All*** *mass measurements should be recorded to the milligram (0.001 g.)*

1. Set up a ring stand with a triangle as demonstrated by your teacher. The small white crucible should fit inside the triangle.

2. Place about 0.5 g of manganese(IV) oxide into a clean, dry small white crucible. Heat the crucible and catalyst with a Bunsen burner for about 3 minutes to drive off any moisture that may be in the catalyst or crucible. Wear safety glasses at all times if a Bunsen burner is operational at your lab bench!

3. When the crucible is cool enough to touch, record the entire mass to the nearest 0.001g.

4. Add between 2.0 to 2.5 grams of the unknown mixture to the crucible. Mix the contents to obtain a somewhat uniform mixture. Record the mass of the crucible plus catalyst plus mixture to the nearest 0.001 g. Be sure to also record your unknown number!

5. Begin heating the crucible gently at first followed by a more aggressive treatment for a total of 10 minutes. Be aware that the sample may begin to bubble and spurt; if this happens, turn the flame down a bit.

6. Allow the sample to cool to room temperature. Record the mass to the nearest 0.001 g.

7. Reheat your sample for 5 minutes. Cool and record the mass. If your mass is within 0.005 g of the mass after the previous heating with the unknown sample, congratulations, you can move on to calculations. If not, you should reheat, cool, and weigh until you have two successive masses within 0.005 g of each other. Clean up and put away your equipment (all waste in this lab can be washed down the drain with water.)

Percent Potassium Chlorate in a Mixture

YOUR NAME: ____

DATA: *Watch the video (***http://mhchem.org/w/5.htm***) to acquire these values:*

mass of empty crucible (g): ______________________

crucible after heating with MnO_2 (g): ______________________

crucible, MnO_2 & $KClO_3$ mixture before heating (g): ______________________

crucible, MnO_2 & $KClO_3$ mixture after heating for ten minutes (g): ______________________

crucible, MnO_2 & $KClO_3$ mixture after heating for additional ten minutes (g):

____

CALCULATIONS: *The video (***http://mhchem.org/w/5.htm***) might help you with these calculations. Clearly* ***show all work*** *in the area provided,* ***watch significant figures*** *and circle final answers*

1. Write a balanced equation for the decomposition of potassium chlorate into potassium chloride and oxygen gas.

2. Using the data from the video, determine the **mass of the $KClO_3$ mixture** used in this experiment (no MnO_2!).

3. Using the data from the video, determine the **mass of oxygen lost** upon heating the mixture. This answer will be the **α** (below) in the equation.

4. Determine the **molar mass** of **oxygen** (**O_2**) to 0.01 g/mol. This answer will be the **β** (below) in the equation.

5. Determine the **molar mass** of **potassium chlorate** (**$KClO_3$**) to 0.01 g/mol. This answer will be the **δ** (below) in the equation.

6. Use the balanced equation and your values of **α** (the mass of oxygen lost), **β** (the molar mass of oxygen) and **δ** (the molar mass of potassium chlorate) to **determine the mass of potassium chlorate present in the original mixture** (this is the $KClO_3$ that decomposed in this experiment and is represented by **λ**, below, in the equation.) *Show your work! This is the "grams - moles - moles - grams' application we will be talking a lot about soon!* The equation to use:

$$\lambda \text{ g KClO}_3 = (\alpha \text{ g O}_2 \text{ lost}) * \left(\frac{1 \text{ mol O}_2}{\beta \text{ g O}_2}\right) * \left(\frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2}\right) * \left(\frac{\delta \text{ mol KClO}_3}{1 \text{ mol KClO}_3}\right)$$

7. **Determine the percentage of potassium chlorate in the original white mixture** using your answers from step 6 (the pure KClO3) and step 2 (the mass of the original mixture.)

POSTLAB QUESTIONS:

1. A white powder is a mixture of magnesium carbonate and magnesium oxide. Upon heating, the magnesium carbonate decomposes into magnesium oxide and carbon dioxide. If you have 1.897 g of the mixture and after heating are left with 1.494 g of magnesium oxide, calculate the weight percent of magnesium carbonate in the original mixture. *Hint:* Start by writing a balanced reaction, and remember the 1.897 g value is not pure!

2. Calculate the % oxygen by mass for the following (show calculations): a) $LiNO_3$ b) $NaHCO_3$ *Hint:* first find the molar mass (to 0.01 g/mol) of the compound!

3. If we had doubled the mass of the original mixture and completed the lab as written, would the calculated $\%KClO_3$ have changed? Explain.

CH 221 Fall 2024:
“Net Ionic Reactions”
(online) ***Lab*** *- Instructions*

Note: **This is the lab for section W1 of CH 221 only.**

- *If you are taking section 01 or section H1 of CH 221*, please use this link: http://mhchem.org/s/6a.htm

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Step One:

Watch the lab video for the “Net Ionics” lab, found here: **http://mhchem.org/w/6.htm** Write down the “Observations” data towards the end of the lab and include it in the appropriate place in your lab report.

Step Two:

Complete pages Ib-6-9 through Ib-6-12 using the “Net Ionics” video and the actual lab instructions on pages Ib-6-2 through Ib-6-8. Include your name on page Ib-6-9!

Step Three:

Submit your lab (pages Ib-6-9 through Ib-6-12 *only* to avoid a point penalty) **as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, November 6 by 11:59 PM.** I recommend a free program (ex: CamScanner, https://camscanner.com) or a website (ex: CombinePDF, https://combinepdf.com) to convert your work to a PDF file.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Net Ionic Reactions in Aqueous Solutions

Double replacements are among the most common of the simple chemical reactions. Consider the hypothetical reaction:

$$\mathbf{AB + CD \rightarrow AD + CB}$$

where AB exists as A^+ and B^- ions in solution and CD exists as C^+ and D^- ions in solution. As the ions come in contact with each other, there are six possible combinations that might conceivably cause a chemical reaction. Two of these combinations are the meeting of ions of like charge; that is, A^+ with C^+ and B^- with D^-. Since particles with like electrical charges repel each other, no reaction will occur. Two other possible combinations are those of the original two compounds; that is A^+ with B^- and C^+ with D^-. This combination would lead to no change. Thus the only possibilities for chemical reaction are the combination of each of the positive ions with the negative ion of the other compound; that is, A^+ with D^- and C^+ with B^-.

Example 1: When solutions of sodium chloride and silver nitrate are mixed, the combination of silver cations and chloride anions form silver chloride, which precipitates and settles to the bottom of the container. Note that the states of matter are included: (aq) substance is soluble in water; (s) substance is insoluble in water (solid precipitate)

$$\mathbf{NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)}$$

This combination – the **molecular equation** - of chemicals is referred to as a **precipitation reaction** since an insoluble solid, AgCl, is present as a product.

Example 2: When solutions of potassium chloride and sodium nitrate are mixed, the equation for the hypothetical double replacement reaction is:

$$\mathbf{KCl(aq) + NaNO_3(aq) \rightarrow KNO_3 + NaCl}$$

But has there been a reaction? Double replacement reactions occur when one of the following is formed as a product of the reaction:

a. an **insoluble solid** (precipitate) - check the solubility table in this lab report. If a solid has formed, this is called a **precipitation reaction**.
b. a **gas** - for example, CO_2 (from H_2CO_3), SO_2 (from H_2SO_3), or NH_3 (from NH_4OH). If a gas has formed, this is called a **gas forming reaction**.
c. **water** from an acid (source of H^+) and a base (source of OH^{-1}). If water forms from an acid and a base (along with an ionic "salt"), this is called an **acid-base reaction**.

Using the solubility table (see below) we find both KNO_3 and NaCl are water-soluble (aqueous, or *aq* for short) products. There is no precipitate, gas or water from this combination. Thus in Example 2, we conclude that even though we can write an equation for a double replacement reaction, no reaction occurs. We simply end up with a solution containing four kinds of ions - Na^+, K^+, Cl^-, and NO_3^-.
Thus the **molecular equation** is more properly written:

$$\mathbf{KCl(aq) + NaNO_3(aq) \rightarrow KNO_3(aq) + NaCl(aq)}$$

but in terms of "if something happens", we should write:

$$\mathbf{KCl(aq) + NaNO_3(aq) \rightarrow No\ Reaction}$$

Aqueous solutions of sodium chloride and silver nitrate will undergo double replacement reaction to produce a white precipitate of silver chloride and aqueous sodium nitrate. What would happen if we just mixed solid silver nitrate and solid sodium chloride together? No apparent reaction occurs. Thus the water performs some necessary function that allows the reaction to proceed. When ionic compounds are dissolved in water, the ions separate and become surrounded by water molecules. This frees the ions from the crystal lattice, allowing them to move throughout the solution and react with appropriate ions of opposite charge.

To clarify what reaction occurs between ions in electrolyte solutions, we write **total ionic equations**. In this type of equation, compounds are written in the form in which they are predominately present in water. Most notably, soluble compounds (aq) are written as ions in solution. Others (s, l, g) are written in their molecular form.

For example, if we write the total ionic equation for the double replacement precipitation reaction (See Example 1) we get the following:

Total Ionic Equation: $Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + AgCl(s)$

Note that during the course of reaction, there has been no change in the Na^+ and NO_3^- ions. These unreacted ions (**spectator ions**) can be left out of the total ionic equation to yield the **net ionic equation**. Net ionic equations tell us only what is actually changing during reaction.

Net Ionic Equation: $Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$

Another example is illustrated below for the reaction of nitric acid and a dilute aqueous solution of barium hydroxide (an **acid-base reaction**):

Molecular Equation: $2\ HNO_3(aq) + Ba(OH)_2(aq) \rightarrow 2\ H_2O(l) + Ba(NO_3)_2(aq)$

Total Ionic Equation: $2\ H^+(aq) + 2\ NO_3^-(aq) + Ba^{2+}(aq) + 2\ OH^-(aq) \rightarrow 2\ H_2O(l) + Ba^{2+}(aq) + 2\ NO_3^-(aq)$

Net Ionic Equation: $2\ H^+(aq) + 2\ OH^-(aq) \rightarrow 2\ H_2O(l)$
or, reduced to simplest form: $\mathbf{H^+(aq) + OH^-(aq) \rightarrow H_2O(l)}$
This is an example of an **acid-base reaction**.

We will use the following solubility table in CH 221:

CH 221 Solubility Table for Ionic Compounds

SOLUBLE COMPOUNDS	EXCEPTIONS
Almost all salts of Na^+, K^+, NH_4^+	
Salts of nitrate, NO_3^- chlorate, ClO_3^- perchlorate, ClO_4^- acetate, $CH_3CO_2^-$	
Almost all salts of Cl^-, Br^-, I^-	Halides of Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds containing F^-	Fluorides of Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
Salts of sulfate, SO_4^{2-}	Sulfates of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

INSOLUBLE COMPOUNDS	EXCEPTIONS
Most salts of carbonate, CO_3^{2-} phosphate, PO_4^{3-} oxalate, $C_2O_4^{2-}$ chromate, CrO_4^{2-}	Salts of NH_4^+ and the alkali metal cations
Most metal sulfides, S^{2-}	
Most metal hydroxides and oxides	

Note: **Use this table for *all* CH 221 solubility questions!**

The following examples should help you understand how the solubility table works and also how to complete the written part of this assignment.

Example: Is $PbSO_4$ soluble in water? What species are present in a water solution?

Answer: To solve, notice how $PbSO_4$ has a sulfate ion (SO_4^{2-}). Most salts of sulfate are soluble (i.e. they dissolve in water, or are (aq)), but salts of sulfate with a Pb^{2+} ion are an exception to this rule. Hence, $PbSO_4$ is insoluble (it does not dissolve in water), and we would write $PbSO_4$ in water as $PbSO_4$(s). Indeed, $PbSO_4$(s) is the only species present in a water solution.

Example: Is Li_2S soluble in water? What species are present in a water solution?

Answer: To solve, notice how Li_2S has a sulfide ion (S^{2-}). Most salts of sulfide are insoluble (i.e. they do not dissolve in water, or are (s)), but salts of alkali metals (which includes Li^+) are an exception to this rule. Hence, Li_2S is soluble (it *does* dissolve in water), and we would write Li_2S in water as Li_2S(aq).

Further, because Li_2S dissolves in water, we should really write it as the dissociation ions – i.e. the molecular Li_2S dissociates into lithium and sulfide ions. The species which are present in a water solution of Li_2S(aq) are 2 Li^+(aq) and S^{-2}(aq) (molecular Li_2S does not exist in water.)

Example: Describe $CaCO_3$, $RbMnO_4$ and $TiCrO_4$ in terms of their solubility in water.

Answer: Use the solubility table to answer these types of questions:
- **$CaCO_3$** has a carbonate ion, and calcium is not an alkali metal or ammonium, so CaCO3 is insoluble in water. We would write it as $CaCO_3$(s).
- **$RbMnO_4$** has an alkali metal ion (rubidium), so by default, all alkali metals are water soluble (no exceptions, at least in CH 221!) So we would write this species as $RbMnO_4$(aq) or, as dissolve ions, we would write it as Rb^+(aq) and MnO_4^-(aq).
- **$TiCrO_4$** has a chromate ion, and titanium is not an alkali metal or ammonium, so $TiCrO_4$ is insoluble in water. We would write it as $TiCrO_4$(s).

Example: Write the balanced molecular equation and net ionic reaction that occurs between potassium nitrate and calcium chloride in water. Classify this reaction type.

Answer: First, we need the chemical equations for potassium nitrate and lithium chloride. They are KNO_3 and $CaCl_2$. Notice the ionic charges on the cations and anions: K^+, NO_3^{-1}, Ca^{2+}, Cl^{-1}.

All of the reactions in this lab are "double displacement" – the reactant cations will switch places, forming new products. *The ionic charges will not change upon going from reactant to product*. Potassium and chloride will come together as KCl (only one Cl^{-1} for every one K^{+1}), and calcium and nitrate will come together as $Ca(NO_3)_2$ (two nitrates being needed for every calcium +2 ion.) Initially, the equation looks like this:

$$KNO_3 + CaCl_2 \rightarrow KCl + Ca(NO_3)_2$$

Notice the parentheses used for more than one polyatomic ion ($Ca(NO_3)_2$) but parentheses are not used when only one polyatomic ion is used (KNO_3).

We need to balance this reaction and add states of matter. Every compound with potassium (an alkali metal) or nitrate will dissolve in water; $CaCl_2$ is also soluble in water (Ca is not Ag, Pb or Hg), leading to:

$$\mathbf{2\ KNO_3(aq) + CaCl_2(aq) \rightarrow 2\ KCl(aq) + Ca(NO_3)_2(aq)}$$

To *classify* this reaction, our options include: precipitate, acid-base, gas forming, or no reaction. Since no solids have formed, it is not a precipitation reaction. Water has not formed from an acid or base, so this is excluded; and H_2CO_3 and NH_4OH have not formed (see next example), so gas forming is excluded. Indeed, all the reactants and products are (aq), so nothing really happens; classify this reaction as "**no reaction**." Nothing needs to be written for a net ionic reaction because nothing happens!

Example: Write the balanced molecular equation and net ionic reaction that occurs between potassium carbonate and hydrobromic acid in water. Classify this reaction type.

Answer: First, we need the chemical equations for the reactants. They are K_2CO_3 and HBr, and they make K^+, CO_3^{-2}, H^{1+}, Br^{-1}. Performing a "double displacement" on these reactants – the reactant cations will switch places – we get: $K_2CO_3 + HBr \rightarrow KBr + H_2CO_3$

Balancing this reaction and adding states of matter, we get:

$$\mathbf{K_2CO_3(aq) + 2\ HBr(aq) \rightarrow 2\ KBr(aq) + H_2CO_3(aq)}$$

At first, it looks like this is a "no reaction" classification – all states are aqueous – but **make sure you check for H_2CO_3 and NH_4OH** – these two species are the hallmarks of the **gas forming reaction** since both are unstable compounds and further decompose to new products. ***Be watchful for H_2CO_3 and NH_4OH!***

Carbonic acid (H_2CO_3 breaks down into water and carbon dioxide, so really you should write:

$$\mathbf{H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)}$$

and ammonium hydroxide (NH_4OH), which breaks down into water and ammonia, should be written:

$$\mathbf{NH_4OH(aq) \rightarrow H_2O(l) + NH_3(g)}$$

So in this **gas forming** example,

$$\mathbf{K_2CO_3(aq) + 2\ HBr(aq) \rightarrow 2\ KBr(aq) + H_2CO_3(aq)}$$

should be written as a net ionic equation in the following manner:

$$\mathbf{CO_3^{2-}(aq) + 2\ H^+(aq) \rightarrow H_2O(l) + CO_2(g)}$$

PROCEDURE and LAB REPORT:

Use the attached sheets to complete this week's lab. The purpose, conclusion, etc. can be omitted this week, and typing is not required as long as your handwriting is legible. For each reaction,

- **Mix** 1.0 mL (20 drops) of each of the two indicated solutions (below) in a clean (but not necessarily dry) **small** test tube and record observations that might indicate a chemical change has occurred **(color, precipitate, bubbles of a gas, or heat released.)**

- Write the **balanced molecular equation** (double displacement or exchange reaction) for each reaction. Show **states of matter** (**use the solubility table in this lab report for your answers**) and **ionic charges (on ions, *not* molecules)** for all species.

- Write the **total ionic equation** and the **net ionic equation** for each reaction. Be sure to include all states of matter and ionic charges. If all the products are aqueous, no reaction has occurred, and you should write **no reaction** in place of the net ionic equation. Note that even *if no reaction occurs*, **you will still be required to write a balanced molecular equation and the total ionic equation**.

- Finally, **classify each reaction** as **precipitation, acid-base, gas forming or** (if nothing happened) **no reaction**. Remember that gas forming reactions often create unstable precursors (such as $\mathbf{H_2CO_3}$ (which creates CO_2(g) and H_2O(l)) and $\mathbf{NH_4OH}$ (which creates NH_3(g) and H_2O(l)).)

The reactions:

1. Barium Nitrate + Magnesium Sulfate
2. Barium Nitrate + Hydrochloric Acid
3. Barium Nitrate + Sodium Carbonate
4. Iron(III) Chloride + Sodium Hydroxide
5. Iron(III) Chloride + Sodium Phosphate
6. Iron(III) Chloride + Magnesium Sulfate

7. Magnesium Sulfate + Sodium Hydroxide
8. Magnesium Sulfate + Sodium Carbonate
9. Ammonium Oxalate + Barium Nitrate
10. Hydrochloric Acid + Sodium Hydroxide
11. Hydrochloric Acid + Sodium Carbonate
12. Silver Nitrate + Potassium Chromate

13. Silver Nitrate + Iron(III) Chloride
14. Sodium Hydroxide + Ammonium Chloride
15. Sodium Hydroxide + Sulfuric Acid
16. Copper(II) Sulfate + Iron(III) Chloride
17. Copper(II) Sulfate + Sodium Phosphate
18. Acetic Acid + Sodium Carbonate

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Net Ionic Reactions *Worksheet* **Name:**

Complete the following worksheet using the instructions provided. Remember to show states of matter and charges where appropriate. **M** = **Molecular Equation**, **T** = **Total Ionic Equation**, and **N** = **Net Ionic Equation**. **Section W1 must show "Total Ionic" work (the "T" line) for all problems**.

1. Barium Nitrate + Magnesium Sulfate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

2. Barium Nitrate + Hydrochloric Acid Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

3. Barium Nitrate + Sodium Carbonate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

4. Iron(III) Chloride + Sodium Hydroxide Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

5. Iron(III) Chloride + Sodium Phosphate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

6. Iron(III) Chloride + Magnesium Sulfate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

7. Magnesium Sulfate + Sodium Hydroxide Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

8. Magnesium Sulfate + Sodium Carbonate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

9. Ammonium Oxalate + Barium Nitrate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

10. Hydrochloric Acid + Sodium Hydroxide Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

---

11. Hydrochloric Acid + Sodium Carbonate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

---

12. Silver Nitrate + Potassium Chromate Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

---

13. Silver Nitrate + Iron(III) Chloride Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

---

14. Sodium Hydroxide + Ammonium Chloride Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

15. Sodium Hydroxide + Sulfuric Acid Observations: ______________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

---

16. Copper(II) Sulfate + Iron(III) Chloride Observations: ________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

---

17. Copper(II) Sulfate + Sodium Phosphate Observations: ________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

---

18. Acetic Acid + Sodium Carbonate Observations: ________________________

M: ____

T: ____

N: ____

***Type** of reaction (circle **one**):* **precipitation** **acid-base** **gas forming** **no reaction**

Bonus! On a *separate piece of paper*, add an original poem or short story of at least 50 words in length for extra credit… content will not be criticized, but the poem must be original, and short haikus written at the bottom of this page will not count (although the instructor will find them fun to read! ☺) Original music will also count!

CH 221 Fall 2024:
“Unknown Chloride”
(online) ***Lab*** *- Instructions*

Note: **This is the lab for section W1 of CH 221 only.**

- *If you are taking section 01 or section H1 of CH 221*, please use this link: http://mhchem.org/s/7a.htm

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Step One:

Watch the lab video for the “Unknown Chloride” lab, found here:
http://mhchem.org/w/7.htm
Record the data found at the ***end*** of the lab video on page Ib-7-5.

Step Two:

Complete pages Ib-7-5 through Ib-7-9 using the “Unknown Chloride” video and the actual lab instructions on pages Ib-7-2 through Ib-7-3. Include your name on page Ib-7-5!

Step Three:

Submit your lab (pages Ib-7-5 through Ib-7-9 *only* to avoid a point penalty**) as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, November 13 by 11:59 PM.** I recommend a free program (ex: CamScanner, https://camscanner.com) or a website (ex: CombinePDF, https://combinepdf.com) to convert your work to a PDF file.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Determination of an Unknown Chloride

The determination of a soluble chloride salt concentration is a classic titrametric analysis. A titration involves delivering a measured amount of a solution whose concentration is known accurately (the **titrant**) into a solution whose concentration is not known (the **titrate**). The purpose of the titration is to determine the number of moles of titrate present. When the reaction is complete, some physical change is observed, indicating the **endpoint** of the titration. The endpoint of a titration occurs when stoichiometric ratios of reactants are present and must be determined accurately.

In the titration in this lab, a dilute solution of silver nitrate with a known concentration acts as the titrant. It is added to a salt solution with an unknown amount of chloride, i.e. the titrate. Silver chloride, a white insoluble solid will precipitate from the solution. In order to detect when all the AgCl has been precipitated, another reagent is used as an **indicator**. The indicator in this lab, potassium chromate, is yellow and reacts with silver ions to form a bright orange silver chromate precipitate. This solid is slightly more soluble than the silver chloride so it does not form until essentially all the chloride has precipitated from the solution.

$$\mathbf{Ag^+(aq) + Cl^-(aq) \rightleftharpoons AgCl(s)} \text{ at } \textbf{low} \text{ silver ion concentration}$$

$$\mathbf{2\,Ag^+(aq) + CrO_4^{2-}(aq) \rightleftharpoons Ag_2CrO_4(s)} \text{ at } \textbf{higher} \text{ silver ion concentration}$$

All standard solutions must first be standardized using a **primary standard** because of potential evaporation. A primary standard is a solid that is stable and does not pick up water. The primary standard in this experiment is purified sodium chloride.

In this lab you will perform six titrations. In the first three titrations you will use a known amount of a pure NaCl sample to determine the exact concentration of an approximate 0.05 M silver nitrate solution. The endpoint is the first permanent orange-red color of Ag_2CrO_4. From this information one can determine the concentration of the $AgNO_3$. The last three titrations will allow you to find the percentage of chloride in your salt when used in conjunction with the average silver nitrate concentration.

Note: Silver is a heavy metal toxin and should never be flushed down the drain. Dispose of all silver waste (silver nitrate and silver chloride) in the waste bottles provided.

PROCEDURE:

Part A: ***Standardizing the Silver Nitrate Solution***

1. Clean a 50 mL buret with soap and water, then rinse well with water.

2. Fill your buret with silver nitrate from an amber bottle. To prevent contamination, ***never*** add anything to the amber bottle. Fill the buret to the 0.00 mL mark with the $AgNO_3$ solution. Drain 5 mL from the buret into a beaker (to remove air bubbles) and fill to 0.00 mL again. *Note:* $AgNO_3$ is the only solution that will be placed in your buret! Also, do not dispose of $AgNO_3$ in the sink – place this heavy metal in a waste container.

3. Use an analytical balance to weigh three 0.1000 - 0.1200 gram samples of purified NaCl. Record exact mass.

4. Add about 50 mL of distilled water to each sample in a 125 mL Erlenmeyer flask (or larger) to dissolve the NaCl sample. Add about three drops of indicator (K_2CrO_4).

5. Titrate with 0.05 M $AgNO_3$ solution as you continually swirl the flask to a lovely peach end point. As you add the silver nitrate solution initially in short bursts you will see the orange-red color form and disappear as the solution is swirled. As you approach the end point (which should be between 20-40 mL) the color should begin to persist. At this point you should be adding the solution dropwise. Read the buret to the nearest 0.01 mL. Stop when the sample has a permanent faint peach color.

6. Repeat the titration with the second and third samples.

Part B: ***Determination of Percent Chloride***

1. Obtain an unknown chloride salt and record the ID number in your lab notes. Use an analytical balance to weigh **three** 0.1000 - 0.1200 gram samples.

2. To **each** sample add 50 mL of distilled water and 3 drops of K_2CrO_4 indicator solution in an Erlenmeyer flask. Titrate each sample with the standardized silver nitrate solution as in part A.

3. When done, place excess $AgNO_3$ in the waste container and rinse the buret with water before leaving the lab.

CALCULATIONS:

For Part A, calculate the molarity of the silver nitrate solution ([$AgNO_3$]) for each titration. Calculate the Parts Per Thousand (PPT) for [$AgNO_3$] using the "Parts Per Thousand" handout in the “Lab Notes” of the Companion. If your PPT is greater than 30 for the three trials, consider omitting a deviant molarity value to improve your PPT.

For Part B, calculate the percent chloride. (*Note:* Use the average molarity of $AgNO_3$ as determined in part A.) Average your three percent chloride values and find the PPT for the %Cl values. As in Part A, if one trial is quite different from the other two, report data from all three trials, but only average two trials.

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Determination of an Unknown Chloride

YOUR NAME: ____

DATA: *Watch the video (***http://mhchem.org/w/7.htm***) to acquire these values using the data at the very end:*

Part A

NaCl sample #1 (g): ____________________ mL of $AgNO_3$ #1 (mL): ____________________

NaCl sample #2 (g): ____________________ mL of $AgNO_3$ #2 (mL): ____________________

NaCl sample #3 (g): ____________________ mL of $AgNO_3$ #3 (mL): ____________________

Part B

Unknown sample #1 (g): ____________________ mL of $AgNO_3$ #1 (mL): ____________________

Unknown sample #2 (g): ____________________ mL of $AgNO_3$ #2 (mL): ____________________

Unknown sample #3 (g): ____________________ mL of $AgNO_3$ #3 (mL): ____________________

<u>Part A Calculations</u>: *Determination of an Unknown Chloride*

Show all work, use significant figures and circle the final answer for full credit.

Use the space below to calculate three values for the molarity of the $AgNO_3$ solution. The lab video offers help with this calculation. *Show all work! Watch significant figures!*

Using the three calculated values of [$AgNO_3$] from above, calculate the **average concentration** of $AgNO_3$, the **average deviation** of your molarity calculations and the precision in **parts per thousand**. (The parts per thousand handout can be found here: **http://mhchem.org/ppt**)

Part B Calculations: *Determination of an Unknown Chloride*

Show all work, use significant figures and circle the final answer for full credit.

Use the space below to calculate three values for the %Cl of your unknown sample. The lab video offers help with this calculation. *Show all work! Watch significant figures!*

Using the three calculated values of %Cl from above, calculate the **average %Cl**, the **average deviation** of your %Cl calculations and the precision in **parts per thousand**. (ppt handout: http://mhchem.org/ppt)

<u>Part C Postlab Questions</u>: ***Determination of an Unknown Chloride***

Show all work, use significant figures and circle the final answer for full credit.

1. 35.46 mL of a silver nitrate solution was used to reach the chromate end point with a 50 mL solution containing 0.1165 g of pure NaCl. What is the molarity of the $AgNO_3$ solution?

2. How many mL of the silver nitrate solution used in question 1 above will react with 0.2595 g of $BaCl_2$ dissolved in 50 mL of water?

3. A solid chloride sample weighing 0.09969 g required 18.25 mL of 0.05205 M $AgNO_3$ to reach the chromate end point. What is the % chloride in this sample?

<u>Part C Postlab Questions</u>: ***Determination of an Unknown Chloride - continued***
Show all work, use significant figures and circle the final answer for full credit.

4. How would the following hypothetical errors affect the calculated % chloride (increase, decrease or no change)? Explain.

   a. The pure sodium chloride was left open in the scale room and absorbed moisture.

   This will **increase** **decrease** **not change** the percent chloride. ***(circle one)***

   Explain your answer:

   b. The calculated molarity of the silver nitrate solution was 5% too high.

   This will **increase** **decrease** **not change** the percent chloride. ***(circle one)***

   Explain your answer:

   c. Two mL of $AgNO_3$ are added beyond the chromate end in titrating the unknown chloride.

   This will **increase** **decrease** **not change** the percent chloride. ***(circle one)***

   Explain your answer:

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CH 221 Fall 2024:
"Calorimetry" *(online)*
Lab *- Instructions*

Note: **This is the lab for section W1 of CH 221 only.**

- *If you are taking section 01 or section H1 of CH 221*, please use this link:
  http://mhchem.org/s/8a.htm

Step One:

Watch the lab video for the "Calorimetry" lab, found here: **http://mhchem.org/w/8.htm**
Record the data found at the ***end*** of the lab video on page Ib-8-7.

Step Two:

Complete pages Ib-8-7 through Ib-8-14 using the "Calorimetry" video and the actual lab instructions on pages Ib-8-2 through Ib-8-5. Include your name on page Ib-8-7!

Step Three:

Submit your lab (pages Ib-8-7 through Ib-8-14 *only* to avoid a point penalty) **as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, November 20 by 11:59 PM.** I recommend a free program (ex: CamScanner, https://camscanner.com) or a website (ex: CombinePDF, https://combinepdf.com) to convert your work to a PDF file.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Calorimetry

Thermal energy, usually called heat, is one of the most familiar forms of energy. We observe this form of energy when it is passed from an object of higher temperature to one at a lower temperature when the objects are in contact. Heat flow can be measured using a device called a **calorimeter**. A calorimeter is a device that is insulated from the surroundings so that essentially no heat can flow in or out of the device. Within the calorimeter, heat can be transferred from one system to another and this transfer causes a temperature change.

Specific Heat of an Unknown Metal

When heat flows into a substance, the temperature of the substance increases. The quantity of heat, **q**, required to cause a temperature change **ΔT** of any substance varies directly with the mass, **m**, of the substance such that:

$$\mathbf{q = mC\Delta T}$$

The **specific heat, C,** can be considered the amount of heat required to raise the temperature of one gram of a pure substance by one degree Celsius. The calorie is a unit of heat as well as the joule. The **specific heat of water** is 1.000 calorie/g ˚C which equals **4.184 J/g ˚C**. The joule (J) is directly related to mechanical work and is the S.I. unit of energy; hence, we shall use the joule almost exclusively.

The specific heat of a metal can be easily measured with a calorimeter. A weighed amount of metal is heated to a specific temperature and is quickly added to a measured amount of water at a known temperature in a calorimeter. The metal loses heat to the water and eventually the metal and water equilibrate at a temperature above the original temperature of the water but below that of the hot metal. If we assume no heat loss to the surroundings or the walls of the container, the heat lost by the hot metal should equal the heat gain of the water. For heat flow *q*,

$$q_{\text{water}} = -q_{\text{metal}} \quad \text{(the negative sign indicates the opposite flow of the heat)}$$

If we now express the heat flow in terms of specific heat (C) for both the metal as well as the water we get:

$$\mathbf{(C_{water})(m_{water})(\Delta T_{water}) = -(C_{metal})(m_{metal})(\Delta T_{metal})} \qquad \textit{\textbf{equation A}}$$

Knowing the specific heat of water we can use this equation to solve for the specific heat of the metal. The specific heat of a metal is related to its molar mass by a simple relationship. Dulong and Petit discovered that 25 joules is required to raise the temperature of one mole of many metals by 1 ˚C. This relationship, shown below, is known as the **Law of Dulong and Petit**:

$$\textbf{Molar Mass (g / mol)} = \mathbf{25 / C_{metal}}\ \textbf{(J/g ˚C)} \qquad \textit{\textbf{equation B}}$$

In part A of this lab you will determine the specific heat and molar mass of an unknown metal.

Heat of Reaction and Hess's Law

When a physical or chemical change occurs, it is usually accompanied by a change in the heat content (**enthalpy**) of the material in question. Enthalpy (**H**) is defined as the heat content of a given set of conditions, called a **state**. Since there is only one value of enthalpy for any given state, the enthalpy is one of a number of thermodynamic variables called **state functions**. Because many factors internally contribute to the enthalpy of a substance there is no way to measure the enthalpy of a pure substance. Instead we can determine the change in the enthalpy (**ΔH**) when a chemical or physical change occurs. When a chemical reaction occurs in water solutions, the situation is similar to that which is present when a hot metal sample is put into water. As in the specific heat experiment the heat flow for the reaction mixture is equal in magnitude but opposite in sign to that for the water.

$$q_{\text{reaction}} = -q_{\text{water}} \qquad \textit{and} \qquad \Delta H_{\text{reaction}} = q_{\text{reaction}} / \text{mol}$$

By measuring the mass of the water used as solvent, and by observing the temperature change that the water undergoes, we can find q_{water} and therefore $\Delta H_{\text{reaction}}$. An increase in water temperature indicates that heat is given off by the reaction; the reaction is **exothermic**, and $\Delta H_{\text{reaction}}$ is negative. Conversely, if the temperature decreases, heat is absorbed by the reaction from the surroundings, the reaction is **endothermic**, and $\Delta H_{\text{reaction}}$ is positive.

Hess's law further states that when two or more chemical equations are *combined* to produce a balanced chemical equation, the enthalpy changes combined in the same manner will yield the enthalpy change of the new reaction. This will enable us to determine the enthalpy change for a reaction that may not be easily performed in the laboratory, i.e. the enthalpy of formation of acetylene gas (C_2H_2).

The reaction we are trying to determine is:	$2\ C(s) + H_2(g) \rightarrow C_2H_2(g)$	$\Delta H = ?$
By taking 2 x the heat of formation of CO_2:	$2\ C(s) + 2\ O_2(g) \rightarrow 2\ CO_2(g)$	$\Delta H = -787.0$ kJ
$^1/_2$ x the heat of formation of H_2O:	$H_2(g) + {}^1/_2\ O_2(g) \rightarrow H_2O(l)$	$\Delta H = -285.8$ kJ
Reverse the heat of combustion of C_2H_2:	$2\ CO_2(g) + H_2O(l) \rightarrow C_2H_2(g) + {}^5/_2$	$\Delta H = +846.1$ kJ

The **sum** of these enthalpies is **-226.7 kJ**, which is the enthalpy of formation of acetylene.

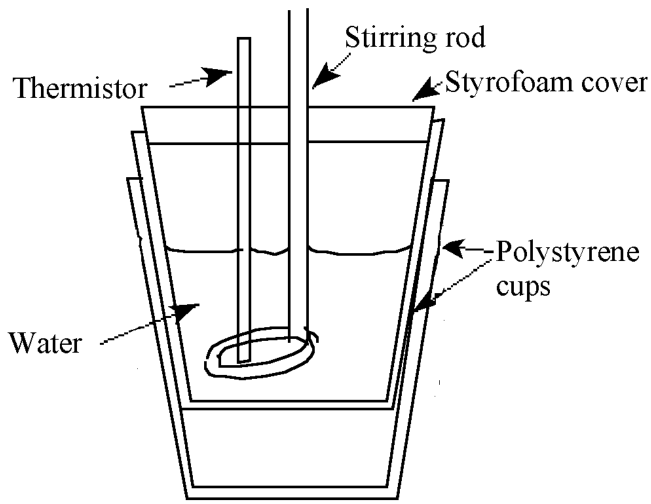
In this experiment we will measure the enthalpy change for the reaction of a metal, zinc, with acid to produce a zinc salt. We will then measure the enthalpy change for zinc oxide reacting with the same acid. From these two reactions along with the value for the reaction of hydrogen with oxygen, one can determine the *heat of combustion of zinc metal* (or the **heat of formation for zinc oxide**):

$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$	ΔH_1 (1)
$ZnO(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$	ΔH_2 (2)
$H_2(g) + {}^1/_2\ O_2(g) \rightarrow H_2O(l)$	$\Delta H_3 = -285.8$ kJ
$Zn(s) + {}^1/_2\ O_2(g) \rightarrow ZnO(s)$	$\Delta H_4 = ?$

As before in the last experiment, you can use a calorimeter to determine the heats (***q***) of reaction and enthalpies for reactions ***1*** and ***2***, above. Combining these enthalpy values with the enthalpy of formation for water (equation ***3***, above), one can use Hess's Law to calculate the heat of formation for Zinc Oxide (equation ***4***.)

PROCEDURE: Part A: SPECIFIC HEAT of an UNKNOWN METAL

1. Assemble the calorimeter as diagrammed. The calorimeter consists of two nested polystyrene coffee cups with a Styrofoam cover. There are two holes in the cover, one for the **thermistor** (which records temperature when connected to a **Vernier LabQuest** apparatus) and one for the glass stirrer provided for this experiment. Weigh the dry calorimeter to 0.001 g. Add about 40 mL of tap water and reweigh the calorimeter and water.



2. Fill a 600 mL beaker 2/3 full with tap water and heat it to boiling. While waiting for the water to boil, weigh a sample of dry metal to the nearest 0.001 g. Return the metal to the dry test tube and clamp the test tube in the boiling water bath so that the metal is below the water line.

3. Record the temperature of the boiling water bath using the LabQuest thermistor probe. Remove the thermistor from the boiling water bath and wipe off all the hot water before placing it in the calorimeter. Record the temperature of the water in the calorimeter.

4. Remove the test tube from the boiling water bath, quickly wipe excess water off the outside of the test tube. Pour the hot metal into the calorimeter without causing the water to splash (tilt the calorimeter). While stirring the water in the calorimeter, monitor the temperature until it remains steady or begins to fall. Record the temperature when it has stabilized.

5. Dry the metal sample, return it to the large test tube and heat it again in the boiling water. **Repeat** the experiment.

Part B: HEAT of REACTION: HESS'S LAW – *all waste should go in a waste container!*

Zinc Reaction

1) Using a graduated cylinder, add about 75.0 mL of 6.00 M HCl in the dry calorimeter. Determine the mass of the HCl solution in the calorimeter to 0.001 g, then record the temperature.
2) Weigh about 0.65 g of Zn to the nearest 0.001 g.
3) Add the metal to the calorimeter, stir and record the highest temperature (when it stabilizes.)

Zinc Oxide Reaction

4) Perform a similar experiment using 75.0 mL of 6.00 M HCl and 1.2 g zinc oxide.

CALCULATIONS for Part A:

1. For each trial, calculate the specific heat of the metal. Use "equation A" on the front page of this lab.
2. Determine the **average** specific heat and deviation in **parts per thousand**.
3. **Estimate the molar mass** using the law of Dulong and Petit (equation B, front page). What is the identity of your metal?

CALCULATIONS for Part B:

1. Calculate the heat change (q) for the Zn and ZnO reactions.
   Example for q_{Zn}: **q_{Zn} = -(Heat capacity of HCl)(g HCl *solution*)(ΔT)** *notice the negative sign!*
   ***Heat capacity** for HCl is 3.86 J/g·˚C. Assume the **density** of the HCl solution is 1.00 g/mL (if needed.)

2. Calculate the heat of reaction (ΔH) for the Zn and ZnO reactions. Watch the sign of your value!
   Example for Zn: **$\Delta H = q_{Zn}$ / mol Zn**

3. Write balanced equations for the two reactions performed in lab, including your experimentally determined ΔH. *Hint:* See the second page of this lab, towards the bottom.

4. Use Hess's Law to determine the heat of formation for zinc oxide: $Zn_{(s)} + \frac{1}{2} O_{2\,(g)} \rightarrow ZnO_{(s)}$ (***Hint**:* See the second page of this lab, towards the bottom! You will need the heat of formation for water to calculate the heat of formation for zinc oxide.)

5. Look up the value for the heat of formation of $ZnO_{(s)}$ in your text. Calculate your **percent error**.
   Percent error = absolute value{ (actual - experimental) / actual }*100%. Remember to explain (in your conclusion) any discrepancies.

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Calorimetry

YOUR NAME: ______________________________

DATA: *Watch the video (***http://mhchem.org/w/8.htm***) to acquire these values using the data at the very end:*

PART A:

<u>Trial 1</u>

metal sample (g): ____________

initial temperature of metal (°C): ____________

water sample (g): ____________

initial temperature of water (°C): ____________

final temperature of metal plus water (°C): ____________

<u>Trial 2</u>

metal sample (g): ____________

initial temperature of metal (°C): ____________

water sample (g): ____________

initial temperature of water (°C): ____________

final temperature of metal plus water (°C): ____________

PART B:

<u>Zinc</u>

mass Zn (g): ____________

Volume HCl (mL): ____________

initial temperature of HCl (°C): ____________

final temperature of solution (°C): ____________

<u>Zinc Oxide</u>

mass ZnO (g): ____________

Volume HCl (mL): ____________

initial temperature of HCl (°C): ____________

final temperature of solution (°C): ____________

<u>Part A Calculations</u>: *Determining the Specific Heat of an Unknown Metal*

Show all work, use significant figures and circle the final answer for full credit.

- Use the data from Part A to calculate the **specific heat of the metal for both trial #1 and trial #2** (two separate calculations.) *Hint:* Use **equation A** on page I-8-2 to solve for C_{metal}. Use this page and the following page to show ***all*** calculations; attach additional pages of work if necessary.
- Find the **average specific heat** of your metal and the **parts per thousand** (http://mhchem.org/ppt).
- Use the **Law of Dulong and Petit** (**equation B** on page I-8-2) to estimate the molar mass of the unknown metal and then predict it's identity using the periodic table.

Part A Calculations (continued, if needed):

Trial #1 Specific Heat of Unknown Metal (J g^{-1} K^{-1}): ____________________________

Trial #2 Specific Heat of Unknown Metal (J g^{-1} K^{-1}): ____________________________

Average Specific Heat of Unknown Metal (J g^{-1} K^{-1}): ____________________________

Parts Per Thousand: _____________________________________

Molar Mass of Unkonwn Metal using the Law of Dulong and Petit (g/mol): ___________

Probable identity of the Unknown Metal (name and symbol): __________________________

Explain briefly if you agree with the probable identity of the metal

Part B Calculations: *Heat of a Reaction / Hess's Law*

For the Zn reaction:

1. Assuming that the density of the HCl solution is **1.00 g/mL**, calculate the mass (g) of the HCl solution for the Zn reaction. *Hint:* this will be a number larger than twenty grams!

2. Calculate the change in temperature (ΔT) for the zinc reaction.

3. Using a *heat capacity* of 3.86 J g^{-1} K^{-1} , find q_{Zn} in **Joules (J)** using the equation:

$$q_{Zn} = -(\textit{heat capacity} \text{ of HCl})(\text{g HCl solution})(\Delta T)$$

4. Calculate the moles of Zn (**mol Zn**) used using the grams of Zinc used and the molar mass of Zn.

5. Find the heat of reaction for the Zn reaction (**ΔH_{Zn}**) ***in kJ/mol*** using the equation:

$$\Delta H_{Zn} = q_{Zn} / \text{mol Zn}$$

For the ZnO reaction:

1. Assuming that the density of the HCl solution is **1.00 g/mL**, calculate the mass (g) of the HCl solution for the ZnO reaction. *Hint:* this will be a number larger than twenty grams!

2. Calculate the change in temperature (ΔT) for the zinc oxide reaction.

3. Using a *heat capacity* of 3.86 J g^{-1} K^{-1} , find q_{ZnO} in **Joules (J)** using the equation:

$$q_{ZnO} = -(\textit{heat capacity} \text{ of HCl})(\text{g HCl solution})(\Delta T)$$

4. Calculate the molar mass for zinc oxide (g/mol).

5. Calculate the moles of ZnO (**mol ZnO**) used using the grams of zinc oxide used and the molar mass of zinc oxide.

6. Find the heat of reaction for the ZnO reaction (**ΔH_{ZnO}**) ***in kJ/mol*** using the equation:

$$\Delta H_{ZnO} = q_{ZnO} / \text{mol ZnO}$$

Hess's Law:

1. Use your previously calculated values of $\boldsymbol{\Delta H_{Zn}}$ and $\boldsymbol{\Delta H_{ZnO}}$ to complete the missing values in the equations below:

$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ ΔH_{Zn} (kJ/mol) = ___________________

$ZnO(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$ ΔH_{ZnO} (kJ/mol) = ___________________

$H_2(g) + {}^1/_2\ O_2(g) \rightarrow H_2O(l)$ ΔH_3 (kJ/mol) = -285.8 kJ

2. Use Hess's Law and the three equations above (ΔH_{Zn}, ΔH_{ZnO}, ΔH_3) to find ΔH_f for the following equation:

$Zn(s) + {}^1/_2\ O_2(g) \rightarrow ZnO(s)$ ΔH_f (kJ/mol) = ___________________

3. Look up the value of ΔH_f for ZnO in your textbook (or here: **https://mhchem.org/thermo**). Calculate the percent error (% error) using the *actual* value (from the textbook or table) and your *experimental* value (answer #2 above.) Why are the two values not equal? Explain briefly.

Percent error = absolute value{ (actual - experimental)/ actual }*100%

Postlab Questions:

Show all work, use significant figures and circle the final answer for full credit.

1. A metal sample weighing 45.23 g is heated to 100.00 ˚C. It is placed in 38.64 g of water in a calorimeter at 25.25 ˚C. At equilibrium, the temperature of the water and metal is 33.05 ˚C. Calculate the specific heat of the metal.

2. When 0.5000 g of Zn are added to 50.00 mL of 6.00 M HCl, the temperature of the solution rises from 20.00 ˚C to 32.56 ˚C. Using the assumptions and procedures recommended in Part B of this lab concerning the HCl density and specific heat, calculate $\boldsymbol{q_{Zn}}$ in **Joules (J)** and ΔH_{Zn} in **kilojoules/mol (kJ/mol)** for the reaction.

<u>Postlab Questions</u>: ***Continued***

3. Calculate the enthalpy change for the allotropic transformation of graphite into diamond using the following data:

$$C\ (graphite) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -390.5\ kJ$$

$$C\ (diamond) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -393.5\ kJ$$

4. Using the following equations:

$$C(s) + 2\ Cl_2(g) \rightarrow CCl_4(g) \qquad \Delta H^0 = -135.4\ kJ$$
$$H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g) \qquad \Delta H^0 = -184.6\ kJ$$
$$2\ H_2(g) + C(s) \rightarrow CH_4(g) \qquad \Delta H^0 = -74.8\ kJ$$

Calculate the standard enthalpy of reaction for the process:

$$CH_4(g) + 4\ Cl_2(g) \rightarrow CCl_4(g) + 4\ HCl(g) \qquad \Delta H^0 = ?$$

5. Using the Law of Dulong and Petit, calculate the heat capacity of pure bismuth.

CH 221 Fall 2024:
“Hydrogen Spectrum”
(online) ***Lab*** *- Instructions*

Note: **This is the lab for section W1 of CH 221 only.**

- *If you are taking section 01 or section H1 of CH 221*, please use this link: http://mhchem.org/s/9a.htm

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Step One:

Watch the lab video for the “Hydrogen Spectrum” lab, found here:
http://mhchem.org/w/9.htm
Record the data found at the ***end*** of the lab video on page Ib-9-4.

Step Two:

Complete pages Ib-9-5 through Ib-9-7 using the “Hydrogen Spectrum” video and the actual lab instructions on pages Ib-9-2 through Ib-9-3. Include your name on page Ib-9-5!

Step Three:

Submit your lab (pages Ib-9-5 through Ib-9-7 *only* to avoid a point penalty**) as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, December 4 by 11:59 PM.** I recommend a free program (ex: CamScanner, https://camscanner.com) or a website (ex: CombinePDF, https://combinepdf.com) to convert your work to a PDF file.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

The Atomic Spectrum of Hydrogen

When atoms are excited, either in an electric discharge or with heat, they tend to give off light. The light is emitted only at certain wavelengths that are characteristic of the atoms in the sample. These wavelengths constitute what is called the atomic spectrum of the excited element and reveal much of the detailed information we have regarding the electronic structure of atoms.

Atomic spectra are interpreted in terms of quantum theory, which states that atoms can exist only at certain states that correspond to fixed energy levels. When an atom changes its state, it must absorb or emit an amount of energy that is just equal to the difference between the energies of the initial and final states. This energy may be absorbed or emitted in the form of light. The emission spectrum of an atom is obtained when excited atoms fall from a higher to a lower energy level. Since there are many such levels, the atomic spectra of most elements are very complex.

Light is absorbed or emitted by atoms in the form of photons, each with a specific amount of **energy**, **E**. This energy is related to the **frequency (ν)** and **wavelength** (λ) of light by the following equation:

$$E_{photon} = h\nu = \frac{hc}{\lambda}$$

where **h** = **Planck's constant** = 6.62608 x 10^{-34} J·s *and* **c** = **speed of light** = 2.997925 x 10^{8} m / s

The law of conservation of energy states that total energy is conserved. Thus, the change in energy of the atom must equal the energy change of the photon emitted. The energy change of the atom is equal to the energy of the upper energy level minus the energy of the lower level.

$$\Delta E_{atom} = (E_{final} - E_{initial}) = E_{photon} = \frac{hc}{\lambda}$$

The amount of energy in a photon given off when an atom changes from one level to another is very small, of the order of 10^{-19} joules. To avoid working with such small numbers, we will work with one mole of atoms. The above equation is multiplied by **Avogadro's number**, **N**:

$$\Delta E\ (\text{kJ/mol}) = \frac{Nhc}{\lambda} = \frac{(6.02214 \text{ x } 10^{23})(6.62608 \text{ x } 10^{-34}\,\text{J}\cdot\text{sec})(2.997925 \text{ x } 10^{8}\,\text{m/sec})}{\lambda} = \frac{1.19627 \text{ x } 10^{5}}{\lambda\ (\text{nm})}$$

The above equation is useful in the interpretation of atomic spectra. For example, in the study of the atomic spectrum of sodium, a strong yellow line of wavelength 589.16 nm is observed. The above equation can be used to determine the change in energy. This in turn is corresponded to the transition in energy levels.

The simplest of all atomic spectra is that of the hydrogen atom. In 1886 Balmer showed that the lines in the spectrum of the hydrogen atom had wavelengths that could be expressed by a rather simple equation. In 1913, Bohr explained the spectrum on a theoretical basis with his model of the hydrogen atom. According to Bohr's theory, the energies allowed to a hydrogen atom are given by the so-called **Bohr's Equation**:

$$E_n = \frac{-B}{n^2} = \frac{-1312.04}{n^2}$$

where **B** = a constant (1312.04 kJ/mol) and **n** = the quantum number (1, 2, 3, ...)
Bohr's equation allows you to calculate quite accurately the energy levels for hydrogen. Transitions between these levels give rise to the wavelengths in the atomic spectrum of hydrogen. These wavelengths are also known very accurately.

Given both the energy levels and the wavelengths, it is possible to determine the actual levels associated with each wavelength. In this experiment, your task will be to measure the wavelengths of the hydrogen spectrum and then determine the transition in energy levels associated with each wavelength. This lab consists of a worksheet that needs to be completed for credit; no formal typed laboratory report is due this week.

PROCEDURE: Part A: Visual Observation of a Hydrogen Discharge Tube Using a Spectroscope

The instructor will set up a spectroscope with a hydrogen discharge tube. Note the color of the emitted light without using the spectroscope.

Now view the hydrogen tube using the spectroscope. How many lines do you see? What color does each line have?

Now we shall use the Vernier system and an emission spectrometer to determine the wavelengths of the hydrogen lines you viewed in the spectroscope.

Part B: The Emission Spectrum of Hydrogen Using the Vernier LabQuest 2

Assemble the emission spectrometer and Lab Quest 2 per your instructor's instructions. In the Lab Quest program, you should see "USB: Intensity rel" if everything is connected correctly.

Place the close to (but not touching!) the middle of a hydrogen discharge tube using a LabJack. Start the data collection by pushing the green "start" button (a green triangle) in the lower left of the Vernier LabQuest 2. **You should see at least three distinct peaks to perform the analysis**: the main central peak (656.3 nm) and at least two to the left of the main 656.3 nm peak (at 486.1 and 434.1 nm.) Ignore any peaks to the right (i.e. larger than 700 nm) – these are due to impurities in the hydrogen tube Once you have at least three peaks visible, **stop** your experiment and turn off the hydrogen discharge tube.

Use the LabQuest 2 tools to find the wavelengths of the four emission peaks for hydrogen. Once the experiment is stopped, go "Graph – Graph Options", then change the following: **Left** to 400, **Right** to 700 and **Top** to 0.400, then press **OK**. Your data points will be visible, and a fourth point should be apparent at about 410 nm.

Determine the four experimental emission wavelengths using the "left" and "right" buttons (which control the cursor; alternatively, use the pointer to select the exact point.) **Record** these values and **compare** them to the theoretical wavelengths for hydrogen to calculate the **percent error** for each line on the next page.

Part C: Calculations for the Energy Levels of Hydrogen Atom

Next, you can use the hydrogen wavelengths to calculate the **energy change** for each line in the observed hydrogen spectrum. Using Bohr's equation, calculate the **energy levels (ε_n)** in kJ/mole for each of the eight lowest allowed levels of the hydrogen atom starting with n=1 to n=8. Note that all the energies are negative, so that the lowest energy will have the largest allowed negative value.

The energy levels will allow you to determine the **energy transition (ΔE)** that corresponds to the observed wavelengths. Determine the **quantum numbers** for the initial (n_{hi}) and final (n_{low}) states for these transitions.

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The Atomic Spectrum of Hydrogen: *Worksheet* **Name:**

Complete the following worksheets. *Use this form!*

Part A: *Visual Observations*
What **color** was the hydrogen discharge tube when you looked directly at it? ________________________

What four colors did you observe through the spectroscope? ______________________________________

Part B: *The Emission Spectrum of Hydrogen* – Record the hydrogen wavelengths at the end of the video.

Color	λ (Vernier, nm)	λ (theoretical, nm)	Percent Error
Red	____________	_656.3_	___________
Blue	____________	_486.1_	___________
Violet	____________	_434.1_	___________
Violet	____________	_410.2_	___________

Recall:

$$\text{Percent Error} = \frac{\text{absolute value (theoretical value - experimental value)}}{\text{theoretical value}} * 100\%$$

Part C: Calculations for the Energy Levels of Hydrogen Atom
Find the energy level ε_n (in kJ/mol) for each quantum number from 1 through 8 using the following equation:

$$\varepsilon_n = \frac{-1312.04 \text{ (kJ/mol)}}{n^2}$$

Example: The n=1 energy level can be calculated as follows: $\varepsilon_n = (-1312.04/1^2)$ = **-1312.04 kJ/mol**
Example: The n=2 energy level can be calculated as follows: $\varepsilon_n = (-1312.04/2^2)$ = **-328.010 kJ/mol**

n value	ε_n (kJ/mol)	n value	ε_n (kJ/mol)
1	-1312.04	5	______
2	-328.010	6	______
3	______	7	______
4	______	8	______

Now complete the table below. The first row has been partially completed for you.

Color	**λ (actual, nm)**	**ν (s^{-1})**	**ΔE (J/photon)**	**ΔE (kJ/mol)**	**n_{hi}**	**n_{low}**
red	**656.3**	**4.568×10^{14}**	**3.027×10^{-19}**	**182.3**	_____	_____
blue	**486.1**	________	________	______________	_____	_____
violet	**434.1**	________	________	______________	_____	_____
violet	**410.2**	________	________	______________	_____	_____

Convert the wavelength values into ΔE (in kJ/mol) using the following equation:

$$\Delta E\ (\text{kJ/mole}) = \frac{1.19627 \times 10^5}{\lambda\ (\text{nm})}$$

The calculated ΔE values correspond to a transition between the various energy levels, ε_n, calculated previously. Determine which transition they correspond to by finding the change in energy (i.e. ΔE) between levels.

Example: Find the change in energy in a transition of hydrogen between the n=2 and n=1 energy levels.

The energy level, ε_n, for n=2 is **–328.010** kJ/mol, and the energy level, ε_n, for n=1 is **–1312.04** kJ/mol. A change in energy, **ΔE**, corresponds to the final energy state minus the initial energy state, or:

$\Delta E = \varepsilon_{final} - \varepsilon_{initial} = \varepsilon_1 - \varepsilon_2 = -1312.04 - (-328.010) =$ **-984.03 kJ/mol**

If your calculated value of ΔE is about –984.03 kJ/mol, then your n_{hi} would be **2** (the higher value of n) and your n_{low} value would be **1** (the lower value of n).

Show the frequency calculation for the red line below:

Show the ΔE (J/photon) calculation for the red line below:

Show the ΔE (kJ/mol) calculation for the red line below:

<u>Post Lab Questions</u>:

1. When Balmer found his famous series for hydrogen in 1886, he was limited experimentally to wavelengths in the visible and near ultraviolet regions from 250 nm to 700 nm, as in your experiment. What common characteristic do the lines in the Balmer series have?

2. In the hydrogen atom, the electron is in its lowest energy state, n=1. The maximum electron energy that a hydrogen atom can have is 0 kJ/mole, at which point the electron would essentially be removed from the atom and it would become a H^+ ion. How much energy does it take to ionize one hydrogen atom in **kilojoules per mole** *and* in **Joules per atom**? (*Hint:* calculate ΔE where ε_{final} is zero and $\varepsilon_{initial}$ is –1312.04 kJ/mol.)

<u>Questions #3 through #5</u> will use the equation below for the helium ion. The **helium ion**, He^+, has energy levels similar to those of the hydrogen atom, since both species have only one electron. The energy levels of the helium ion are given by the following equation:

$$E_n = \frac{-5248.16}{n^2} \text{ kJ/mol where } n = 1, 2, 3...$$

3. Calculate the energies in kJ/mole for the four lowest energy levels of the helium ion using the equation above.

ε_1 __________ ε_3 __________

ε_2 __________ ε_4 __________

4. One of the most important transitions for the helium ion involves a jump from the n = 2 to the n = 1 level. Calculate the **change in energy** in kJ/mole for this transition. (*Hint:* $\Delta E = \varepsilon_1 - \varepsilon_2$). Use the equation found in Part C of the worksheet to calculate the **wavelength** (in nm) of this transition.

5. Three of the strongest lines in the helium ion spectrum are observed at the following wavelengths. Find the quantum numbers of the initial and final energy states for the transitions that give rise to these three lines:

<u>λ</u>	<u>ΔE (kJ/mol)</u>	$\underline{n}_{hi}$	$\underline{n}_{low}$
121.57 nm			
164.12 nm			
468.90 nm			

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CH 221 Fall 2024:

Problem Set #1

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #1 by **reading Chapter 1** of the textbook and/or by watching the videos found on the website (https://mhchem.org/221)
- **Try the problems** for Problem Set #1 found on the next pages on your own first. **Write out the answers (and show your work) by hand (on a tablet or paper)**; do not type your answers (and work) to avoid a point penalty. If you write the answers on the problem set itself, you will receive fewer points. Include your name on your problem set!

Step Two:

Section 01 and H1: We will go over Problem Set #1 during recitation. ***Self correct all problems*** of your problem set before turning it in at the end of recitation.

- *Section 01:* due **Monday, September 30 at 1:10 PM**
- *Section H1:* due **Wednesday, October 2 at 1:10 PM**

Section W1: **Watch the recitation video** for Problem Set #1:
http://mhchem.org/w/a.htm

- **Self correct *all* of the problems** while viewing the video. Mark correct problems with a star (or other similar mark), and correct all incorrect problems (show the correct answer and the steps required to achieve it.)
- **Submit Problem Set #1 via email (mike.russell@mhcc.edu) as a single PDF file** (use CamScanner (https://camscanner.com), CombinePDF (https://combinepdf.com), etc.) **by 11:59 PM Wednesday, October 2.**

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

CH 221 Problem Set #1

** Complete problem set on separate pieces of paper showing all work, circling final answers, etc.*
** Self correct your work before turning it in to the instructor.*

Covering: **Chapter One** *and* **Chapter Guide One**
Important Tables and/or Constants: 1 cm^3 = 1 mL, **273.15** *(temperature); know these metric prefixes:* **nano** (10^{-9}), **micro** (10^{-6}), **milli** (10^{-3}), **centi** (10^{-2}) and **kilo** (10^3).

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1. Give the name of each of the following elements: Mn, Cu, Na, K, Xe, Fe
2. Give the symbol for each of the following elements: silver, fluorine, plutonium, tin, technetium, krypton
3. In each of the following pairs, decide which is an element and which is a compound:
    a. $Pt(NH_3)_2Cl_2$ and Pt
    b. copper and copper(II) oxide
    c. silicon and silane
4. A piece of silver metal has a mass of 2.365 g. If the density of silver is 10.5 g/cm^3, what is the volume of the silver?
5. Make the following temperature conversions:
    a. 77 K to °C
    b. 63 °C to K
    c. 1450 K to °C
    d. 67.6 °F to °C *Helpful equation:* °F = 1.8*(°C) + 32
6. A compact disc has a diameter of 11.8 cm. Calculate the surface area of one side of the disc in square centimeters and square meters. (area of a circle = πr^2 where r = radius; ignore the center hole.)
7. The separation between carbon atoms in diamond is 0.154 nm. Express this distance in meters and picometers.
8. The solder once used by plumbers to fasten copper pipes together consists of 67% lead and 33% tin by mass. What is the mass of lead in a 250 g block of solder?
9. You have a white crystalline solid known to be one of the potassium compounds listed below. To determine which, you measure the solid's density. You measure out 18.82 g and transfer it to a graduated cylinder containing kerosene (which doesn't dissolve the salts.) The liquid level rises from 8.50 mL to 15.30 mL. Calculate the density of the solid and identify the compound from the following list:
    a. KF, density = 2.48 g/cm^3
    b. KCl, density = 1.98 g/cm^3
    c. KBr, density = 2.75 g/cm^3
    d. KI, density = 3.13 g/cm^3
10. Four balloons are each filled with a different gas. If the density of air is 1.12 g/L, which balloon or balloons float in the air?
    a. He, density = 0.164 g/L
    b. Ne, density = 0.825 g/L
    c. Ar, density = 1.633 g/L
    d. Kr, density = 4.425 g/L

Problem Set #1 continues on the next page

11. Give the number of significant figures in each of the following numbers:
    a. 0.00546 g
    b. 1600 mL
    c. 2.300 x 10^{-4} g
    d. 2.34 x 10^{9} atoms
    e. 400. km
12. Carry out the following calculation and report the answer to the correct number of significant figures. $$(1.68)\left[\frac{23.56-2.3}{1.248 \text{ x } 10^{3}}\right]$$
13. Copper has a density of 8.96 g/cm^3. An ingot of copper with a mass of 57 kg (126 lb) is drawn into a wire with a diameter of 9.50 mm. What length of wire (in meters) can be produced? [Volume of the wire = πr^2(length)]
14. When you heat popcorn, it pops because it loses water explosively. Assume a kernel of corn with a mass of 0.125 g has a mass of only 0.106 g after popping.
    a. What percentage of its mass did the kernel lose on popping?
    b. Popcorn is sold by the pound in the United States. Using 0.125 g as the average mass of a popcorn kernel, how many kernels are there in a pound of popcorn? [*helpful conversion:*1 lb = 453.6 g]
15. The fluoridation of city water supplies has been practiced in the United States in many major cities for several decades. It is accomplished by continuously adding sodium fluoride to water as it comes from a reservoir. Assume you live in a medium-sized city of 150,000 people and that 660 L (170 gal) of water is consumed per person per day. What mass of sodium fluoride (in kilograms) must be added to the water supply each year (365 days) to have the required fluoride concentration of 1 ppm (part per million) - that is, 1 kilogram of fluoride per 1 million kilograms of water? (Sodium fluoride is 45.0% fluoride, and water has a density of 1.00 g/cm^3.)
16. Automobile batteries are filled with an aqueous solution of sulfuric acid. What is the mass of acid (in grams) in 500. mL of the battery acid solution if the density of the solution is 1.285 g/cm^3 and if the solution is 38.08% sulfuric acid by mass?

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CH 221 Fall 2024:

Problem Set #2

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #2 by **reading Chapter 2 and Chapter 3 (3.1 especially)** of the textbook and/or by watching the videos found on our website (https://mhchem.org/221)
- **Try the problems** for Problem Set #2 found on the next pages on your own first. **Write out the answers (and show your work) by hand (on a tablet or paper)**; do not type your answers (and work) to avoid a point penalty. If you write the answers on the problem set itself, you will receive fewer points. Include your name on your problem set!

Step Two:

Section 01 and H1: We will go over Problem Set #2 during recitation. ***Self correct all problems*** of your problem set before turning it in at the end of recitation.

- *Section 01:* due **Monday, October 7 at 1:10 PM**
- *Section H1:* due **Wednesday, October 9 at 1:10 PM**

Section W1: **Watch the recitation video** for Problem Set #2:

http://mhchem.org/w/n.htm

- **Self correct *all* of the problems** while viewing the video. Mark correct problems with a star (or other similar mark), and correct all incorrect problems (show the correct answer and the steps required to achieve it.)
- **Submit Problem Set #2 via email (mike.russell@mhcc.edu) as a single PDF file** (use CamScanner (https://camscanner.com), CombinePDF (https://combinepdf.com), etc.) **by 11:59 PM Wednesday, October 9.**

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

CH 221 Problem Set #2

** Complete problem set on separate pieces of paper showing all work, circling final answers, etc.*
** Self correct your work before turning it in to the instructor.*

Covering: **Chapter Two, Chapter 3.1** *and* **Chapter Guide Two**
Important Tables and/or Constants: 1 mol = 6.022 x 10^{23}, periodic table (http://mhchem.org/pertab)

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1. Give the mass number of:
   a. a nickel atom with 31 neutrons
   b. a plutonium atom with 150 neutrons, and
   c. a tungsten atom with 110 neutrons
2. Give the complete symbol ($^{A}_{Z}X$) for each of the following atoms:
   a. fluorine with 10 neutrons
   b. chromium with 28 neutrons
   c. xenon with 78 neutrons
3. Strontium has four stable isotopes. Strontium-84 has a very low natural abundance, but ^{86}Sr, ^{87}Sr and ^{88}Sr are all reasonably abundant. Knowing that the atomic weight of strontium is 87.62, which of the more abundant isotopes predominates?
4. Copper exists as two isotopes: ^{63}Cu (62.9298u) and ^{65}Cu (64.9278u). What is the approximate percentage of ^{65}Cu in samples of the element?
   a. 10%
   b. 30%
   c. 50%
   d. 70%
   e. 90%
5. Antimony has two stable isotopes, ^{121}Sb and ^{123}Sb, with masses of 120.9038u and 122.9042u, respectively. Calculate the percent abundances of these isotopes of antimony.
6. Calculate the mass in grams of:
   a. 4.24 mol of gold
   b. 15.6 mol of He
   c. 0.063 mol of platinum
   d. 3.63 x 10^{-4} mol of Pu
7. Calculate the amount (moles) represented by each of the following:
   a. 16.0 g of Na
   b. 0.876 g of tin
   c. 0.0034 g of platinum
   d. 0.983 g of Xe
8. Here are the symbols for five of the seven elements whose names begin with the letter B: **B, Ba, Bk, Bi** and **Br**. Match each symbol with one of the descriptions below:
   a. a radioactive element
   b. a liquid at room temperature
   c. a metalloid
   d. an alkaline earth element
   e. a Group 5A element

Problem Set #2 continues on next page

9. Fill in the blanks in the table (one column per element):

Symbol	**^{65}Cu**	**^{86}Kr**		
Number of protons			**78**	
Number of neutrons			**117**	**46**
Number of electrons in the neutral atom				**35**
Name of the element				

10. The recommended daily allowance (RDA) of iron in your diet is 15 mg. How many moles is this? How many atoms?
11. In an experiment, you need 0.125 mol of sodium metal. Sodium can be cut easily with a knife, so if you cut out a block of sodium, what should the volume of the block be in cubic centimeters? If you cut a perfect cube, what is the length of the edge of a cube? (The density of sodium metal is 0.971 g/cm^3.)
12. An object is coated with a layer of chromium 0.015 cm thick. The object has a surface area of 15.2 cm^2. How many atoms of chromium are used in the coating? (The density of chromium = 7.19 g/cm^3.)
13. Consider at atom of ^{64}Zn:
    a. Calculate the density of the nucleus in g/cm^3 knowing that the nuclear radius is 4.8 x 10^{-6} nm and the mass of the ^{64}Zn atom is 1.06 x 10^{-22} g. [Recall that the volume of a sphere = $^4/_3\pi r^3$]
    b. Calculate the density (in g/cm^3) of the space occupied by the electrons in the zinc atom, given that the atomic radius is 0.125 nm and the mass of a single electron is 9.11 x 10^{-28} g. Assume the zinc atom is neutral.
    c. Having calculated these densities, what statement can you make about the relative densities of the parts of the atom?
14. Match the name on the left with the description on the right.

a. Democritus	1. ___ Discovered the neutron
b. Aristotle	2. ___ The oil drop experiment for electron charge
c. Dalton	3. ___ Proposed a value for the mole
d. Becquerel	4. ___ Observed radioactivity on photographic plates
e. Curie (Marie)	5. ___ "The world is made of fire, earth, water and air"
f. Avogadro	6. ___ Discovered the nucleus is very dense
g. JJ Thomson	7. ___ Plum pudding model for the atom
h. Millikan	8. ___ Discovered types of radiation, 2 Nobel Prizes
i. Rutherford	9. ___ Matter made of atoms, proposed atomic mass scale
j. Chadwick	10 ___ First to propose the concept of the atom
k. alpha	11 ___ Radioactive negative electron
l. beta	12 ___ Electromagnetic radiation, pure energy, massless
m. gamma	13 ___ Radioactive positive helium nucleus

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CH 221 Fall 2024:

Problem Set #3

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #3 by **reading Chapter 2 and Chapter 3 (3.1 - 3.3 especially)** of the textbook and/or by watching the videos found on our website (https://mhchem.org/221)
- **Try the problems** for Problem Set #3 found on the next pages on your own first. **Write out the answers (and show your work) by hand (on a tablet or paper)**; do not type your answers (and work) to avoid a point penalty. If you write the answers on the problem set itself, you will receive fewer points. Include your name on your problem set!

Step Two:

Section 01 and H1: We will go over Problem Set #3 during recitation. ***Self correct all problems*** of your problem set before turning it in at the end of recitation.

- *Section 01:* due **Monday, October 14 at 1:10 PM**
- *Section H1:* due **Wednesday, October 16 at 1:10 PM**

Section W1: **Watch the recitation video** for Problem Set #3:
http://mhchem.org/w/r.htm

- **Self correct *all* of the problems** while viewing the video. Mark correct problems with a star (or other similar mark), and correct all incorrect problems (show the correct answer and the steps required to achieve it.)
- **Submit Problem Set #3 via email (mike.russell@mhcc.edu) as a single PDF file** (use CamScanner (https://camscanner.com), CombinePDF (https://combinepdf.com), etc.) **by 11:59 PM Wednesday, October 16.**

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

CH 221 Problem Set #3

** Complete problem set on separate pieces of paper showing all work, circling final answers, etc.*
** Self correct your work before turning it in to the instructor.*

Covering: **Chapter Two, Chapter 3.1-3.2** *and* **Chapter Guide Three**
Important Tables and/or Constants: 1 mol = 6.022×10^{23}, "Have No Fear Of Ice Clear Brew" (7 Diatomics)

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1. What charges are most commonly observed for monatomic ions of the following elements: **selenium**, **potassium**, and **iron.** What are the seven diatomics?
2. Give the symbol, including the correct charge, for each of the following ions:
   a. permanganate ion
   b. nitrite ion
   c. phosphate ion
   d. dihydrogen phosphate ion
   e. ammonium ion
   f. sulfite ion
3. For each of the following compounds, give the formula, charge and the number of each ion that makes up the compound:
   a. $Mg(CH_3CO_2)_2$
   b. $Al(OH)_3$
   c. $CuCO_3$
   d. $Ti(SO_4)_2$
   e. KH_2PO_4
4. Platinum is a transition element and forms Pt^{2+} and Pt^{4+} ions. Write the formulas for the compounds of each of these ions with a) chloride ions and b) sulfide ions.
5. Which of the following are correct formulas for ionic compounds? For those that are not, give the correct formula.
   a. Ca_2O
   b. $SrBr_2$
   c. Li_2O
   d. Fe_2O_5
6. Name each of the following ionic compounds:
   a. $Ca(CH_3CO_2)_2$
   b. $Al(OH)_3$
   c. KH_2PO_4
   d. $Ni_3(PO_4)_2$
7. Give the formula for each of the following ionic compounds:
   a. calcium hydrogen carbonate
   b. potassium permanganate
   c. magnesium perchlorate
   d. potassium hydrogen phosphate
   e. sodium sulfite
8. Consider the two ionic compounds NaCl and CaO. In which compound are the cation-anion attractive forces stronger? Explain your answer.

Problem Set #3 continues on next page

9. Name each of the following binary nonionic compounds containing nitrogen and oxygen:
    a. N_2O_5
    b. NO_2
    c. N_2O_4
    d. N_2O
10. Give the formula for each of the following compounds:
    a. bromine trifluoride
    b. xenon difluoride
    c. diphosphorus tetrafluoride
    d. ammonia
    e. hydrazine
11. Calculate the molar mass to 0.01 g/mol for each of the following compounds:
    a. $Fe(C_6H_{11}O_7)_2$, iron(II) gluconate, a dietary supplement
    b. $CH_3CH_2CH_2CH_2SH$, butanethiol, has a skunk-like odor
    c. $C_{20}H_{24}N_2O_2$, quinine, used as an antimalarial drug
12. Assume you have 0.123 mol of each of the following compounds.
    a. What mass in grams would you have if the compound is $C_{14}H_{10}O_4$, benzoyl peroxide, which is used in acne medications?
    b. How many molecules of $Pt(NH_3)_2Cl_2$ (cisplatin, a cancer chemotherapy drug) would you have?
13. Calculate the weight percent of iron in Fe_2O_3, iron(III) oxide. What mass of iron (in grams) is present in 25.0 g of Fe_2O_3?
14. Nitrogen and oxygen form a series of oxides with the general formula N_xO_y. One of them, a blue solid, contains 36.84% N. What is the empirical formula of this oxide?
15. An organic compound has the empirical formula C_2H_4NO. If its molar mass is 116.1 g/mol, what is the molecular formula of the compound?
16. You combine 1.25 g of germanium with excess chlorine (Cl_2). The mass of product, Ge_xCl_y, is 3.69 g. What is the empirical formula of the product?
17. If Epsom salt, $MgSO_4 \cdot x\ H_2O$, is heated to 250 °C, all the water of hydration is lost. On heating a 1.687 g sample of the hydrate, 0.824 g of $MgSO_4$ remains. How many molecules of water occur per formula unit of $MgSO_4$? Name the hydrated compound.
18. Fluorocarbonyl hypofluorite is composed of 14.6% C, 39.0% O and 46.3% F. If the molar mass of the compound is 82 g/mol, determine the empirical and molecular formulas for the compound.

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CH 221 Exam Prep I Worksheet Name: ______________________________

Worksheet due dates: **<u>Mon, 10/21</u>, 1:10 PM** *(01)* **, <u>Wed, 10/23</u>, 1:10 PM** *(H1)* ***or*** **11:59 PM** *(W1, email)*. To complete, show ***detailed*** *steps* on how to get the given answer for each problem. *Failure to use this form for work and answers will result in a point penalty.*

<u>Problem 1</u>: Lithium has two stable isotopes with masses of 6.0152 amu and 7.0160 amu. The average molar mass of Li is 6.9410 amu. What is the percent abundance of each isotope?

Answer to Problem #1: **7.49% ^{6}Li and 92.51% ^{7}Li**

<u>Problem 2</u>: A given sample of xenon fluoride contains molecules of a single type XeF_n, where *n* is a whole number. If $9.35 * 10^{20}$ molecules of XeF_n weigh 0.322 g, calculate the most likely value of *n*.

Answer to Problem #2: ***n*** **= 4**

<u>Problem 3</u>: Complete the following problems using correct significant figures:

20.42 + 1.322 + 83.1 = ____________ 15.5 x 27.3 x 5.4 = ____________ 320.5 – 6104.5/2.3 = ____________

Problem 4: A nail is coated with a 0.042 cm thick layer of zinc. The surface area of the nail is 9.17 cm^2. The density of zinc is approximately 7.13 g/cm^3. How many zinc atoms are used in the coating?

Answer to Problem #4: **$2.5 * 10^{22}$ atoms**

Problem 5: In a chemical reaction, 1.000 g of sulfur combines with 3.963 g of copper to give a pure compound. What is the empirical formula for this compound?

Answer to Problem #5: **Cu_2S**

Problem 6: Cyclooctene is a hydrocarbon containing only C and H atoms. When burned in oxygen, 1.000 g of cyclooctene produces 3.195 g of CO_2 and 1.144 g of water. Mass spectrometry shows a molar mass value of 110.2 g/mol. What is the empirical and molecular formula of cyclooctene?

Answer to Problem #6: **C_4H_7 and C_8H_{14}**

CH 221 Fall 2024:

Problem Set #4

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #4 by **reading Chapter 4** of the textbook and/or by watching the videos found on our website (https://mhchem.org/221)
- **Try the problems** for Problem Set #4 found on the next pages on your own first. **Write out the answers (and show your work) by hand (on a tablet or paper)**; do not type your answers (and work) to avoid a point penalty. If you write the answers on the problem set itself, you will receive fewer points. Include your name on your problem set!

Step Two:

Section 01 and H1: We will go over Problem Set #4 during recitation. ***Self correct all problems*** of your problem set before turning it in at the end of recitation.

- *Section 01:* due **Monday, October 28 at 1:10 PM**
- *Section H1:* due **Wednesday, October 30 at 1:10 PM**

Section W1: **Watch the recitation video** for Problem Set #4:
http://mhchem.org/w/m.htm

- **Self correct *all* of the problems** while viewing the video. Mark correct problems with a star (or other similar mark), and correct all incorrect problems (show the correct answer and the steps required to achieve it.)
- **Submit Problem Set #4 via email (mike.russell@mhcc.edu) as a single PDF file** (use CamScanner (https://camscanner.com), CombinePDF (https://combinepdf.com), etc.) **by 11:59 PM Wednesday, October 30.**

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

CH 221 Problem Set #4

** Complete problem set on separate pieces of paper showing all work, circling final answers, etc.*
** Self correct your work before turning it in to the instructor.*

Covering: **Chapter Four** *and* **Chapter Guide Four**
Important Tables and/or Constants: **Solubility Table** (in the "Net Ionics" lab or here: **https://mhchem.org/sol**) - *Use the Net Ionics solubility table when answering questions about solubility in CH 221*)

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1. Balance the following equations:
   a. $Cr(s) + Cl_2(g) \rightarrow CrCl_3(s)$
   b. $SiO_2(s) + C(s) \rightarrow Si(s) + CO(g)$
   c. $Fe(s) + H_2O(g) \rightarrow Fe_3O_4(s) + H_2(g)$
2. Balance the following equations and name each reactant and product:
   a. $SF_4(g) + H_2O(\ell) \rightarrow SO_2(g) + HF(\ell)$
   b. $NH_3(aq) + O_2(aq) \rightarrow NO(g) + H_2O(\ell)$
   c. $BF_3(g) + H_2O(\ell) \rightarrow HF(aq) + H_3BO_3(aq)$
3. What mass of HCl, in grams, is required to react with 0.750 g of $Al(OH)_3$? What mass of water, in grams, is produced? What mass of $AlCl_3$, in grams, is produced? The equation: $Al(OH)_3(s) + 3\ HCl(aq) \rightarrow AlCl_3(aq) + 3\ H_2O(\ell)$
4. Hexane (C_6H_{14}) burns in air (O_2) to give CO_2 and H_2O. Write a balanced equation for this reaction. If 215 g of C_6H_{14} is mixed with 215 g of O_2, what masses of CO_2 and H_2O are produced in the reaction? What mass of excess reactant remains at the end of the reaction?
5. Consider the reaction: $2\ CH_3SH + CO \rightarrow CH_3COSCH_3 + H_2S$. If you begin with 10.0 g of CH_3SH and excess CO,
   a. What is the theoretical yield of CH_3COSCH_3?
   b. If 8.65 g of CH_3COSCH_3 is isolated, what is the percent yield?
6. A metal M reacts with O_2 according to the equation below. If 0.356 g of the metal M reacts with an excess of oxygen to make 0.452 g of the metal oxide MO_2, use this information to find the identity of the metal M. **$M(s) + O_2(g) \rightarrow MO_2(s)$**
7. Saccharin, an artificial sweetener, has the formula $C_7H_5NO_3S$. Suppose you have a sample of a saccharin-containing sweetener with a mass of 0.2140 g. After decomposition to free sulfur and converting it to the SO_4^{2-} ion, the sulfate ion is trapped as the water-insoluble $BaSO_4$. The quantity of $BaSO_4$ obtained is 0.2070 g. What is the mass percent of saccharin in the sample of sweetener?
8. To find the formula of a compound composed of iron and carbon monoxide, $Fe_x(CO)_y$, the compound is burned in pure oxygen to give Fe_2O_3 and CO_2. If you burn 1.959 g of $Fe_x(CO)_y$ and obtain 0.799 g of Fe_2O_3 and 2.200 g of CO_2, what is the empirical formula of $Fe_x(CO)_y$?
9. Mesitylene is a liquid hydrocarbon with formula C_xH_y. Burning 0.115 g of the compound in oxygen gives 0.379 g of CO_2 and 0.1035 g of H_2O. What is the empirical formula of mesitylene?
10. Benzoquinone, a chemical used in the dye industry and in photography, is an organic compound containing only C, H and O. What is the empirical formula of the compound if 0.105 g of the compound gives 0.257 g of CO_2 and 0.0350 g of H_2O when burned completely in oxygen? What is the molecular formula if the molar mass of the compound = 108 g/mol?

Problem Set #4 continues on next page

Note: For questions #11 - 13, use the **solubility table** found in the "**Net Ionic Reactions**" Lab, available in the Chemistry 221 Companion or on the website (http://mhchem.org/221/classroom/lab.htm) or here (https://mhchem.org/sol).

11. Decide whether each of the following is water-soluble. If soluble, tell what ions are produced. Describe them as strong electrolyte, weak electrolyte or non-electrolyte when placed in water.
    a. $NiCl_2$
    b. $Cr(NO_3)_3$
    c. ethanol
    d. ammonia
    e. $BaSO_4$
12. Predict the products of each precipitation reaction. Balance the completed equation, and then write the net ionic equation.
    a. $Pb(NO_3)_2(aq) + KBr(aq) \rightarrow$
    b. $Ca(NO_3)_2(aq) + KF(aq) \rightarrow$
13. Balance the following equations, and then write the net ionic equation. Identify the spectator ions, if any.
    a. $Mg(OH)_2(s) + HCl(aq) \rightarrow MgCl_2(aq) + H_2O(\ell)$
    b. $HNO_3(aq) + CaCO_3(s) \rightarrow Ca(NO_3)_2(aq) + H_2O(\ell) + CO_2(g)$

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CH 221 Fall 2024:

Problem Set #5

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #5 by **reading Chapter 3 and Chapter 5** of the textbook and/or by watching the videos found on our website (https://mhchem.org/221)
- **Try the problems** for Problem Set #5 found on the next pages on your own first. **Write out the answers (and show your work) by hand (on a tablet or paper)**; do not type your answers (and work) to avoid a point penalty. If you write the answers on the problem set itself, you will receive fewer points. Include your name on your problem set!

Step Two:

<u>*Section 01 and H1*</u>*:* We will go over Problem Set #5 during recitation. ***Self correct all problems*** of your problem set before turning it in at the end of recitation.

- *Section 01:* due **Monday, November 4 at 1:10 PM**
- *Section H1:* due **Wednesday, November 6 at 1:10 PM**

<u>*Section W1*</u>*:* **Watch the recitation video** for Problem Set #5:
http://mhchem.org/w/k.htm

- **Self correct *all* of the problems** while viewing the video. Mark correct problems with a star (or other similar mark), and correct all incorrect problems (show the correct answer and the steps required to achieve it.)
- **Submit Problem Set #5 via email (mike.russell@mhcc.edu) as a single PDF file** (use CamScanner (https://camscanner.com), CombinePDF (https://combinepdf.com), etc.) **by 11:59 PM Wednesday, November 6.**

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

CH 221 Problem Set #5

** Complete problem set on separate pieces of paper showing all work, circling final answers, etc.*
** Self correct your work before turning it in to the instructor.*

Covering: **Chapter Three (3.3-3.4), Chapter Five** *and* **Chapter Guide Five**
Important Tables / Constants: **$C(H_2O) = 4.184\ J\ g^{-1}\ K^{-1}$, $\log_{10} x = \ln x / \ln 10$** *and the* ***Thermodynamic Values*** *found after this problem set and here:* **http://mhchem.org/thermo**

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1. Determine the oxidation number of each element in the following ions or compounds.
   a. $O_2(g)$
   b. CuO
   c. UO^{2+}
   d. $H_2AsO_4^{-1}$
   e. OF_2
   f. XeO_4^{2-}
2. Which of the following reactions are oxidation-reduction reactions? Explain your answer briefly. Classify the remaining reactions.
   a. $CdCl_2(aq) + Na_2S(aq) \rightarrow CdS(s) + 2\ NaCl(aq)$
   b. $2\ Ca(s) + O_2(g) \rightarrow 2\ CaO(s)$
   c. $4\ Fe(OH)_2(aq) + 2\ H_2O(\ell) + O_2(g) \rightarrow 4\ Fe(OH)_3(aq)$
   d. $MgCO_3(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2O(\ell) + CO_2(g)$
3. In the following reactions, decide which reactant is oxidized and which is reduced. Designate the oxidizing agent and the reducing agent.
   a. $Cr_2O_7^{2-}(aq) + 3\ Sn^{2+}(aq) + 14\ H^+(aq) \rightarrow 2\ Cr^{3+}(aq) + 3\ Sn^{4+}(aq) + 7\ H_2O(\ell)$
   b. $FeS(s) + 3\ NO_3^{-1}(aq) + 4\ H^+(aq) \rightarrow 3\ NO(g) + SO_4^{2-}(aq) + Fe^{3+}(aq) + 2\ H_2O(\ell)$
4. What is the mass of solute, in grams, in 125 mL of a 1.023×10^{-3} M solution of Na_3PO_4? What are the molar concentrations of Na^+ and PO_4^{3-} ions?
5. Twelve (12.0) mL of a 0.125 M $BaCl_2$ solution is diluted with 9.0 mL of water to a total volume of 21.0 mL. What is the concentration (M) of the new solution? What is the concentration (M) of Cl^{-1} in the final solution? How many grams of $BaCl_2$ are in the final solution?
6. A saturated solution of milk of magnesia, $Mg(OH)_2$, has a pH of 10.5. What is the hydrogen ion concentration of the solution? Is the solution acidic or basic?
7. What mass of Na_2CO_3, in grams, is required for complete reaction with 50.0 mL of 0.125 M HNO_3? $Na_2CO_3(aq) + 2\ HNO_3(aq) \rightarrow 2\ NaNO_3(aq) + CO_2(g) + H_2O(\ell)$
8. Potassium hydrogen phthalate, $KHC_8H_4O_4$, is used to standardize solutions of bases. The acidic anion reacts with strong bases according to the net ionic equation shown below. If a 0.902 g sample of potassium hydrogen phthalate is dissolved in water and titrated to the equivalence point with 26.45 mL of NaOH, what is the molar concentration of the NaOH?
$$HC_8H_4O_4^{-1}(aq) + OH^{-1}(aq) \rightarrow C_8H_4O_4^{-2}(aq) + H_2O(\ell)$$
9. What quantity of heat is required to raise the temperature of 50.00 mL of water from 25.52 °C to 28.75 °C? The density of water at this temperature is 0.997 g/mL.
10. After absorbing 1.850 kJ of heat, the temperature of a 0.500 kg block of copper is 37 °C. What was the initial temperature of the copper? $C_{Cu} = 0.385$ J/g°C

Problem Set #5 continues on next page

11. A 45.5 g sample of copper at 99.8 °C is dropped into a beaker containing 152 g of water at 18.5 °C. What is the final temperature when thermal equilibrium is reached? C_{Cu} = 0.385 J/g°C
12. A piece of chromium metal with a mass of 24.26 g is heated in boiling water to 98.3 °C and then dropped into a coffee cup calorimeter containing 82.3 g of water at 23.3 °C. When thermal equilibrium is reached, the final temperature is 25.6 °C. Calculate the specific heat of the chromium.
13. Chloromethane, CH_3Cl, arises from the oceans and from microbial fermentation and is found throughout the environment. It is used in the manufacture of various chemicals and has been used as a topical anesthetic. What quantity of heat must be absorbed to convert 92.5 g of liquid to a vapor at its boiling point, -24.09 °C? The heat of vaporization of CH_3Cl is 21.40 kJ/mol.
14. Calcium carbide, CaC_2, is manufactured by the reaction of CaO with carbon at high temperatures. Calcium carbide can then be used to make acetylene. Using the reaction below, is this reaction endothermic or exothermic? If 10.0 g of CaO is allowed to react with an excess of carbon, what quantity of heat is absorbed or evolved by the reaction?
$$CaO(s) + 3\ C(s) \rightarrow CaC_2(s) + CO(g) \quad \Delta H^\circ_{rxn} = +464.8\ kJ$$
15. The enthalpy changes of the following reactions can be measured:
$$C_2H_4(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(\ell) \quad \Delta H^\circ = -1411.1\ kJ$$
$$C_2H_5OH(\ell) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(\ell) \quad \Delta H^\circ = -1367.5\ kJ$$
Use these values and Hess's law to determine the enthalpy change for the reaction:
$$C_2H_4(g) + H_2O(\ell) \rightarrow C_2H_5OH(\ell)$$
16. You wish to know the enthalpy change for the formation of liquid PCl_3 from the elements:
$$\tfrac{1}{4}\ P_4(s) + \tfrac{3}{2}\ Cl_2(g) \rightarrow PCl_3(\ell)$$
The enthalpy change for the formation of solid PCl_5 from the elements can be determined experimentally, as can the enthalpy change for the reaction of $PCl_3(\ell)$ with more chlorine to give $PCl_5(s)$. Use the equations below to calculate the enthalpy change for the formation of 1.00 mol $PCl_3(\ell)$ from phosphorus and chlorine.
$$P_4(s) + 10\ Cl_2(g) \rightarrow 4\ PCl_5(s) \quad \Delta H^\circ = -1774.0\ kJ$$
$$PCl_3(\ell) + Cl_2(g) \rightarrow PCl_5(s) \quad \Delta H^\circ = -123.8\ kJ$$
17. Write a balanced chemical equation for the formation of $CaCO_3(s)$ from the elements in their standard states. Find the value of ΔH_f° for $CaCO_3(s)$ in the table of thermodynamic values. If 10.0 g of $CaCO_3(s)$ forms from the elements, how much energy is required or will be released?
18. The first step in the production of nitric acid from ammonia involves the oxidation of NH_3:
$$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$$
    a. Use standard enthalpies of formation to calculate the standard enthalpy change for this reaction.
    b. Using this reaction, what quantity of heat is evolved or absorbed in the *formation* of 10.0 g of NH_3?

Standard Thermodynamic Properties for Selected Substances

Substance	ΔH°_f (kJ/mol)	ΔG°_f (kJ/mol)	S° (J/K•mol)
aluminum			
$Al(s)$	0	0	28.3
$Al(g)$	324.4	285.7	164.54
$Al_2O_3(s)$	–1676	–1582	50.92
$AlF_3(s)$	–1510.4	–1425	66.5
$AlCl_3(s)$	–704.2	–628.8	110.67
$AlCl_3 \cdot 6H_2O(s)$	–2691.57	–2269.40	376.56
$Al_2S_3(s)$	–724.0	–492.4	116.9
$Al_2(SO_4)_3(s)$	–3445.06	–3506.61	239.32
antimony			
$Sb(s)$	0	0	45.69
$Sb(g)$	262.34	222.17	180.16
$Sb_4O_6(s)$	–1440.55	–1268.17	220.92
$SbCl_3(g)$	–313.8	–301.2	337.80
$SbCl_5(g)$	–394.34	–334.29	401.94
$Sb_2S_3(s)$	–174.89	–173.64	182.00
$SbCl_3(s)$	–382.17	–323.72	184.10
$SbOCl(s)$	–374.0	-	-
arsenic			
$As(s)$	0	0	35.1
$As(g)$	302.5	261.0	174.21
$As_4(g)$	143.9	92.4	314
$As_4O_6(s)$	–1313.94	–1152.52	214.22
$As_2O_5(s)$	–924.87	–782.41	105.44
$AsCl_3(g)$	–261.50	–248.95	327.06
$As_2S_3(s)$	–169.03	–168.62	163.59
$AsH_3(g)$	66.44	68.93	222.78
$H_3AsO_4(s)$	–906.3	—	—
barium			
$Ba(s)$	0	0	62.5
$Ba(g)$	180	146	170.24
$BaO(s)$	–548.0	–520.3	72.1
$BaCl_2(s)$	–855.0	–806.7	123.7
$BaSO_4(s)$	–1473.2	–1362.3	132.2
beryllium			
$Be(s)$	0	0	9.50
$Be(g)$	324.3	286.6	136.27
$BeO(s)$	–609.4	–580.1	13.8
bismuth			
$Bi(s)$	0	0	56.74
$Bi(g)$	207.1	168.2	187.00
$Bi_2O_3(s)$	–573.88	–493.7	151.5
$BiCl_3(s)$	–379.07	–315.06	176.98
$Bi_2S_3(s)$	–143.1	–140.6	200.4
boron			
$B(s)$	0	0	5.86
$B(g)$	565.0	521.0	153.4
$B_2O_3(s)$	–1273.5	–1194.3	53.97
$B_2H_6(g)$	36.4	87.6	232.1
$H_3BO_3(s)$	–1094.33	–968.92	88.83
$BF_3(g)$	–1136.0	–1119.4	254.4
$BCl_3(g)$	–403.8	–388.7	290.1
$B_3N_3H_6(l)$	–540.99	–392.79	199.58

Standard Thermodynamic Properties for Selected Substances

Substance	$\Delta H°_f$ (kJ/mol)	$\Delta G°_f$ (kJ/mol)	$S°$ (J/K•mol)
boron *continued*			
$HBO_2(s)$	–794.25	–723.41	37.66
bromine			
$Br_2(l)$	0	0	152.23
$Br_2(g)$	30.91	3.142	245.5
$Br(g)$	111.88	82.429	175.0
$BrF_3(g)$	–255.60	–229.45	292.42
$HBr(g)$	–36.3	–53.43	198.7
cadmium			
$Cd(s)$	0	0	51.76
$Cd(g)$	112.01	77.41	167.75
$CdO(s)$	–258.2	–228.4	54.8
$CdCl_2(s)$	–391.5	–343.9	115.3
$CdSO_4(s)$	–933.3	–822.7	123.0
$CdS(s)$	–161.9	–156.5	64.9
calcium			
$Ca(s)$	0	0	41.6
$Ca(g)$	178.2	144.3	154.88
$CaO(s)$	–634.9	–603.3	38.1
$Ca(OH)_2(s)$	–985.2	–897.5	83.4
$CaSO_4(s)$	–1434.5	–1322.0	106.5
$CaSO_4 \cdot 2H_2O(s)$	–2022.63	–1797.45	194.14
$CaCO_3(s)$ (calcite)	–1220.0	–1081.4	110.0
$CaSO_3 \cdot H_2O(s)$	–1752.68	–1555.19	184.10
carbon			
$C(s)$ (graphite)	0	0	5.740
$C(s)$ (diamond)	1.89	2.90	2.38
$C(g)$	716.681	671.2	158.1
$CO(g)$	–110.52	–137.15	197.7
$CO_2(g)$	–393.51	–394.36	213.8
$CH_4(g)$	–74.6	–50.5	186.3
$CH_3OH(l)$	–239.2	–166.6	126.8
$CH_3OH(g)$	–201.0	–162.3	239.9
$CCl_4(l)$	–128.2	–62.5	214.4
$CCl_4(g)$	–95.7	–58.2	309.7
$CHCl_3(l)$	–134.1	–73.7	201.7
$CHCl_3(g)$	–103.14	–70.34	295.71
$CS_2(l)$	89.70	65.27	151.34
$CS_2(g)$	116.9	66.8	238.0
$C_2H_2(g)$	227.4	209.2	200.9
$C_2H_4(g)$	52.4	68.4	219.3
$C_2H_6(g)$	–84.0	–32.0	229.2
$CH_3CO_2H(l)$	–484.3	–389.9	159.8
$CH_3CO_2H(g)$	–434.84	–376.69	282.50
$C_2H_5OH(l)$	–277.6	–174.8	160.7
$C_2H_5OH(g)$	–234.8	–167.9	281.6
$C_3H_8(g)$	–103.8	–23.4	270.3
$C_6H_6(g)$	82.927	129.66	269.2
$C_6H_6(l)$	49.1	124.50	173.4
$CH_2Cl_2(l)$	–124.2	–63.2	177.8
$CH_2Cl_2(g)$	–95.4	–65.90	270.2
$CH_3Cl(g)$	–81.9	–60.2	234.6

Standard Thermodynamic Properties for Selected Substances

Substance	$\Delta H°_f$ (kJ/mol)	$\Delta G°_f$ (kJ/mol)	$S°$ (J/K•mol)
carbon *continued*			
$C_2H_5Cl(l)$	–136.52	–59.31	190.79
$C_2H_5Cl(g)$	–112.17	–60.39	276.00
$C_2N_2(g)$	308.98	297.36	241.90
$HCN(l)$	108.9	125.0	112.8
$HCN(g)$	135.5	124.7	201.8
chlorine			
$Cl2(g)$	0	0	223.1
$Cl(g)$	121.3	105.70	165.2
$ClF(g)$	–54.48	–55.94	217.78
$ClF_3(g)$	–158.99	–118.83	281.50
$Cl_2O(g)$	80.3	97.9	266.2
$Cl_2O_7(l)$	238.1	—	—
$Cl_2O_7(g)$	272.0	—	—
$HCl(g)$	–92.307	–95.299	186.9
$HClO_4(l)$	–40.58	—	—
chromium			
$Cr(s)$	0	0	23.77
$Cr(g)$	396.6	351.8	174.50
$Cr_2O_3(s)$	–1139.7	–1058.1	81.2
$CrO_3(s)$	–589.5	—	—
$(NH_4)_2Cr_2O_7(s)$	–1806.7	—	—
cobalt			
$Co(s)$	0	0	30.0
$CoO(s)$	–237.9	–214.2	52.97
$Co_3O_4(s)$	–910.02	–794.98	114.22
$Co(NO_3)_2(s)$	–420.5	—	—
copper			
$Cu(s)$	0	0	33.15
$Cu(g)$	338.32	298.58	166.38
$CuO(s)$	–157.3	–129.7	42.63
$Cu_2O(s)$	–168.6	–146.0	93.14
$CuS(s)$	–53.1	–53.6	66.5
$Cu_2S(s)$	–79.5	–86.2	120.9
$CuSO_4(s)$	–771.36	–662.2	109.2
$Cu(NO_3)_2(s)$	–302.9	—	—
fluorine			
$F_2(g)$	0	0	202.8
$F(g)$	79.4	62.3	158.8
$F_2O(g)$	24.7	41.9	247.43
$HF(g)$	–273.3	–275.4	173.8
hydrogen			
$H_2(g)$	0	0	130.7
$H(g)$	217.97	203.26	114.7
$H_2O(l)$	–285.83	–237.1	70.0
$H_2O(g)$	–241.82	–228.59	188.8
$H_2O_2(l)$	–187.78	–120.35	109.6
$H_2O_2(g)$	–136.3	–105.6	232.7
$HF(g)$	–273.3	–275.4	173.8
$HCl(g)$	–92.307	–95.299	186.9
$HBr(g)$	–36.3	–53.43	198.7
$HI(g)$	26.48	1.70	206.59

Standard Thermodynamic Properties for Selected Substances

Substance	ΔH°_f (kJ/mol)	ΔG°_f (kJ/mol)	S° (J/K•mol)
hydrogen *continued*			
$H_2S(g)$	–20.6	–33.4	205.8
$H_2Se(g)$	29.7	15.9	219.0
iodine			
$I_2(s)$	0	0	116.14
$I_2(g)$	62.438	19.3	260.7
$I(g)$	106.84	70.2	180.8
$IF(g)$	95.65	–118.49	236.06
$ICl(g)$	17.78	–5.44	247.44
$IBr(g)$	40.84	3.72	258.66
$IF_7(g)$	–943.91	–818.39	346.44
$HI(g)$	26.48	1.70	206.59
iron			
$Fe(s)$	0	0	27.3
$Fe(g)$	416.3	370.7	180.5
$Fe_2O_3(s)$	–824.2	–742.2	87.40
$Fe_3O_4(s)$	–1118.4	–1015.4	146.4
$Fe(CO)_5(l)$	–774.04	–705.42	338.07
$Fe(CO)_5(g)$	–733.87	–697.26	445.18
$FeCl_2(s)$	–341.79	–302.30	117.95
$FeCl_3(s)$	–399.49	–334.00	142.3
$FeO(s)$	–272.0	–255.2	60.75
$Fe(OH)_2(s)$	–569.0	–486.5	88.
$Fe(OH)_3(s)$	–823.0	–696.5	106.7
$FeS(s)$	–100.0	–100.4	60.29
$Fe_3C(s)$	25.10	20.08	104.60
lead			
$Pb(s)$	0	0	64.81
$Pb(g)$	195.2	162.	175.4
$PbO(s)$ (yellow)	–217.32	–187.89	68.70
$PbO(s)$ (red)	–218.99	–188.93	66.5
$Pb(OH)_2(s)$	–515.9	—	—
$PbS(s)$	–100.4	–98.7	91.2
$Pb(NO_3)_2(s)$	–451.9	—	—
$PbO_2(s)$	–277.4	–217.3	68.6
$PbCl_2(s)$	–359.4	–314.1	136.0
lithium			
$Li(s)$	0	0	29.1
$Li(g)$	159.3	126.6	138.8
$LiH(s)$	–90.5	–68.3	20.0
$LiOH(s)$	–487.5	–441.5	42.8
$LiF(s)$	–616.0	–587.5	35.7
$Li_2CO_3(s)$	–1216.04	–1132.19	90.17
manganese			
$Mn(s)$	0	0	32.0
$Mn(g)$	280.7	238.5	173.7
$MnO(s)$	–385.2	–362.9	59.71
$MnO_2(s)$	–520.03	–465.1	53.05
$Mn_2O_3(s)$	–958.97	–881.15	110.46
$Mn_3O_4(s)$	–1378.83	–1283.23	155.64
mercury			
$Hg(l)$	0	0	75.9

Standard Thermodynamic Properties for Selected Substances

Substance	$\Delta H°_f$ (kJ/mol)	$\Delta G°_f$ (kJ/mol)	$S°$ (J/K•mol)
mercury *continued*			
Hg(*g*)	61.4	31.8	175.0
HgO(*s*) (red)	–90.83	–58.5	70.29
HgO(*s*) (yellow)	–90.46	–58.43	71.13
$HgCl_2(s)$	–224.3	–178.6	146.0
$Hg_2Cl_2(s)$	–265.4	–210.7	191.6
HgS(*s*) (red)	–58.16	–50.6	82.4
HgS(*s*) (black)	–53.56	–47.70	88.28
$HgSO_4(s)$	–707.51	–594.13	0.00
nitrogen			
$N_2(g)$	0	0	191.6
N(*g*)	472.704	455.5	153.3
NO(*g*)	90.25	87.6	210.8
$NO_2(g)$	33.2	51.30	240.1
$N_2O(g)$	81.6	103.7	220.0
$N_2O_3(g)$	83.72	139.41	312.17
$N_2O_4(g)$	11.1	99.8	304.4
$N_2O_5(g)$	11.3	115.1	355.7
$NH_3(g)$	–45.9	–16.5	192.8
$N_2H_4(l)$	50.63	149.43	121.21
$N_2H_4(g)$	95.4	159.4	238.5
$NH_4NO_3(s)$	–365.56	–183.87	151.08
$NH_4Cl(s)$	–314.43	–202.87	94.6
$NH_4Br(s)$	–270.8	–175.2	113.0
$NH_4I(s)$	–201.4	–112.5	117.0
$NH_4NO_2(s)$	–256.5	—	—
$HNO_3(l)$	–174.1	–80.7	155.6
$HNO_3(g)$	–133.9	–73.5	266.9
oxygen			
$O_2(g)$	0	0	205.2
O(*g*)	249.17	231.7	161.1
$O_3(g)$	142.7	163.2	238.9
phosphorus			
$P_4(s)$	0	0	164.4
$P_4(g)$	58.91	24.4	280.0
P(*g*)	314.64	278.25	163.19
$PH_3(g)$	5.4	13.5	210.2
$PCl_3(g)$	–287.0	–267.8	311.78
$PCl_5(g)$	–374.9	–305.0	364.4
$P_4O_6(s)$	–1640.1	—	—
$P_4O_{10}(s)$	–2984.0	–2697.0	228.86
$HPO_3(s)$	–948.5	—	—
$H_3PO_2(s)$	–604.6	—	—
$H_3PO_3(s)$	–964.4	—	—
$H_3PO_4(s)$	–1279.0	–1119.1	110.50
$H_3PO_4(l)$	–1266.9	–1124.3	110.5
$H_4P_2O_7(s)$	–2241.0	—	—
$POCl_3(l)$	–597.1	–520.8	222.5
$POCl_3(g)$	–558.5	–512.9	325.5
potassium			
K(*s*)	0	0	64.7
K(*g*)	89.0	60.5	160.3

Standard Thermodynamic Properties for Selected Substances

Substance	$\Delta H°_f$ (kJ/mol)	$\Delta G°_f$ (kJ/mol)	$S°$ (J/K•mol)
potassium *continued*			
$KF(s)$	–576.27	–537.75	66.57
$KCl(s)$	–436.5	–408.5	82.6
silicon			
$Si(s)$	0	0	18.8
$Si(g)$	450.0	405.5	168.0
$SiO_2(s)$	–910.7	–856.3	41.5
$SiH_4(g)$	34.3	56.9	204.6
$H_2SiO_3(s)$	–1188.67	–1092.44	133.89
$H_4SiO_4(s)$	–1481.14	–1333.02	192.46
$SiF_4(g)$	–1615.0	–1572.8	282.8
$SiCl_4(l)$	–687.0	–619.8	239.7
$SiCl_4(g)$	–662.75	–622.58	330.62
SiC(*s, beta cubic*)	–73.22	–70.71	16.61
SiC(*s, alpha hexagonal*)	–71.55	–69.04	16.48
silver			
$Ag(s)$	0	0	42.55
$Ag(g)$	284.9	246.0	172.89
$Ag_2O(s)$	–31.05	–11.20	121.3
$AgCl(s)$	–127.0	–109.8	96.3
$Ag_2S(s)$	–32.6	–40.7	144.0
sodium			
$Na(s)$	0	0	51.3
$Na(g)$	107.5	77.0	153.7
$Na_2O(s)$	–414.2	–375.5	75.1
$NaCl(s)$	–411.2	–384.1	72.1
sulfur			
$S_8(s)$ (rhombic)	0	0	256.8
$S(g)$	278.81	238.25	167.82
$SO_2(g)$	–296.83	–300.1	248.2
$SO_3(g)$	–395.72	–371.06	256.76
$H_2S(g)$	–20.6	–33.4	205.8
$H_2SO_4(l)$	–813.989	690.00	156.90
$H_2S_2O_7(s)$	–1273.6	—	—
$SF_4(g)$	–728.43	–684.84	291.12
$SF_6(g)$	–1220.5	–1116.5	291.5
$SCl_2(l)$	–50	—	—
$SCl_2(g)$	–19.7	—	—
$S_2Cl_2(l)$	–59.4	—	—
$S_2Cl_2(g)$	–19.50	–29.25	319.45
$SOCl_2(g)$	–212.55	–198.32	309.66
$SOCl_2(l)$	–245.6	—	—
$SO_2Cl_2(l)$	–394.1	—	—
$SO_2Cl_2(g)$	–354.80	–310.45	311.83
tin			
$Sn(s)$	0	0	51.2
$Sn(g)$	301.2	266.2	168.5
$SnO(s)$	–285.8	–256.9	56.5
$SnO_2(s)$	–577.6	–515.8	49.0
$SnCl_4(l)$	–511.3	–440.1	258.6
$SnCl_4(g)$	–471.5	–432.2	365.8

Standard Thermodynamic Properties for Selected Substances

Substance	$\Delta H°_f$ (kJ/mol)	$\Delta G°_f$ (kJ/mol)	$S°$ (J/K•mol)
titanium			
Ti(*s*)	0	0	30.7
Ti(*g*)	473.0	428.4	180.3
TiO_2(*s*)	–944.0	–888.8	50.6
$TiCl_4$(*l*)	–804.2	–737.2	252.4
$TiCl_4$(*g*)	–763.2	–726.3	353.2
tungsten			
W(*s*)	0	0	32.6
W(*g*)	849.4	807.1	174.0
WO_3(*s*)	–842.9	–764.0	75.9
zinc			
Zn(*s*)	0	0	41.6
Zn(*g*)	130.73	95.14	160.98
ZnO(*s*)	–350.5	–320.5	43.7
$ZnCl_2$(*s*)	–415.1	–369.43	111.5
ZnS(*s*)	–206.0	–201.3	57.7
$ZnSO_4$(*s*)	–982.8	–871.5	110.5
$ZnCO_3$(*s*)	–812.78	–731.57	82.42
complexes			
cis-$[Co(NH_3)_4(NO_2)_2]NO_3$	–898.7	—	—
trans-$[Co(NH_3)_4(NO_2)_2]NO_3$	–896.2	—	—
$NH_4[Co(NH_3)_2(NO_2)_4]$	–837.6	—	—
$[Co(NH_3)_6][Co(NH_3)_2(NO_2)_4]_3$	–2733.0	—	—
cis-$[Co(NH_3)_4Cl_2]Cl$	–874.9	—	—
trans-$[Co(NH_3)_4Cl_2]Cl$	–877.4	—	—
cis-$[Co(en)_2(NO_2)_2]NO_3$	–689.5	—	—
cis-$[Co(en)_2Cl_2]Cl$	–681.2	—	—
trans-$[Co(en)_2Cl_2]Cl$	–677.4	—	—
$[Co(en)_3](ClO_4)_3$	–762.7	—	—
$[Co(en)_3]Br_2$	–595.8	—	—
$[Co(en)_3]I_2$	–475.3	—	—
$[Co(en)_3]I_3$	–519.2	—	—
$[Co(NH_3)_6](ClO_4)_3$	–1034.7	–221.1	615
$[Co(NH_3)_5NO_2](NO_3)_2$	–1088.7	–412.9	331
$[Co(NH_3)_6](NO_3)_3$	–1282.0	–524.5	448
$[Co(NH_3)_5Cl]Cl_2$	–1017.1	–582.5	366.1
$[Pt(NH_3)_4]Cl_2$	–725.5	—	—
$[Ni(NH_3)_6]Cl_2$	–994.1	—	—
$[Ni(NH_3)_6]Br_2$	–923.8	—	—
$[Ni(NH_3)_6]I_2$	–808.3	—	—

CH 221 Exam Prep II Worksheet Name: ____________________

Worksheet due dates: **<u>Mon, 11/18</u>, 1:10 PM** *(01)* **, <u>Wed, 11/13</u>, 1:10 PM** *(H1)* ***or*** **11:59 PM** *(W1, email)*. To complete, show ***detailed*** *steps* on how to get the given answer for each problem. *Failure to use this form for work and answers will result in a point penalty.*

<u>Problem 1</u>: You take an aspirin tablet (which contains only carbon, hydrogen and oxygen) with a mass of 1.000 g and burn it in air to collect 2.20 g of carbon dioxide and 0.400 g of water. A molar mass experiment shows a value between 170 and 190 g/mol. What is the molecular formula for aspirin?

Answer to Problem #1: **$C_9H_8O_4$**

<u>Problem 2</u>: The reaction of 23.1 g of NH_3 and 18.3 g of O_2 produces 4.10 g of NO. What is the percent yield for this reaction? The equation for this reaction is: **$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$** *Note:* Check ***both*** reactants for credit on this problem.

Answer to Problem #2: **29.9%**

Problem 3: What volume of 0.300 M NaCl is required to precipitate all the Pb^{2+} ion from 25.0 mL of aqueous 0.440 M $Pb(NO_3)_2$? The reaction is: **$Pb(NO_3)_2(aq) + 2\ NaCl(aq) \rightarrow PbCl_2(s) + 2\ NaNO_3(aq)$**

Answer to Problem #3: **73.3 mL**

Problem 4: If 1.00 mole of ethanol, CH_3CH_2OH, at 22.0 °C absorbs 1.45 kJ of heat, what is the final temperature of the ethanol? The specific heat capacity for ethanol is 2.44 J/gK.

Answer to Problem #4: **34.9 °C**

Problem 5: The standard molar enthalpy of formation of $NH_3(g)$ is -45.9 kJ/mol. What is the enthalpy change if 5.38 g of $N_2(g)$ and 3.32 g of $H_2(g)$ react to form $NH_3(g)$? *Note:* Check ***both*** reactants for credit on this problem.

Answer to Problem #5: **-17.6 kJ**

CH 221 Fall 2024:

Problem Set #6

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #6 by **reading Chapter 6** of the textbook and/or by watching the videos found on our website (https://mhchem.org/221)
- **Try the problems** for Problem Set #6 found on the next pages on your own first. **Write out the answers (and show your work) by hand (on a tablet or paper)**; do not type your answers (and work) to avoid a point penalty. If you write the answers on the problem set itself, you will receive fewer points. Include your name on your problem set!

Step Two:

Section 01 and H1: We will go over Problem Set #6 during recitation. ***Self correct all problems*** of your problem set before turning it in at the end of recitation.

- *Section 01:* due **Monday, December 2 at 1:10 PM**
- *Section H1:* due **Wednesday, December 4 at 1:10 PM**

Section W1: **Watch the recitation video** for Problem Set #6:
http://mhchem.org/w/c.htm

- **Self correct *all* of the problems** while viewing the video. Mark correct problems with a star (or other similar mark), and correct all incorrect problems (show the correct answer and the steps required to achieve it.)
- **Submit Problem Set #6 via email (mike.russell@mhcc.edu) as a single PDF file** (use CamScanner (https://camscanner.com), CombinePDF (https://combinepdf.com), etc.) **by 11:59 PM Wednesday, December 4.**

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

CH 221 Problem Set #6

** Complete problem set on separate pieces of paper showing all work, circling final answers, etc.*
** Self correct your work before turning it in to the instructor.*

Covering: **Chapter Six** *and* **Chapter Guide Six**
Important Tables and/or Constants: **c = 2.998 x 10^{8} m/s, h = 6.626 x 10^{-34} J s**, the **Electromagnetic Spectrum** *and* **Subshell Filling Order diagrams** *on page 4 of Problem Set #6*. ***Memorize c and h!***

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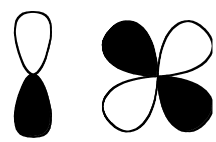
1. Consider the electromagnetic spectrum:
   a. What color of light has photons of greater energy, yellow or blue?
   b. Which color of light has the greater frequency, blue or green?
   c. Place the following types of radiation in order of increasing energy per photon.
      i. radar signals (RADAR = RAdio Detection And Recognition)
      ii. radiation within a microwave oven
      iii. gamma rays from a nuclear reaction
      iv. red light from a neon sign
      v. ultraviolet radiation from a sun lamp
2. The most prominent line in the spectrum of magnesium is 285.2 nm; other lines are found at 383.8 and 518.4 nm. In which regions of the electromagnetic spectrum are these lines found? Which is the most energetic line? What is the frequency and energy (in both Joules per photon and kJ per mol) of the wavelength of the most energetic line?
3. Consider only transitions involving the n = 1 through n = 4 energy levels for the hydrogen atom:
   a. How many emission lines are possible?
   b. Photons of the lowest energy are emitted in a transition from the level with n = ___ to a level with n = ___.
   c. The emission line having the shortest wavelength corresponds to a transition from the level with n = ___ to the level with n = ___.
4. A beam of electrons (m = 9.11 x 10^{-31} kg/electron) has an average speed of 1.3 x 10^{8} m s^{-1}. What is the wavelength of electrons having this average speed? *(Note to physics fans: no relativity in this problem!)*
5. Answer the following questions:
   a. When $n = 4$, $\ell = 2$ and $m_l = -1$, to what orbital type does this refer? (Use the subshell label, such as 1*s*.)
   b. How many orbitals occur in the n = 5 electron shell? How many subshells? What are the letter labels of the subshells?
   c. If a subshell is labeled *f*, how many orbitals occur in the subshell? What are the values of m_l?
6. Explain briefly why each of the following is not a possible set of quantum numbers for an electron in an atom. In each case, change the incorrect value(s) to make the set valid.
   a. $n = 2, \ell = 2, m_l = 0, m_s = +1/2$
   b. $n = 2, \ell = 1, m_l = -1, m_s = 0$
   c. $n = 3, \ell = 1, m_l = +2, m_s = +1/2$

Problem Set #6 continues on next page

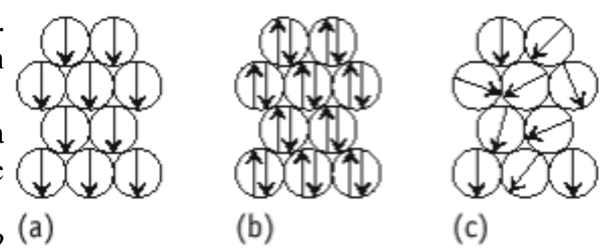
7. What is the maximum number of orbitals that can be identified by each of the following sets of quantum numbers?
   a. $n = 4$, $\ell = 3$
   b. $n = 4$
   c. $n = 2$, $\ell = 2$
   d. $n = 3$, $\ell = 1$, $m_l = -1$
8. State which of the following are incorrect designations for orbitals according to the quantum theory: 3*p*, 4*s*, 2*f*, and 1*p*. Briefly explain your answers.
9. How many nodal surfaces (planar *and* spherical) are associated with each of the following atomic orbitals?
   a. 4*f*
   b. 2*p*
   c. 6*s*
10. Answer the following questions:
    a. The quantum number n describes the _____ of an atomic orbital and the quantum number ℓ describes its _____.
    b. When $n = 3$, the possible values of ℓ are _____.
    c. What type of subshell corresponds to $l = 3$?
    d. For a 4*d* orbital, the value of n is ___, the value of ℓ is ___, and a possible value of m_l is ___.
    e. For each of the orbitals shown in the diagram on the right, give the letter designation for the orbital, the value of ℓ, and the number of planar nodal surfaces.



    f. An atomic orbital with three planar nodal surfaces is ______(letter).
    g. Which of the following orbitals cannot exist according to modern quantum theory? 2*s*, 3*p*, 2*d*, 3*f*, 5*p*, 6*p*
    h. Which of the following is *not* a valid set of quantum numbers?
       i. $n = 3$, $\ell = 2$, $m_l = 1$
       ii. $n = 2$, $\ell = 1$, $m_l = 2$
       iii. $n = 4$, $\ell = 3$, $m_l = 0$
    i. What is the maximum number of orbitals that can be associated with each of the following sets of quantum numbers?
       i. $n = 2$ and $\ell = 1$
       ii. $n = 3$
       iii. $n = 3$ and $\ell = 3$
       iv. $n = 2$, $\ell = 1$, $m_l = 0$
    j. Place the following subshells in order (from lowest to highest energy) using the $n + l$ rule: **1s 2s 2p 3s 3p 3d 4s 4p 4d 4f 5s 5p 5d 6s 6p**

Problem Set #6 continues on next page

11. In principle, which of the following can be determined?
    a. The energy of a high-speed electron in the H atom with high precision and accuracy
    b. The position of a high-speed electron with high precision and accuracy
    c. At the same time, both the position and the energy of a high-speed electron with high precision and accuracy.
12. Write the electron configuration for neutral Mg and Ar using both *spdf* notation and orbital box diagrams.
13. Using *spdf* and noble gas notations, write electron configurations for neutral atoms of the following elements:
    a. Strontium, Sr. This element is named for a town in Scotland.
    b. Zirconium, Zr. The metal is exceptionally resistant to corrosion and so has important industrial applications. Moon rocks show a surprisingly high zirconium content compared with rocks on earth.
    c. Rhodium, Rh. This metal is used in jewelry and in catalysts in industry.
    d. Tin, Sn. The metal was used in the ancient world. Alloys of tin (solder, bronze, pewter) are important.
14. Using orbital box diagrams, depict an electron configuration for each of the following ions:
    a. Na^{+}
    b. Al^{3+}
    c. Ge^{2+}
    d. F^{-}
15. Explain each answer briefly:
    a. Arrange the following elements in order of increasing size: Ca, Rb, P, Ge, Sr
    b. Which has the largest first ionization energy: O, S, or Se?
    c. Which has the most negative electron affinity: Se, Cl or Br?
    d. Which has the largest radius: O^{2-}, F^{-1} or F?
    e. Which is most paramagnetic: Fe^{3+} or Cr^{3+}? Explain.
16. The diagrams on the right represent a small section of a solid. Each circle represents an atom and an arrow represents an electron.
    a. Which represents a diamagnetic solid, which is a paramagnetic solid, and which is a ferromagnetic solid?
    b. Which is most strongly attracted to a magnetic field? Which is least attracted?



17. *Briefly* describe the notable achievements of the following individuals in relation to quantum theory: **Maxwell, Planck, Einstein, Bohr, de Broglie, Schrödinger, Heisenberg, Dirac, Pauli, Hund**

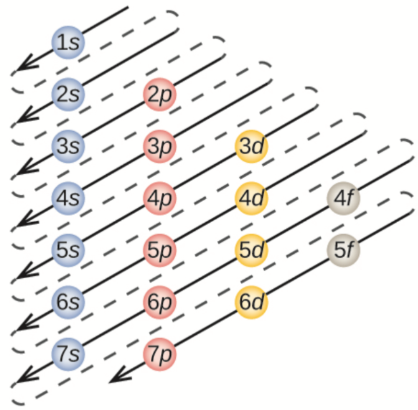


Figure 6.27 The arrow leads through each subshell in the appropriate filling order for electron configurations. This chart is straightforward to construct. Simply make a column for all the *s* orbitals with each *n* shell on a separate row. Repeat for *p*, *d*, and *f*. Be sure to only include orbitals allowed by the quantum numbers (no 1*p* or 2*d*, and so forth). Finally, draw diagonal lines from top to bottom as shown.

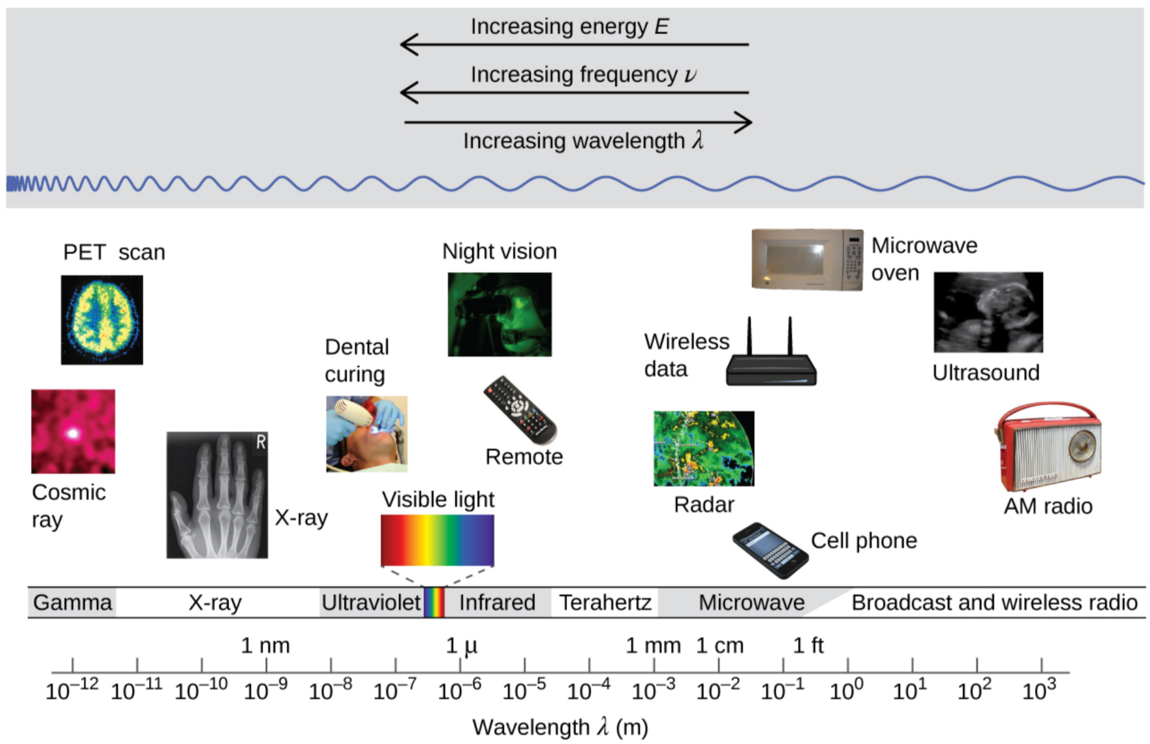


Figure 6.3 Portions of the electromagnetic spectrum are shown in order of decreasing frequency and increasing wavelength. Examples of some applications for various wavelengths include positron emission tomography (PET) scans, X-ray imaging, remote controls, wireless Internet, cellular telephones, and radios. (credit “Cosmic ray": modification of work by NASA; credit “PET scan": modification of work by the National Institute of Health; credit “X-ray": modification of work by Dr. Jochen Lengerke; credit “Dental curing": modification of work by the Department of the Navy; credit “Night vision": modification of work by the Department of the Army; credit “Remote": modification of work by Emilian Robert Vicol; credit “Cell phone": modification of work by Brett Jordan; credit “Microwave oven": modification of work by Billy Mabray; credit “Ultrasound": modification of work by Jane Whitney; credit “AM radio": modification of work by Dave Clausen)

Both images from “Chemistry” by OpenStax College (978-1-938168-39-0)

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CH 221 Final Exam Prep Worksheet Name: ______________________________

Worksheet due dates: **<u>At the time of your Lecture Final</u>** *(01, H1)* **, <u>Wed, 12/11</u>, 11:59 PM** *(W1, email)*. To complete, show ***detailed*** *steps* on how to get the given answer for each problem. *Failure to use this form for work and answers will result in a point penalty.*

<u>Problem 1</u>: If the de Broglie wavelength of an electron is 555 nm, what is its velocity? The mass of an electron is $9.1 * 10^{-31}$ kg.

Answer to Problem #1: **$1.3 * 10^{3}$ m/s**

<u>Problem 2</u>: What is the energy of a photon of blue light that has a wavelength of 450 nm? What is the energy per mole?

Answer to Problem #2: **$4.4 * 10^{-19}$ J; 260 kJ** *(270 kJ ok if more than 2 sig J value used)*

Problem 3: What is the electron configuration for Cr^{2+}? Cr^{3+}? Which is more paramagnetic? How many unpaired electrons does each paramagnetic ion have? (Use **orbital box notation** and **give the electron configuration for both ions** to receive credit!)

Answer to Problem #3: **The more paramagnetic species, Cr^{2+}, has four unpaired electrons.**

Problem 4: Photography lenses incorporate small amounts of silver(I) chloride in the glass of the lens. The following reaction occurs in the light, causing the lens to darken: **$AgCl(s) \rightarrow Ag(s) + Cl$**
The enthalpy change for this reaction is $3.10 * 10^2$ kJ/mol. Assuming all this energy is supplied by light, what is the maximum wavelength of light that can cause this reaction?

Answer to Problem #4: **$3.86 * 10^{-7}$ m**

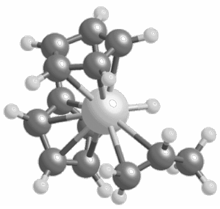
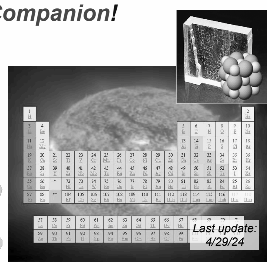
Problem 5: Using a strict interpretation of the $n + l$ rule, how many protons would an atom need to create a ground state electron configuration with one electron in a 5g orbital? (Give the **electron configuration starting with [Rn]** for the atom *in proper electron filling order* to receive credit)

Answer to Problem #5: **121 protons**

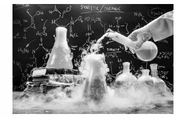

Welcome to Chemistry 221!

http://mhchem.org/221

Get the CH 221 Companion!

CH 221: Lectures and Labs



Lectures: **MWF** from **9 - 9:50 AM** in **AC 1303** *(this room)*
- Lectures recorded, available soon afterwards
- Lecture notes to print available (under "Problem Sets and Handouts", mhchem.org/221) and in ***Chemistry 221 Companion*** (get it!)

Labs (Section 01): **Mondays from 1:10 - 5 PM**
- **Start** in **room AC 2501**
- Move to **AC 2507** ("the lab") around **3 PM**
- *For first day*, **bring a printed copy of the "Eight Bottles" Lab** (mhchem.org/221) and your **calculator**
- Some labs will require safety glasses (Dollar store ok)

MAR

...more on Monday afternoon

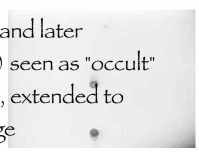
The Nature of Chemistry

What is Chemistry?
- "Keme" (earth)
- "Kehmeia" (transmutation)
- "Al-Khemia" (Arabic)
- "alchemy" (Europe's Dark Age)
- "chymistry" (Boyle's 1661 publication)
- "chemistry" (modern)

M.

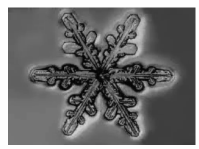
Khemeia - ancient Egyptian processes for embalming the dead, later extended to metallurgy

Khemeia (and later chemistry) seen as "occult" by laymen, extended to modern age



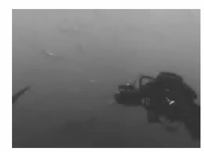
What is Matter?

The Nature of Chemistry



How does Matter Change?

How does Matter Interact?



Why Study Chemistry?

MAR

The Art (?) of Chemistry

Chemistry and Art?!?

Dr. Roald Hoffman,
1981 Nobel Prize in Chemistry

Stick to the chemistry, Roald!

"There was no question that
the reaction worked
but transient colors were seen
in the slurry of sodium methoxide in
dichloromethane
and we got a whole lot of products
for which we can't sort out the
kinetics
the next slide show will show
the most important part
very rapidly
within two minutes
and I forgot to say on further
warming
we get in fact the ketone…"

MAR

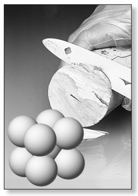
- Organic **- carbon, nitrogen, oxygen**
- Inorganic **- metals, everything "non-carbon"**
- Analytical **- Spectroscopy, "how much", "what kind"**
- Physical **- measurement, where physics meets chemistry**
- Biochemical **- the chemistry of life**
- ***many others!***

The Branches of Chemistry

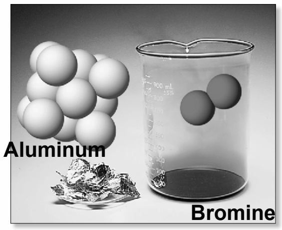
MAR

The Language of Chemistry

CHEMICAL ELEMENTS - pure substances that cannot be decomposed by ordinary means to other substances.



Sodium



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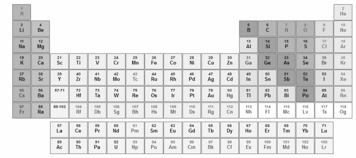
The Language of Chemistry

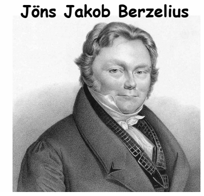
The elements, their names, and symbols are given on the PERIODIC TABLE

***Berzelius* - first to use letter symbols for atoms**

How many elements are there?



Jöns Jakob Berzelius



***pre-Biblical elements:* Au *(sun),* Ag *(moon),* Cu *(Venus),* Fe *(Mars),* Sn *(Jupiter),* Pb *(Saturn),* Hg *(Mercury),* S, C**

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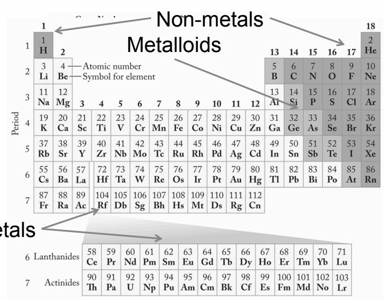
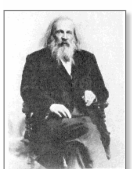
The Periodic Table



Periodic table originally organized by mass, now by atomic number

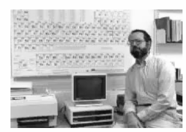


Dmitri Mendeleev (1834 - 1907)

Predicted Ga, Ge, Sc and Tc!

Dr. Frank DiSalvo (Cornell University)

"On the importance of the periodic table"



Number of compounds possible is virtually limitless!!!

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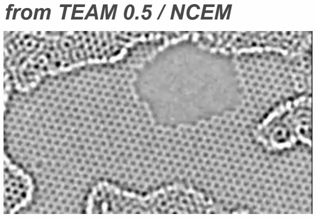
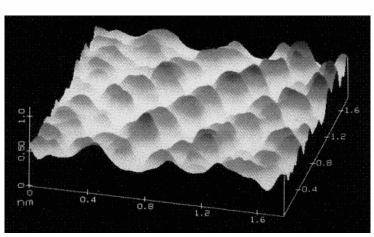
An atom is the smallest particle of an element that has the chemical properties of the element.

Real time carbon atoms from TEAM 0.5 / NCEM





Copper atoms on a silica surface
Distance across = 1.8 nanometer (1.8 x 10^{-9} m)

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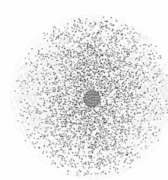
The Atom

An atom consists of a nucleus (of protons and neutrons) and electrons in space about the nucleus.



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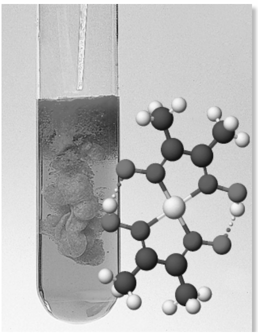


CHEMICAL COMPOUNDS are composed of atoms and so can be decomposed to those atoms.

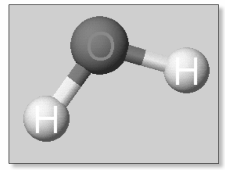
The red compound is composed of
- nickel (Ni) (silver)
- carbon (C) (black)
- hydrogen (H) (white)
- oxygen (O) (red)
- nitrogen (N) (blue)

This type of compound is an ionic compound - unshared electrons

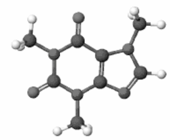
A MOLECULE is the smallest unit of a compound that retains the chemical characteristics of the compound.

Composition of molecules is given by a **MOLECULAR FORMULA**

H_2O



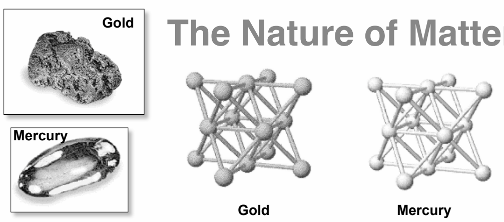
$C_8H_{10}N_4O_2$ - caffeine



Water and caffeine are examples of covalent compound - **shared electrons**

MAR

The Nature of Matter



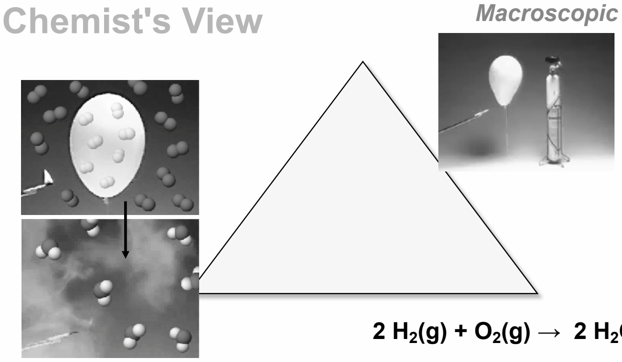
Chemists are interested in the nature of matter and how this is related to its atoms and molecules.

We can explore the MACROSCOPIC world - what we can see - to understand the ATOMIC world - what we cannot see - using SYMBOLS.

MAR

A Chemist's View

Macroscopic

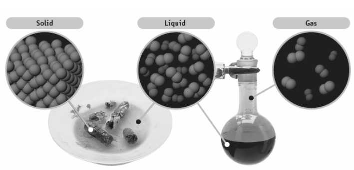


$2\ H_2(g) + O_2(g) \rightarrow\ 2\ H_2O(g)$

Atomic *Symbolic*

MAR

STATES OF MATTER



SOLIDS: rigid shape, fixed volume, reasonably well understood.

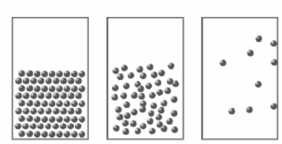
LIQUIDS: no fixed shape, may not fill a container completely, not well understood.

GASES: expand to fill their container, good theoretical understanding.

also PLASMA - more in CH 222!

KINETIC NATURE OF MATTER

Matter consists of atoms and molecules in motion.



Kinetic Molecular Theory describes solids, liquids and gases

~~Test Monkeys?~~ *Er, sorry,* Student volunteers?!?

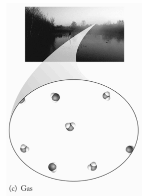
MAR

STATES OF MATTER & ENERGY

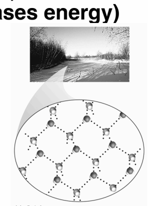
Energy of transitions *important to scientists:*

solid -> liquid -> gas: **endothermic** *(takes energy)*

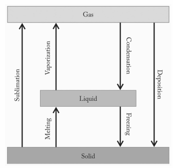
gas -> liquid -> solid: **exothermic** **(releases energy)**

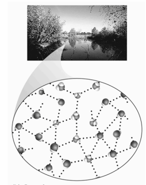


(c) Gas



(a) Solid





(b) Liquid

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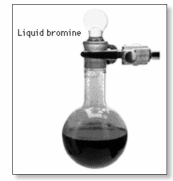
Physical Properties

Physical properties can be observed and measured without changing the composition of a substance. They include:

- **color**
- **melting and boiling point**
- **odor**



Some physical changes would be

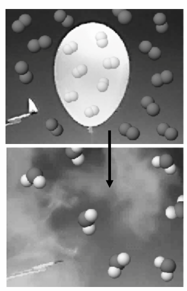
- boiling of a liquid
- melting of a solid
- dissolving a solid in a liquid to give a homogeneous mixture - a SOLUTION.

MAR

Chemical Properties and Chemical Change

Burning hydrogen (H_2) in oxygen (O_2) gives H_2O.

Chemical change **or chemical reaction involves the transformation of one or more atoms or molecules into one or more different molecules.**



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Chemical Properties and Physical Properties

Physical properties **do not change the composition of the substance**

Chemical properties **change the composition of the substance**

MAR

Physical Properties

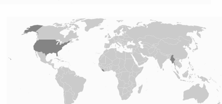
Physical properties useful in separating compounds and elements

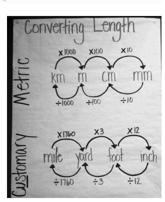
- **density**
- **melting and boiling point**
- **magnetism**

Physical and chemical properties require units **- need**

METRIC SYSTEM!

See the <u>Metric Guide</u>



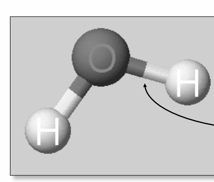


MAR

Units of Length / Conversions

1 kilometer (km)	**=**	**10^3 meters (m)**
1 centimeter (cm)	**=**	**10^{-2} meters (m)**
1 millimeter (mm)	**=**	**10^{-3} meters (m)**
1 micrometers (μm)	**=**	**10^{-6} meters (m)**
1 nanometer (nm)	**=**	**10^{-9} meters (m)**

Know these five metric conversions!



O-H distance =
9.4×10^{-11} m
9.4×10^{-9} cm
9.4×10^{-5} μm
0.094 nm

MAR

UNITS OF MEASUREMENT

We make QUALITATIVE observations of reactions - changes in color and physical state.

We also make QUANTITATIVE MEASUREMENTS, which involve numbers and amount.

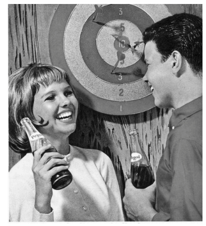
Use SI units - based on the metric system

length	(meter, m)
mass	(kilogram, kg)
time	(second, s)

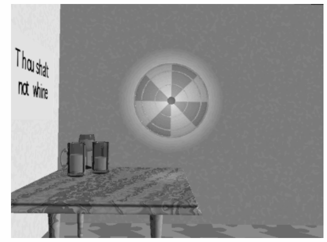
MAR

Accuracy and Precision

Measurements affected by accuracy and precision.



Don't stand next to the dart board, especially with poor precision!



Accurate *and* Precise

MAR

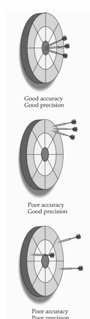
Accuracy versus Precision

Accuracy refers to the proximity of a measurement to the true value of a quantity

Accuracy determined by % error

Precision refers to the proximity (reproducibility) of several measurements to each other.

Determined by average deviation or **parts per thousand**



MAR

Experimental Error

Average deviation:

Step 1: find the absolute value of the difference between each measurement and the average.

Step 2: find the summation of all the deviations and divide by the total number of measurements.

Standard deviation (not used in CH221):

$$\text{Standard deviation} = \sqrt{\frac{\text{sum of squares of deviations}}{(\text{\# of deviations} - 1)}}$$

ppt (parts per thousand):

$$\text{ppt} = \frac{\text{average deviation}}{\text{average}} \times 1000$$

Percent error:

$$\text{\% error} = \frac{\text{experimental value - accepted value}}{\text{accepted value}} \times 100$$

MAR

Experimental Error - *Example*

Trial #	Boiling Point (ºC)	Average (ºC)	Deviations (ºC)	Ave. Dev. (ºC)
1	11.23	11.19	0.04	0.06
2	11.09		0.10	
3	11.27		0.08	
4	11.16		0.03	

Average Deviation = 0.06 ºC (11.19 ± 0.06 ºC)

ppt = (0.06 ºC / 11.19 ºC) x 1000 = 5 ppt

If the literature (accepted) value was 11.25 ºC,

% error = (11.19 ºC – 11.25 ºC / 11.25 ºC) x 100 = -0.5%

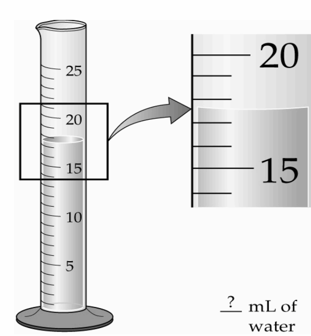
sometimes %error is absolute value (always positive)

MAR

Measurement and Significant Figures

Every experimental measurement, *no matter how precise*, has a degree of uncertainty because there is a limit to the number of digits that can be determined.

Need mathematical system - SIGNIFICANT FIGURES - *very important*, see Chapter One in text and Handout



MAR

Measurement and Significant Figures

To indicate the precision, recorded values should use all the digits known with certainty *plus* one additional estimated digit

Estimated ("doubtful") digit usually considered *uncertain* by plus or minus 1 (± 1)

The total number of digits used to express such a measurement is called the number of *significant figures (sig figs)*.

***Ex:* 65.07 g - four sig figs, 7 "doubtful"**

***Ex:* 54.70318 g - seven sig figs, 8 "doubtful"**

MAR

Rules for Determining Significant Figures

1. **Zeroes in the middle of a number are significant. 69.08 has four sig figs.**
2. **Zeroes at the beginning of a number are not significant. 0.0089 has two sig figs (8 and 9).**
3. **Zeroes at the end of a number and after the decimal point are significant. 2.50 has three sig figs. 25.00 has four sig figs.**
4. **Zeroes at the end of a number and before the decimal point will be significant *only* with a decimal placeholder (period). 1500 has two sig figs, but 1500. has four sig figs.**
5. **Exact conversions (Definitions) have *infinite* sig figs (*ex:* 60 s/1 min, 10 mm/1 cm).**
6. **STUDY! PRACTICE! IMPORTANT!**

MAR

Scientific Notation

Know how your calculator displays scientific notation (and also "regular" notation!)

Always use proper scientific notation when reporting answers in lab, quizzes, etc.

Scientific Notation is a convenient way to write very small or large numbers

Written as a product of a number between 1 and 10, times the number 10 raised to a power. *Examples:*

$$215. = 2.15 \times 10^2$$

Decimal point is moved two places to the left, so exponent is 2.

$$1.56 \times 10^{-8} = 0.000\,000\,015\,6$$

Negative exponent of –8,
so decimal point is moved to the left eight places.

MAR

See: Scientific Notation Handout & Scientific Notation Handout #2

Calculators, Rounding and Sig Figs

Calculators produce large numbers in calculations but the reportable sig figs is usually *much less*! The calculator's large number must be rounded off to a smaller number *keeping only significant figures*.

Once you decide how many numbers to keep *(next slide)*, look at the **first** digit to be dropped:

If the first digit you remove is *between 0 and 4*, drop it and all remaining digits.

If the first digit you remove is *between 5 and 9*, round the number **up** by adding 1 to the digit to the left of the one you drop

Example: 2.4271 becomes 2.4 when rounded to two significant figures

Example: 4.5816 becomes 4.6 when rounded to two significant figures

MAR

Rules for Rounding off Numbers

For multiplication and division: The answer cannot have more significant figures than either of the original numbers.

Three significant figures

$$\frac{278\ \text{mi}}{11.70\ \text{gal}} = 23.8\ \text{mi/gal}$$

Three significant figures

Four significant figures

Actual value: 23.76068…..

MAR

Rules for Rounding off Numbers

For addition and subtraction: The final number must stop at the ***largest doubtful digit***.

Volume of water at start	3.18? ?? L
Volume of water addded	+ 0.013 15 L
Total volume of water	3.19? ?? L

3.18**: 8** is the ***"doubtful digit"***, it stops at the hundredths spot

0.01315**: 5** is the ***"doubtful digit"***, it stops at the hundred thousandths spot

Final answer is 3.19 L:
Answer stops at largest "doubtful digit" *(hundredths vs. hundredth thousandths)*

Actual value: 3.19315

MAR

Density: *the* **ratio** *of a substance's* **mass** *(grams)* **to its volume** *(mL, cm³)*

Substances layer themselves according to their density: **superposition**

Density used to separate materials

$d = 19.3$ g/cm³

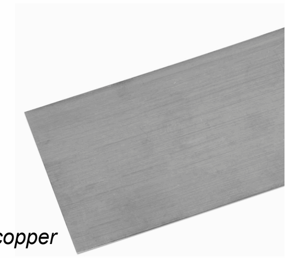
$d = 13.6$ g/cm³

MAR

Density Problem

Problem: A piece of copper has a mass of 57.54 g. It is 9.36 cm long, 7.23 cm wide, and 0.95 mm thick. Calculate density (g/cm³).

$$\text{Density} = \frac{\text{mass (g)}}{\text{volume (cm}^3\text{)}}$$



Sheet of copper

MAR

Density Problem

Problem: A piece of copper has a mass of 57.54 g. It is 9.36 cm long, 7.23 cm wide, and 0.95 mm thick. Calculate density (g/cm³).

SOLUTION

1. Get dimensions in common units.

$$0.95 \text{ mm} \cdot \frac{1 \text{ cm}}{10 \text{ mm}} = 0.095 \text{ cm}$$

2. Calculate volume in cubic centimeters.

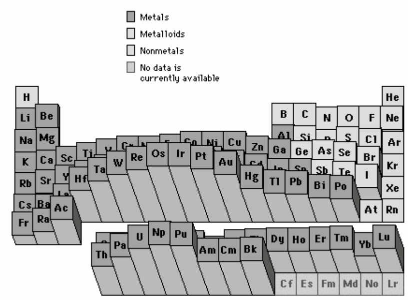
$$(9.36 \text{ cm})(7.23 \text{ cm})(0.095 \text{ cm}) = \mathbf{6.4 \text{ cm}^3}$$

6.42891...

3. Calculate the density.

MAR

Relative Densities of the Elements



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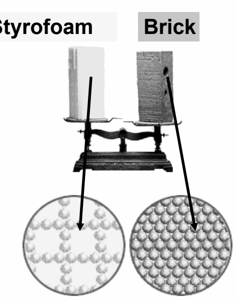
DENSITY

Notice that density is an INTENSIVE property of matter

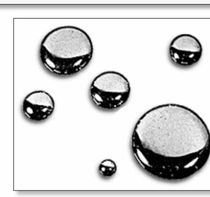
INTENSIVE - does not depend on quantity of matter (density, boiling point, etc.)

Contrast with EXTENSIVE - depends on quantity of matter. Examples include mass and volume



MAR

PROBLEM: Mercury (Hg) has a density of 13.6 g/cm³. What is the mass of 95 mL of Hg? (454 g = 1 lb)



Solve the problem using DIMENSIONAL ANALYSIS ***- see the <u>Dimensional Analysis</u> and <u>Factor Label</u> handouts on the Web***

MAR

PROBLEM: Mercury (Hg) has a density of 13.6 g/cm³. What is the mass of 95 mL of Hg? (454 g = 1 lb)

First, note that 1 cm³ = 1 mL

Then, use dimensional analysis to calculate mass.

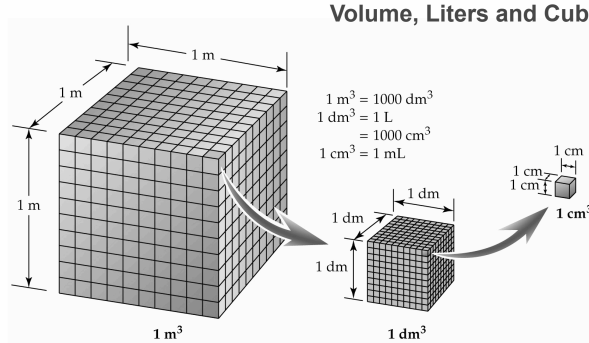
$$95\ \text{cm}^3 \cdot \frac{13.6\ \text{g}}{1\ \text{cm}^3} = 1.3 \times 10^3\ \text{g}$$ *1292*

What is the mass in pounds?

$$1.3 \times 10^3\ \text{g} \cdot \frac{1\ \text{pound}}{454\ \text{g}} = 2.9\ \text{lb}$$ *2.863436...*

MAR

Volume, Liters and Cubed Lengths



cm³ = cm * cm * cm, *etc.*

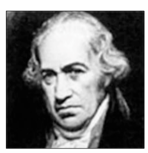
MAR

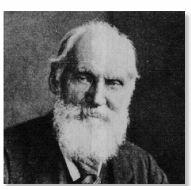
Temperature Scales

- **Fahrenheit (°F)**
- **Celsius (°C)**
- **Kelvin (K)**

Anders Celsius
1701-1744

Daniel Fahrenheit
1686-1736





Lord Kelvin
(William Thomson)
1824-1907

MAR

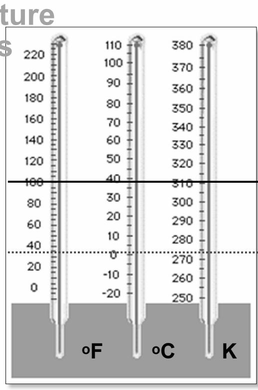
Temperature Scales

	Fahrenheit	Celsius	Kelvin
Boiling Point of water	212	100	373.15
	180°	100°	100°
Freezing Point of water	32	0	273.15

Notice that 1 Kelvin degree = 1 degree Celsius
Difference between Celsius temperatures and Kelvin temperatures the same!

MAR

Temperature Scales



100 ˚F = 38 ˚C = 311 K

32 ˚F = 0 ˚C = 273 K

Note:
no negative K!
no degree K symbol!

MAR

Calculations Using Temperature

Chemistry experiments recorded in Celcius (°C) but calculations *generally* require temperatures in Kelvins (K)

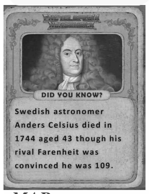
$$T\ (\text{K}) = T\ (^\circ\text{C}) + 273.15$$

Body temp = 37.0 ˚C + 273.15 = 310.2 K *310.15*

Liquid nitrogen = 77 K - 273.15 = -196 ˚C *-196.15*

Memorize 273.15!

MAR



Calculations Using Temperature

Occasionally need Fahrenheit (F) values

Convert using Celsius scale

T (˚F) = $^9/_5$ T (˚C) + 32.00

Liquid He = 4.2 K - 273.15 = -269.0 ˚C *-268.95*

T (˚F) = $^9/_5$ (-269.0 ˚C) + 32.00 = -452.2 ˚F *-452.2*

MAR

Mass Percentages in Chemistry

Often see "30% lead, 70% oxygen"

This means that in 100 grams of the substance

- **30 grams will be lead**
- **70 grams will be oxygen**

In one gram of the substance,

- **0.30 grams will be lead**
- **0.70 grams will be oxygen**

MAR

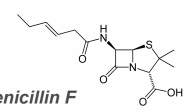
Mass Percentages in Chemistry

Example: **Penicillin F is 53.829% carbon. How much carbon in 75 g of Penicillin F?**

Solution

75 g Penicillin F* (53.829 g C / 100 g Penicillin F) =
40. g carbon *(40.37175)*

Note **that volume percentages occasionally seen, but not often in our class**



Penicillin F

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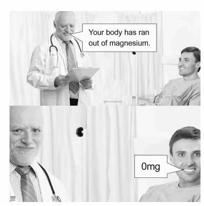
End of Chapter One

See also:

- *Chapter One Study Guide*
- *Chapter One Concept Guide*
- *Math ("Chapter Guide Zero") Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*



"Just a darn minute! Yesterday you said X equals two!"

MAR

Important Equations, Constants, and Handouts from this Chapter:

metric prefixes:
nano (n) = 10^{-9}
micro (μ) = 10^{-6}
milli (m) = 10^{-9}
centi (c) = 10^{-9}
kilo (k) = 10^{-9}

$$\text{Density} = \frac{\text{mass (g)}}{\text{volume (cm3)}}$$

T (K) = T (˚C) + 273.15

1 cm^3 = 1 mL

significant figures!!!

mass percentages

MAR

End of Chapter Problems: ***Test Yourself***

1. 32.32 - 23.2 = _____
2. 32.4 * 37.31 = _____
3. 4.311 / 0.07 = _____
4. Convert 37.0 C to K.
5. Convert 253.6 mL to cm^3.
6. Convert 24 m^3 to cm^3.
7. 235.05 + 19.6 + 2 = _____
8. 58.925 - 19 = _____
9. 2.19 x 4.2 = _____
10. 4.311 ÷ 0.07 = _____
11. The platinum-containing cancer drug cisplatin contains 65.0% platinum. If you have 1.53 g of the compound, what mass of platinum (in grams) is contained in this sample?
12. The anesthetic procaine hydrochloride is often used to deaden pain during dental surgery. The compound is packaged as a 10.% solution (by mass; d = 1.0 g/mL) in water. If your dentist injects 0.50 mL of the solution, what mass of procaine hydrochloride (in milligrams) is injected?

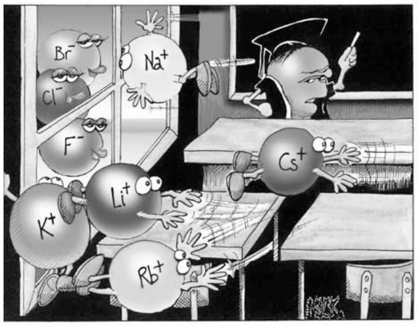
MAR

End of Chapter Problems: *Answers*

1. 9.1
2. 1210
3. 60
4. 310.2 K
5. 253.6 cm^3
6. 2.4 x 10^7 cm^3
7. 257
8. 40.
9. 9.2
10. 60
11. 0.995 g Pt
12. 50. mg

MAR

Atoms, Molecules and Ions



"Perhaps one of you gentlemen would mind telling me just what it is outside the window that you find so attractive...?"

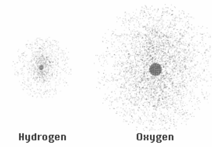
Chapter 2 and Chapter 3 (3.1)

"Chapter 2 Part 1"

Chemistry 221
Professor Michael Russell

Last update: 4/29/24

Atoms and Elements

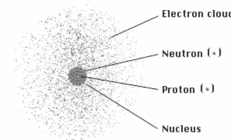


Elements: the building blocks of Nature

Atoms: the smallest pieces of an element

Atoms contain protons, neutrons and electrons

Protons and neutrons in the nucleus



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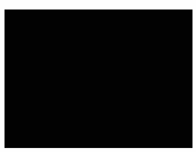
Where Does Matter Come From?

FROM THE



The universe is 13.77 billion years old

Hydrogen and Helium important

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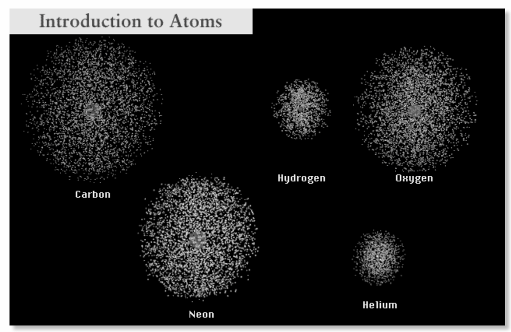

Also Carbon, Oxygen and Neon

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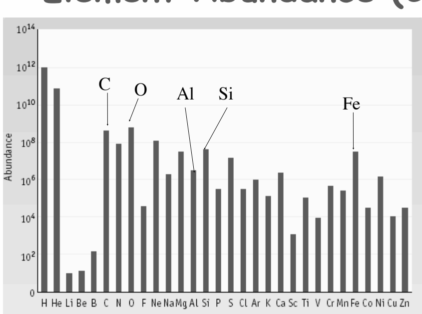
Element Abundance (Cosmic)



Most abundant elements in the Earth's

- **Crust**: Silicon ***and*** oxygen
- **Atmosphere**: nitrogen ***and*** oxygen

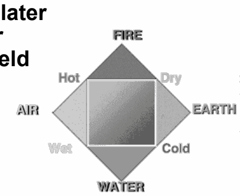
http://www.webelements.com/

MAR

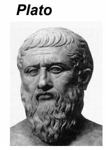
Early Models of the Atom - Empedocles & Aristotle

Empedocles' theory of matter (later modified by Aristotle) with four "elements" (coined by Plato) held for approximately 2000 years!

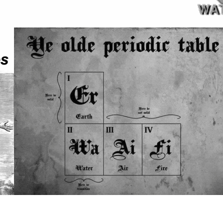
Element "mixtures" produce "properties" of hot, wet, etc.



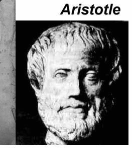
Plato



Empedocles



Aristotle



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Early Models of the Atom - Democritus

DEMOCRITUS (460 - 370 BCE)
was a contemporary of Plato

Atoms have *structure* and *volume*

"Gold can be divided into smaller pieces only so far before the pieces no longer retain the properties of gold"

Smallest unit of matter = *atomos*, atoms

MAR

JOHN DALTON

(1766 - 1844)

The "Newton" of Chemistry

1804 - Proposed Atomic Theory

"Atoms cannot be created or destroyed"

"Atoms of one element are different from other element's atoms" - proposed *relative* scale of atomic masses (now the *amu*)

"Chemical change involves bond breaking, bond making and rearrangement of atoms"

Did not include Democritus' ideas that atoms have structure

1 atomic mass unit (amu) **= 1.66054 * 10^{-24} g =** *1 Dalton (Da)*

MAR

The Discovery of Atomic Structure: Electricity

BEN FRANKLIN:

Key Theories:

- **+ and - charges**
- **Opposites attract, like repel**
- **Charge is *conserved***
- **Force inversely proportional to distance**

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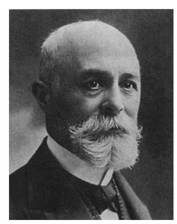


Radioactivity

Henri Becquerel (1896) discovered radioactivity while studying uranium ore

Emits new kind of "ray"

Rays pass unimpeded through many objects

Rays produce image on photographic plate (silver emulsion)

But MARIE CURIE opened the door...

MAR

MARIE CURIE

the "Newton of Radioactivity"

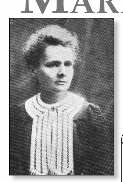
Substances disintegrated upon emission of rays - *radioactive*

Challenged Dalton's idea on "indestructible atoms" - more comprehensive theory

MAR

MARIE CURIE

the "Newton of Radioactivity"



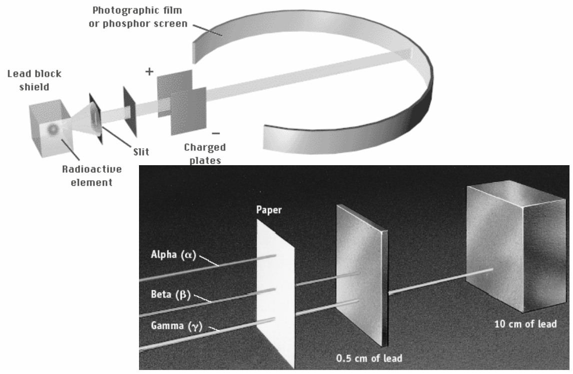
She (and *Rutherford*) found three types of radiative processes:

alpha - a helium cation - α
beta - supercharged electrons - β
gamma - high energy emission - γ

***Note* that α and β are massive and charged, but γ radiation has no charge or mass**

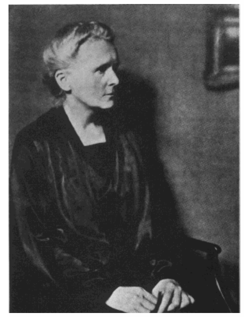
MAR

Discovering the Radioactive Particles



MAR

Marie Curie



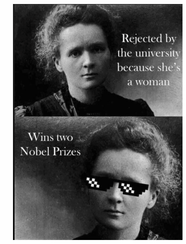
1903 - discovered radium, polonium

1911 - isolated pure radium (bought her own samples!)

1919 - American Association of University Women raised $150K for 1 g of radium, continued work

1934 - died of leukemia *killed by her work*

"Nothing in life is to be feared. It is only to be understood."



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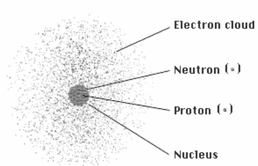
ATOM COMPOSITION



The atom is mostly empty space

protons & neutrons in nucleus

Atom electrically neutral *if* the # protons = # electrons

electrons in space around nucleus

***Extremely small!* One teaspoon of water has 3 times as many atoms as the Atlantic Ocean has teaspoons of water.**

MAR

ATOMIC COMPOSITION

(Three Particles Handout)

Protons
- positive electrical charge
- mass = 1.672623×10^{-24} g
- relative mass = 1.0073 atomic mass units (amu) where 1 amu = $1.66054 * 10^{-24}$ g

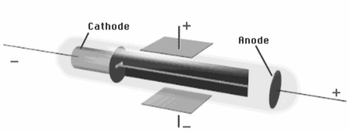
Electrons
- negative electrical charge
- relative mass = 0.0005486 amu

Neutrons
- no electrical charge
- mass = 1.0087 amu

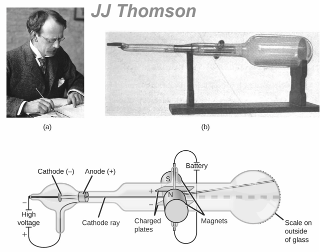
MAR

ELECTRONS

Charge to mass ratio of the electron discovered in 1897 by JJ Thomson using Cathode Ray Tubes (CRT)



JJ Thomson



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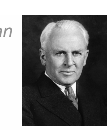
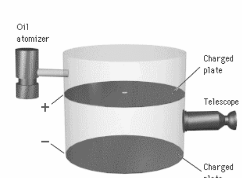
ELECTRONS

Robert Millikan



Charge to mass ratio of the electron discovered in 1897 by JJ Thomson using Cathode Ray Tubes (CRT)

Robert Millikan discovered the mass of the electron in 1913



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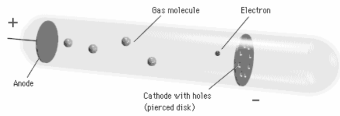
PROTONS

Discovered in 1919 by Rutherford while using canal ray tubes and hydrogen gas



1,837 times more massive than electron
Opposite charge (same magnitude) as electron

MAR

NEUTRONS

Most difficult particle to discover - no charge, no voltage/magnet tests

Chadwick detected neutrons in 1932
n more massive than p or e,
used mass spectrometer

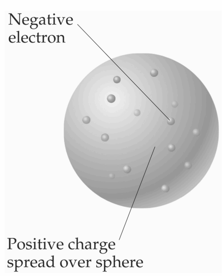
MAR

THE ATOM: *Plum Pudding Model*

JJ Thomson (discoverer of the electron) proposed the "plum pudding" model of the atom in 1904:

Large volume, negative "spheres" in a positive "cloud" of low density

Rutherford proposed the correct model



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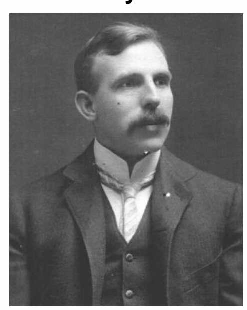
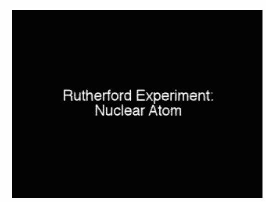
The modern view of the atom was developed by Ernest Rutherford in 1910.





Low density atom with a highly dense, positively charged nucleus

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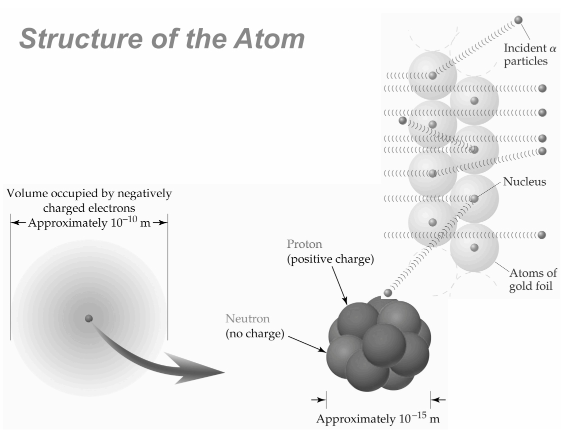
Structure of the Atom



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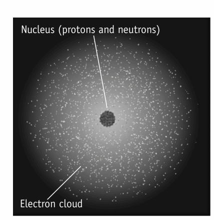
THE ATOM: *Summary*

Protons and neutrons in nucleus; electrons circle outside

Most of the mass of an atom is in the nucleus; electrons have ~0.05% mass

Nucleus very dense; most of atom's volume *empty*

Atom electrically neutral *if* the # protons = # electrons



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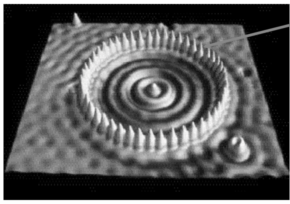
How Large is an Atom?



Circle consists of 48 Fe atoms

Radius of circle is 71 Angstroms where
1 Å = 10^{-10} m

STM image of "quantum corral" of iron atoms

See http://www.almaden.ibm.com/vis/stm for STM or Scanning Tunneling Microscopic images of atoms.

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Atomic Number, Z

All atoms of the same element have the same number of protons in the nucleus, Z.

13	← atomic number
Al	← symbol
26.9815	← atomic weight

Z distinguishes atoms from one another!

MAR

Atomic Number, Z

All atoms of the same element have the same number of protons in the nucleus, Z.

13	← atomic number
Al	← symbol
26.9815	← atomic weight

Henry Moseley determined the atomic number for each element in 1914 using x-ray scattering

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Mass Number, A

Mass Number, A

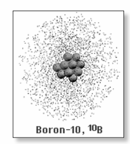
A usually in units of amu

A = # protons + # neutrons

A boron atom can have

A = 5 p + 5 n = 10 amu

Method to display A, Z and element symbol: $\begin{matrix} A \rightarrow \\ Z \rightarrow \end{matrix} {}^{10}_{5}\mathrm{B}$



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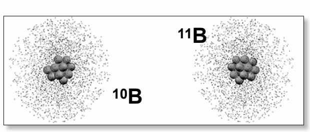
Isotopes

Atoms of the same element (same Z) but different mass number (A).

Boron-10 (^{10}B) has 5 p and 5 n

Boron-11 (^{11}B) has 5 p and 6 n



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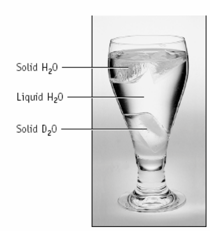
Hydrogen Isotopes

Hydrogen has *three* isotopes

${}^{1}_{1}\mathrm{H}$ 1 proton and 0 neutrons, protium

${}^{2}_{1}\mathrm{H}$ 1 proton and 1 neutron, deuterium

${}^{3}_{1}\mathrm{H}$ 1 proton and 2 neutrons, tritium radioactive



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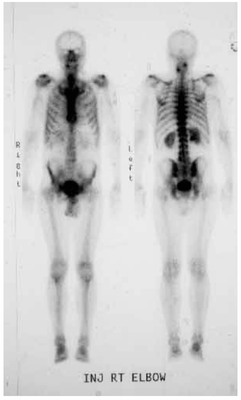

Isotopes & Their Uses

Bone scans with radioactive technetium-99.

$^{99}_{43}Tc$

Emits gamma rays

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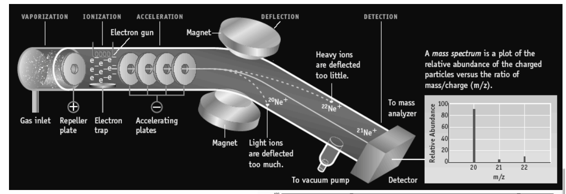
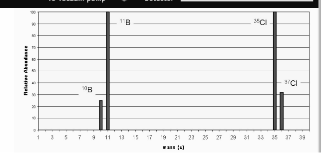
Masses of Isotopes

determined with a mass spectrometer





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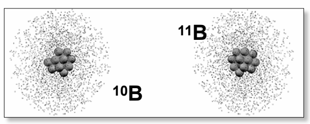
Isotopes



Because of the existence of isotopes, the mass of a collection of atoms has an average value.

Average mass = ATOMIC WEIGHT

Boron is 20% ^{10}B and 80% ^{11}B. That is, ^{11}B is 80 percent abundant on earth.

To calculate the atomic weight for boron:

= (abundance$_1$ * mass$_1$) + (abundance$_2$ * mass$_2$)

= 0.20 (10 amu) + 0.80 (11 amu) = 10.8 amu

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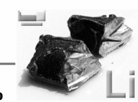
Isotopes & Atomic Weight

Because of the existence of isotopes, the mass of a *collection* of atoms has an average value.

^{6}Li = 7.5% abundant and ^{7}Li = 92.5%



Atomic weight of Li = ______________

^{28}Si = 92.23%, ^{29}Si = 4.67%, ^{30}Si = 3.10%

Atomic weight of Si = ______________

MAR

Isotopes

Example: **Nitrogen has two main isotopes, ^{14}N (14.0031 amu, 99.6299%) and ^{15}N (15.0001 amu, 0.3701%). Calculate the average atomic mass.**

Solution

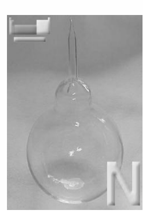
Average atomic mass =

= (abundance$_1$ * mass$_1$) + (abundance$_2$ * mass$_2$)

= (0.996299*14.0031) + (0.003701*15.0001)

= 13.9512745 + 0.05551537

= 14.0068 amu



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Isotopes

Example: Gallium **has two main isotopes, ^{69}Ga (68.9257 amu) and ^{71}Ga (70.9249 amu) with an average atomic mass of 69.723. Calculate the % abundance of each isotope.**

Solution

Average atomic mass =

69.723 = *x*(^{69}Ga)*68.9257 + *y*(^{71}Ga)*70.9249

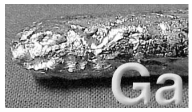
but also

1 = *x*(^{69}Ga) + *y*(^{71}Ga) *(2 percentages equal 100%)*

so y(^{71}Ga) = 1 - x(^{69}Ga)

MAR

Isotopes



Example: **Gallium has two main isotopes, ^{69}Ga (68.9257 amu) and ^{71}Ga (70.9249 amu) with an average atomic mass of 69.723. Calculate the % abundance of each isotope.**

Solution

$69.723 = x(^{69}Ga) * 68.9257 + y(^{71}Ga) * 70.9249$, *or*

$69.723 = x * 68.9257 + (1 - x) * 70.9249$

Solve for x, get:

$x(^{69}Ga) = 0.6012$ (60.12%)

$y(^{71}Ga) = 1 - x = 0.3988$ (39.88%)

MAR

Isotopes

Antimony has two main isotopes:

^{121}Sb (120.9038 amu, 57.20%) *and*

^{123}Sb (122.9042 amu, 42.80%)

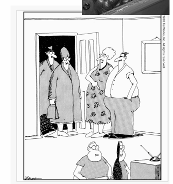
Average atomic mass of Sb: **121.760**

Will you have <u>one atom</u> of antimony with 121.760 amu?

No!

One atom of antimony will have a mass of 120.9038 amu 57.20% of the time

One atom of antimony will have a mass of 122.9042 amu 42.80% of the time

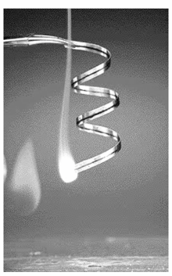


"Bob and Ruth! Come on in. ... Have you met Russell and Bill, our 1.5 children?"

Average kids per family in Oregon: 1.7-1.8 *(2019)*

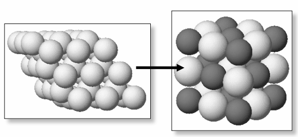
MAR

Counting Atoms



Mg burns in air (O_2) to produce white magnesium oxide, MgO.

How can we figure out how much oxide is produced from a given mass of Mg?



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Counting Atoms

Chemistry is a quantitative science - we need a "counting unit."

The MOLE!

A mole is similar to a *dozen* - you can have a dozen roses, a dozen donuts - you can also have a mole of roses, or a mole of donuts

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Particles in a Mole

Avogadro's Number (N_A), *named for* **Amedeo Avogadro,** *1776-1856*

$6.02214076 \times 10^{23}$

A mole is the amount of *any* substance containing 6.022×10^{23} particles

$$\frac{6.022 \times 10^{23} \text{ Cu atoms}}{1 \text{ mole Cu}} \qquad \frac{1 \text{ mole } CO_2}{6.022 \times 10^{23} \text{ molecules } CO_2}$$

Memorize* 6.022×10^{23}! *Always use this value for N_A!

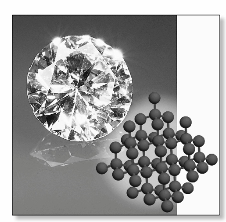
MAR

Molar Mass

1 mol of ^{12}C
= 12.00 g of C
= 6.022×10^{23} atoms of C

12.00 g of ^{12}C is its MOLAR MASS

Taking into account all of the isotopes of C, the molar mass of C is 12.011 g/mol



Try to use at least four sig figs for molar mass

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Molar Mass

1 mol of ^{12}C = 12.00 g of C = 6.022×10^{23} atoms of C

12.00 g of ^{12}C is its MOLAR MASS

Taking into account all of the isotopes of C, the molar mass of C is 12.011 g/mol

Find molar mass from periodic table

13	← atomic number
Al	← symbol
26.9815	← atomic weight

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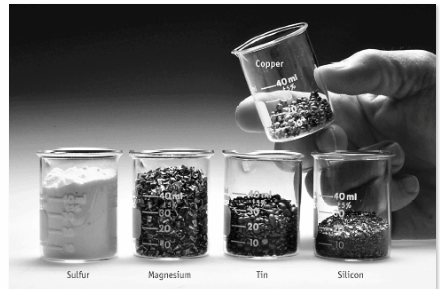
One mole Amounts



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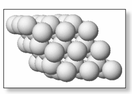

PROBLEM: What amount of Mg is represented by 0.200 g? How many atoms?

Mg has a molar mass of 24.3050 g/mol.

$$0.200 \text{ g} \bullet \frac{1 \text{ mol}}{24.31 \text{ g}} = 8.23 \times 10^{-3} \text{ mol}$$

0.00822706...

How many atoms in this piece of Mg?

$$8.23 \times 10^{-3} \text{ mol} \bullet \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}}$$

$= 4.96 \times 10^{21}$ atoms Mg

4.956106....E21

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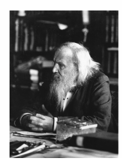
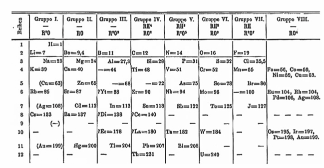
Periodic Table

Dmitri Mendeleev developed the modern periodic table. Argued that element properties are *periodic functions of their atomic weights.*

We now know that element properties are periodic functions of their ATOMIC NUMBERS.



Dmitri Mendeleev



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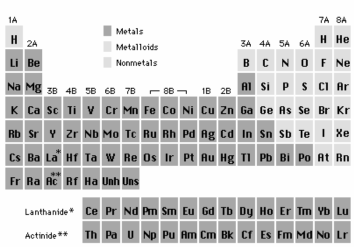
Periods in the Periodic Table



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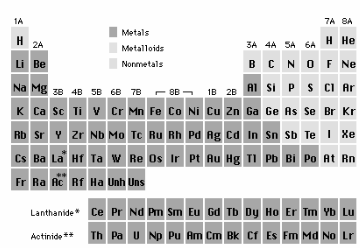
Groups in the Periodic Table



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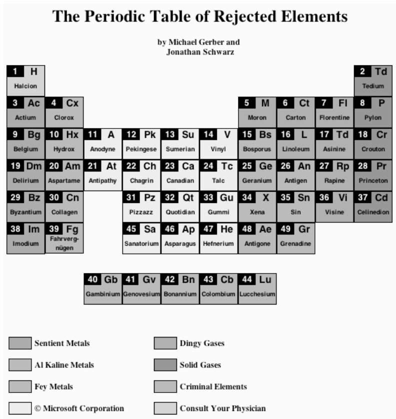
Rejected Elements in the Periodic Table



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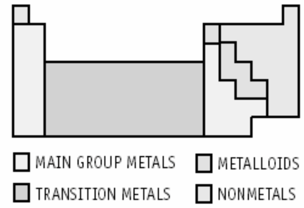
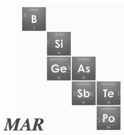
Periodic Table

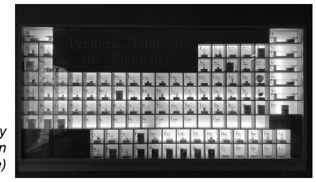
Periodic Table has the following:

- A groups: **main group elements**
- B groups: **transition metals**
- Lanthanides
- Actinides
- metals
- nonmetals
- metalloids





Common Metalloids



Periodic Table Display at University of Oregon (Eugene)

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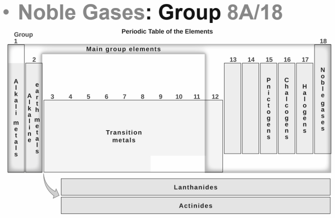
- Alkali Metals: **Group** 1A/1
- Alkaline Earth Metals: **Group** 2A/2
- Icosagens ("Twisted Metals"): **Group** 3A/13
- Crystallogens: **Group** 4A/14
- Pnictogens **("to choke"): Group** 5A/15
- Chalcogens **("chalk formers"): Group** 6A/16
- Halogens **("salt formers"): Group** 7A/17
- Noble Gases: **Group** 8A/18



Periodic Table

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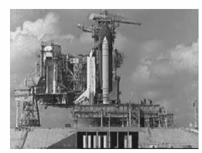
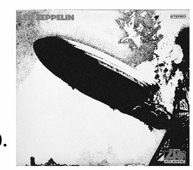
Hydrogen



Shuttle main engines use H_2 and O_2



The Hindenburg crash, May 1939.

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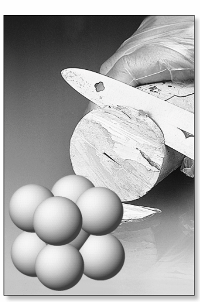
Group 1A: Alkali Metals: Li, Na, K, Rb, Cs, Fr



Extreme reactivity with water!

Sodium cut with a knife

Solids at room temperature, violently react with water

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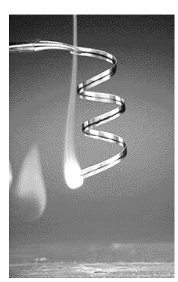
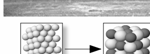
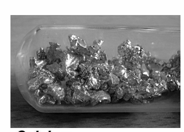
Group 2A: Alkaline Earth Metals Be, Mg, Ca, Sr, Ba, Ra



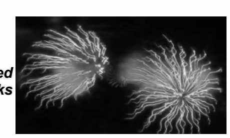


MgO



Calcium

Ba gives green fireworks



Sr gives red fireworks

Alkaline Earth Metals occur naturally only in compounds (except Be)

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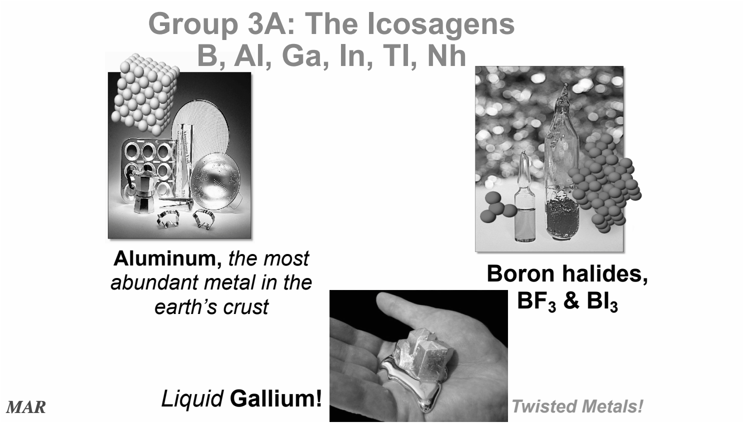
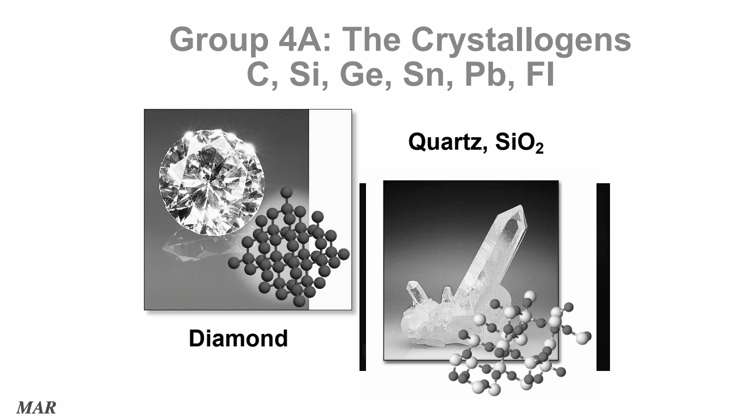
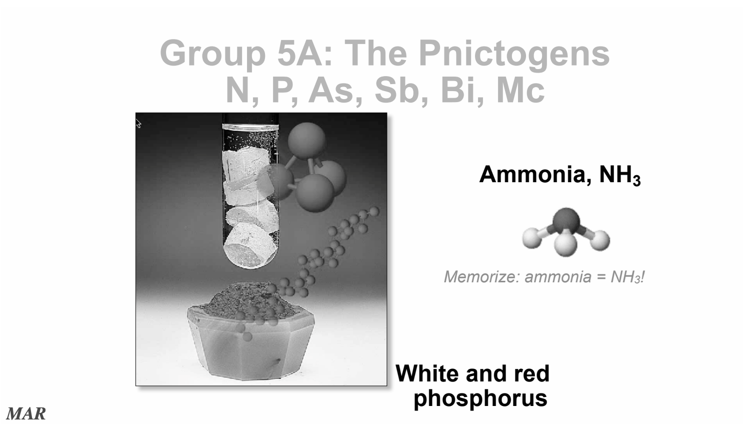
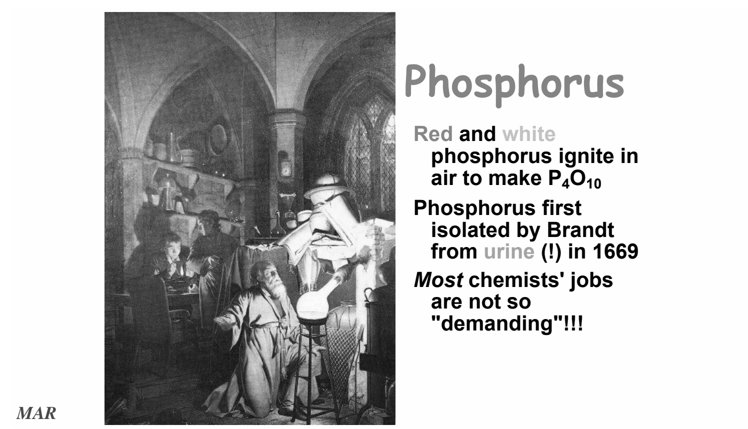
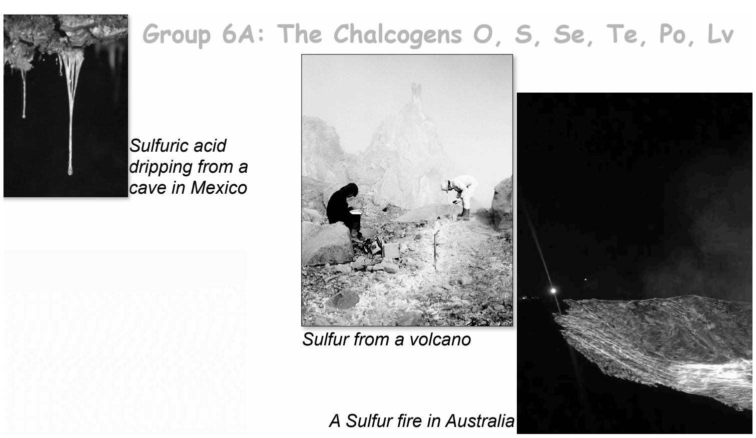
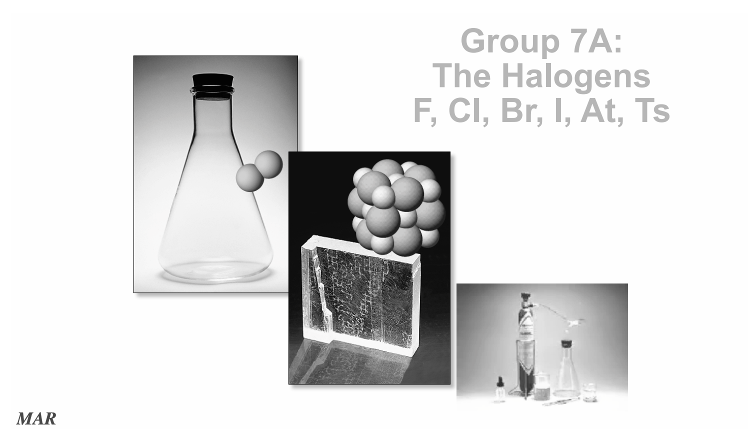

Group 3A: The Icosagens
B, Al, Ga, In, Tl, Nh
Aluminum, the most abundant metal in the earth's crust
Boron halides, BF_3 & BI_3
Liquid Gallium!
Twisted Metals!
MAR


Group 4A: The Crystallogens
C, Si, Ge, Sn, Pb, Fl
Quartz, SiO_2
Diamond
MAR


Group 5A: The Pnictogens
N, P, As, Sb, Bi, Mc
Ammonia, NH_3
Memorize: ammonia = NH_3!
White and red phosphorus
MAR

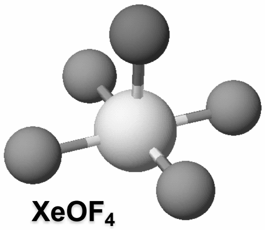

Phosphorus
Red and white phosphorus ignite in air to make P_4O_{10}
Phosphorus first isolated by Brandt from urine (!) in 1669
Most chemists' jobs are not so "demanding"!!!
MAR


Group 6A: The Chalcogens O, S, Se, Te, Po, Lv
Sulfuric acid dripping from a cave in Mexico
Sulfur from a volcano
A Sulfur fire in Australia
MAR

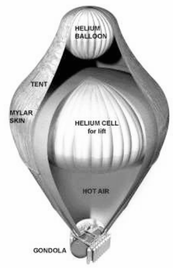

Group 7A:
The Halogens
F, Cl, Br, I, At, Ts

Group 8A: Noble Gases
He, Ne, Ar, Kr, Xe, Rn, Og

All gases at room temperature; considered unreactive until 1962

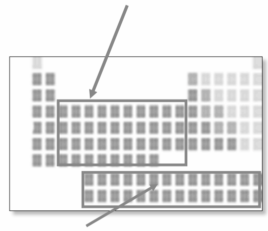


$XeOF_4$

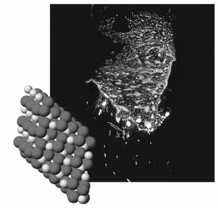


MAR

Transition Metals



Lanthanides and actinides



Iron in air gives iron(III) oxide

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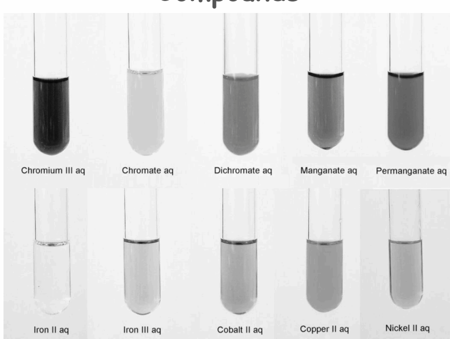
Colors of Transition Metal Compounds



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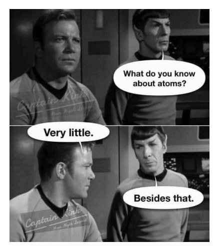
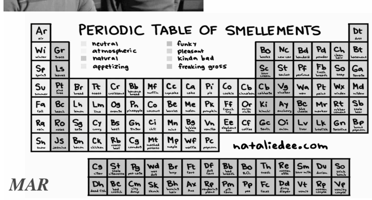
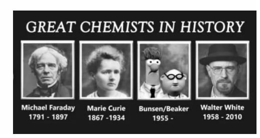
End of Chapter 2 Part 1



See also:

- *Chapter Two Part 1 Study Guide*
- *Chapter Two Part 1 Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*





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Important Equations, Constants, and Handouts from this Chapter:

- **alpha, beta, gamma radiation**
- **the "gold foil experiment"**
- **protons, neutrons, electrons**
- **mass number, atomic number**
- **isotopes**
- **atomic weight and molar mass**
- **Avogadro's number**

$$\begin{array}{l} A \longrightarrow 10 \\ Z \longrightarrow \ \ 5 \end{array} B$$

A mole is the amount of *any* substance containing 6.022 x 10^{23} particles

Periodic table: groups, periods, metals, metalloids, nonmetals, alkali, alkaline earth, halogens, noble gases, transition metals, lanthanides, actinides, *how to find the molar mass of an element!*

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End of Chapter Problems: ***Test Yourself***

1. How many protons in a magnesium atom with 15 neutrons? What is the mass number of this isotope?
2. How many neutrons in: $^{60}_{27}Co$
3. Thallium has two stable isotopes, ^{203}Tl and ^{205}Tl. Knowing that the atomic weight of thallium is 204.4, which isotope is the more abundant of the two?
4. Gallium has two naturally occurring isotopes, ^{69}Ga and ^{71}Ga, with masses of 68.9257 u and 70.9249 u, respectively. Calculate the percent abundances of these isotopes of gallium.
5. Calculate the mass in grams of 2.5 mol of aluminum.
6. Calculate the amount (moles) represented by 0.012 mol Li. How many atoms of Li are present?
7. A cylindrical piece of sodium is 12.00 cm long and has a diameter of 4.5 cm. The density of sodium is 0.971 g/cm³. How many atoms does the piece of sodium contain? (The volume of a cylinder is V = π x r² x length.)
8. In the following list, tell which element is: a metalloid, a transition metal, a halogen, a noble gas, a lanthanide, an alkali metal: Gd, Se, Cs, W, Xe, Cl

MAR

End of Chapter Problems: *Answers*

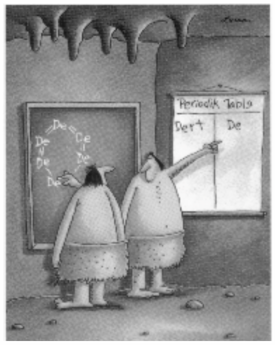
1. 12 protons, mass number = 27
2. 33 neutrons
3. 205
4. ^{69}Ga abundance is 60.12%, ^{71}Ga abundance is 39.88%
5. 68 g Al
6. 1.7×10^{-3} mol Li, 1.0×10^{21} atoms Li
7. 4.9×10^{24} atoms Na
8. A metalloid (Se), a transition metal (W), a halogen (Cl), a noble gas (Xe), a lanthanide (Gd), an alkali metal (Cs)

MAR

Atoms, Molecules and Ions

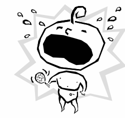
Chapter 2 and Chapter 3 (3.1, 3.2)
"Chapter 2 Part 2"

Chemistry 221
Professor Michael Russell


Early chemists describe the first dirt molecule

MAR Last update: 4/29/24

Poor Auntie Jane!



Auntie Jane fed Baby Nell
What she thought was calomel
What the baby really ate
was Corrosive Sublimate
Not much difference,
I confess,
Just one chlorine more
and one baby less!

calomel = **HgCl** ***(for dysentery)***
Corrosive Sublimate = **$HgCl_2$**

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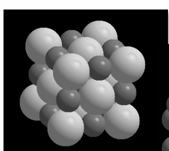
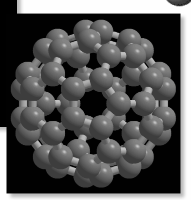
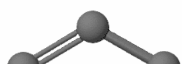
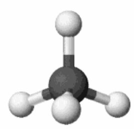
Molecules, Ions and Compounds


NaCl, salt


Buckyball, C_{60}


Ozone, O_3


Methane, CH_4

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Compounds and Molecules

COMPOUNDS are a combination of 2 or more elements in definite ratios by mass.

The character of each element is lost when forming a compound.

MOLECULES are the smallest unit of a compound that retains the characteristics of the compound.

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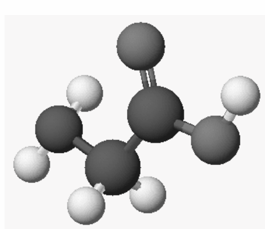
MOLECULAR FORMULA

Formula for glycine is $C_2H_5NO_2$
In one molecule there are
- **2 C atoms**
- **5 H atoms**
- **1 N atom**
- **2 O atoms**



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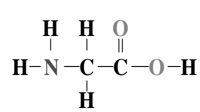
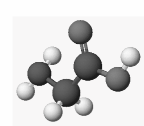
Writing Formulas

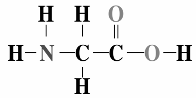
Can also write glycine formula ($C_2H_5NO_2$) as
H_2NCH_2COOH
to show atom ordering
or in the form of a structural formula



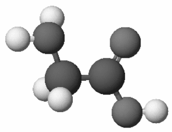


MAR *structural formulas also called "condensed" formulas*

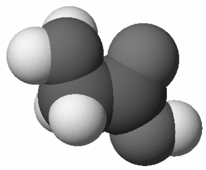
Molecular Modeling



Drawing of glycine



Ball & stick



Space-filling

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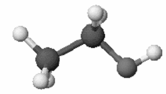
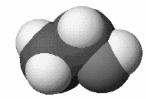
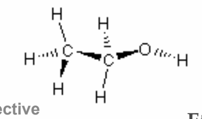
3D Perspectives: Ethanol



Ball and Stick



Space Filling



Perspective

Ethanol = CH_3CH_2OH

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Comparison of Formula Types

Compound	Molecular	Empirical	Structural
Water	H_2O	H_2O	HOH
Hydrogen Peroxide	H_2O_2	HO	HOOH
Ethylene	C_2H_4	CH_2	H_2CCH_2
Ethane	C_2H_6	CH_3	H_3CCH_3
Ethanol	C_2H_6O	C_2H_6O	H_3CCH_2OH
Dimethyl ether	C_2H_6O	C_2H_6O	H_3COCH_3

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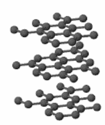
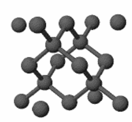
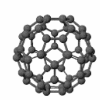
Allotropes of Elements



graphite



diamond



Buckyball (C_{60}) or Buckminsterfullerene

Most elements exist as individual atoms - ***monotomic***

Allotropes are **different versions of the same element**

Carbon *exists naturally as* **graphite, diamond and buckyballs.**

Seven elements exist as *diatomics* *(next slide)*

MAR

Also carbon graphene

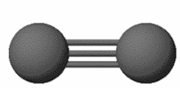
The Seven Diatomic Elements

Have
No
Fear
Of
Ice
Clear
Brew

"HONCl BrIF"

Most elements are monatomic - they exist as individual atoms - but there are these seven exceptions:

Oxygen (O_2) generated on a leaf



Nitrogen, N_2

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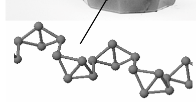
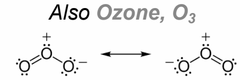
OTHER ELEMENTS THAT EXIST AS POLYATOMIC ELEMENTS

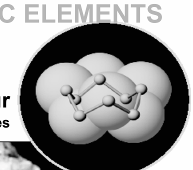


White P_4 and polymeric red phosphorus

Also Ozone, O_3



S_8 sulfur molecules



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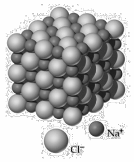
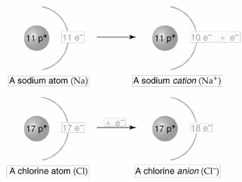
IONS AND IONIC COMPOUNDS

IONS are atoms or groups of atoms with a positive or negative charge.

Taking away electron(s) creates a CATION with a positive charge

Adding electron(s) creates an ANION with a negative charge.





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IONS AND IONIC COMPOUNDS

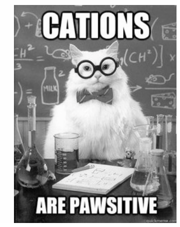
CATIONS have protons > electrons

ANIONS have electrons > protons

Remember:

CATS have PAWS

CATions are PAWSitive



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Formation of Cations and Anions

Mg --> Mg^{2+} + 2 e-

A cation forms when an atom loses one or more electrons.

F + e- --> F^-

An anion forms when an atom gains one or more electrons

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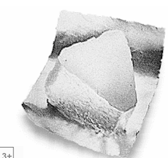
Charges on Metals



Fixed charge metals include:
- *Groups IA, IIA & "stairs" *(next slide)*
- ****charge** = group number *(mostly)***

Na^+	sodium ion
Mg^{2+}	magnesium ion
Al^{3+}	aluminum ion
Ag^+	silver ion

		Al 3+ 13
	Zn 2+ 30	Ga 3+ 31
Ag 1+ 47	Cd 2+ 48	In 3+ 49

the stairs

All other metals ("*variable charge*" metals) --> use Roman number to represent charge on metal

Fe^{2+}	iron(II) ion
Fe^{3+}	iron(III) ion
V^{3+}	vanadium(III) ion

No ferrous or ferric nomenclature!

MAR

The Fixed Charge "Stairs" Metals

- Start with Al (which is +3)
- Go backwards down the "stairs"
- Decrease charge after each "stair"

		Al 3+ 13
	Zn 2+ 30	Ga 3+ 31
Ag 1+ 47	Cd 2+ 48	In 3+ 49

These, and Groups IA and IIA, are the "fixed charge metals", and we always know their ionic charge

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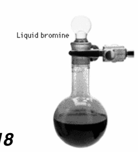
NONMETALS



NONMETAL + n e- ------> X^{n-}
where charge = Group no. - 8

webelements.com: group no. - 18

Group 4A	Group 5A	Group 6A	Group 7A
C^{4-}, carbide	N^{3-}, nitride	O^{2-}, oxide	F^-, fluoride
		S^{2-}, sulfide	Cl^-, chloride
			Br^-, bromide
			I^-, iodide

Name derived by adding -ide to stem

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A "Quick and Dirty" Guide to Ionic Charges

Groups IA, IIA or "the stairs": *fixed charge metals*
Charge = **positive**
Magnitude = **group #** ***mostly!***

Groups VA, VIA or VIIA: *fixed charge nonmetals*
Charge = **negative**
Charge = **group # - 8**

All Other Metals: **Difficult to predict,** use **Roman number** to represent positive charge, *these are the* "Variable Charge metals"

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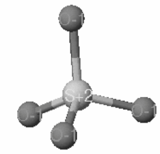
POLYATOMIC IONS



Groups of atoms with a charge.
MEMORIZE the names and formulas in your text and the "Nomenclature" lab.

Charge	Formula	Name	Formula	Name
1–	H^-	Hydride ion	**CH_3COO^-** (or $C_2H_3O_2^-$)	**Acetate ion**
	F^-	**Fluoride ion**	ClO_3^-	Chlorate ion
	Cl^-	**Chloride ion**	**ClO_4^-**	**Perchlorate ion**
	Br^-	**Bromide ion**	**NO_3^-**	**Nitrate ion**
	I^-	**Iodide ion**	MnO_4^-	Permanganate ion
	CN^-	Cyanide ion		
	OH^-	**Hydroxide ion**		
2–	**O^{2-}**	**Oxide ion**	**CO_3^{2-}**	**Carbonate ion**
	O_2^{2-}	Peroxide ion	CrO_4^{2-}	Chromate ion
	S^{2-}	**Sulfide ion**	$Cr_2O_7^{2-}$	Dichromate ion
			SO_4^{2-}	**Sulfate ion**
3–	N^{3-}	Nitride ion	**PO_4^{3-}**	**Phosphate ion**

* The most common ions are in boldface.

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Formulas and Names of Some Common Polyatomic Ions

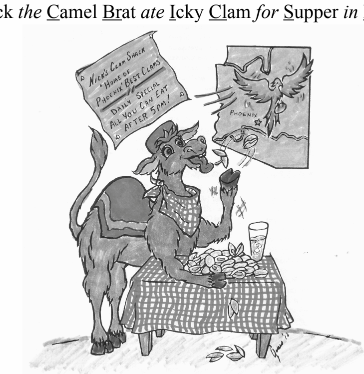
Formula	Name	Formula	Name
CATION: Positive Ion			
NH_4^+	ammonium ion		
ANIONS: Negative Ions			
Based on a Group 4A element		***Based on a Group 7A element***	
CN^-	cyanide ion	ClO^-	hypochlorite ion
$CH_3CO_2^-$	acetate ion	ClO_2^-	chlorite ion
CO_3^{2-}	carbonate ion	ClO_3^-	chlorate ion
HCO_3^-	hydrogen carbonate ion (or bicarbonate ion)	ClO_4^-	perchlorate ion
Based on a Group 5A element		***Based on a transition metal***	
NO_2^-	nitrite ion	CrO_4^{2-}	chromate ion
NO_3^-	nitrate ion	$Cr_2O_7^{2-}$	dichromate ion
PO_4^{3-}	phosphate ion	MnO_4^-	permanganate ion
HPO_4^{2-}	hydrogen phosphate ion		
$H_2PO_4^-$	dihydrogen phosphate ion		
Based on a Group 6A element			
OH^-	hydroxide ion		
SO_3^{2-}	sulfite ion		
SO_4^{2-}	sulfate ion		
HSO_4^-	hydrogen sulfate ion (or bisulfate ion)		

Note: many O containing anions have names ending in –ate (or -ite).

MAR

Introducing: Nick the Camel!

Nick *the* Camel Brat *ate* Icky Clam *for* Supper *in* Phoenix



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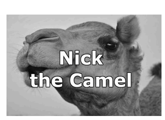
Nick the Camel



Nick *the* Camel Brat *ate* Icky Clam *for* Supper *in* Phoenix

	Consonants = Oxygen	Vowels = Charge	Polyatomic Ion
Nick = Nitrate	3	-1	NO_3^-
Camel = Carbonate	3	-2	CO_3^{2-}
Brat = Bromate	3	-1	BrO_3^-
Icky = Iodate	3	-1	IO_3^-
Clam = Chlorate	3	-1	ClO_3^-
Supper = Sulfate	4	-2	SO_4^{2-}
Phoenix = Phosphate	4	-3	PO_4^{3-}

Did Nick have Crepes for dessert too? :)

Crepes = chromate	4	-2	CrO_4^{2-}

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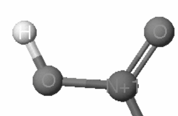
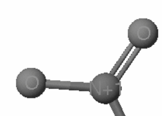
Some Common Polyatomic Ions



HNO_3
nitric acid



NO_3^-
nitrate ion

Many polyatomic ions related by a hydrogen ion (H^+) to an acid

Potassium nitrate somewhat common!

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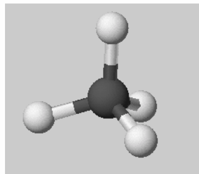
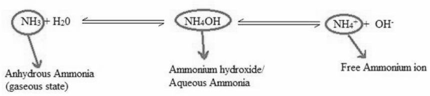
Some Common Polyatomic Ions

NH_4^+ **is the** ammonium **ion**

One of the few common polyatomic *cations*

ammonia (NH_3) *plus* acid (H^+) *gives the* ammonium cation:





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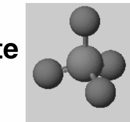
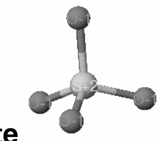
Some Common Polyatomic Ions

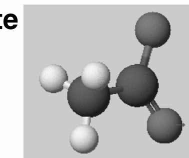
PO_4^{3-} phosphate



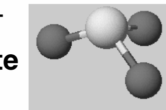
SO_4^{2-} sulfate



$CH_3CO_2^-$ acetate



SO_3^{2-} sulfite



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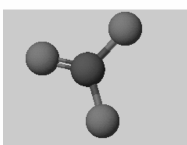
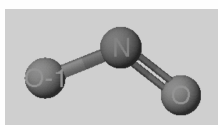
Some Common Polyatomic Ions

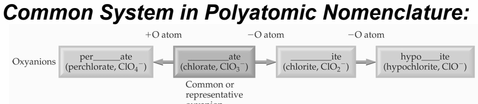
NO_3^- nitrate





NO_2^- nitrite

Common System in Polyatomic Nomenclature:



MAR

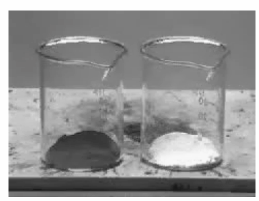
See: "Common Polyatomic Ions and the Corresponding Acids" handout

COMPOUNDS FORMED FROM IONS
(ionic bonding)

CATION + ANION → COMPOUND

Zn^{+2} + S^{-2} → ZnS

A neutral compound requires equal number of positive and negative charges.



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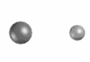
COMPOUNDS FORMED FROM IONS
(ionic bonding)

CATION + ANION → COMPOUND

Na^+ + Cl^- → NaCl

A neutral compound requires equal number of positive and negative charges.



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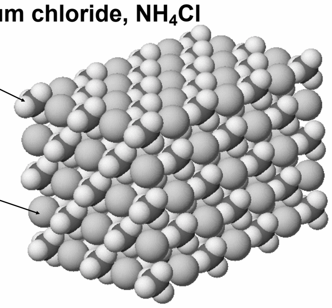
IONIC COMPOUNDS

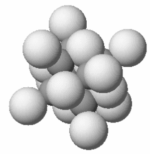
ammonium chloride, NH_4Cl



MAR

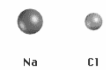
Some Ionic Compounds

Ca^{2+} + F^- --->
CaF_2
Name = **calcium fluoride**
Mg^{2+} + NO_3^- ---->
$Mg(NO_3)_2$
Name = **magnesium nitrate**
Fe^{2+} + PO_4^{3-} ---->
$Fe_3(PO_4)_2$
Name = **iron(II) phosphate**
also: **$FePO_4$** *Name* = **iron(III) phosphate,**
etc.


calcium fluoride

MAR

Properties of Ionic Compounds
Forming NaCl from Na and Cl_2

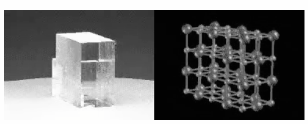


A metal atom can transfer an electron to a nonmetal.

The resulting cation and anion are attracted to each other by electrostatic forces.

MAR

Electrostatic Forces



The oppositely charged ions in ionic compounds are attracted to one another by ELECTROSTATIC FORCES.

These forces are governed by COULOMB'S LAW.

MAR

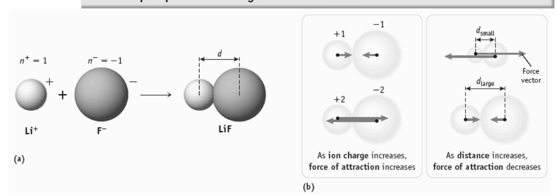
Electrostatic Forces

COULOMB'S LAW

charge on + and − ions; charge on electron

$$\text{Force} = -k\,\frac{(n^+e)(n^-e)}{d^2}$$

proportionality constant; distance between ions



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Electrostatic Forces

COULOMB'S LAW

charge on + and − ions; charge on electron

$$\text{Force} = -k\,\frac{(n^+e)(n^-e)}{d^2}$$

proportionality constant; distance between ions

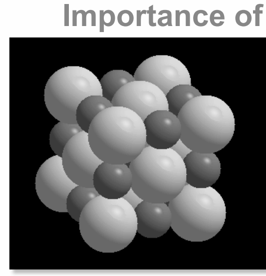
As ion charge increases, the attractive force _______________.

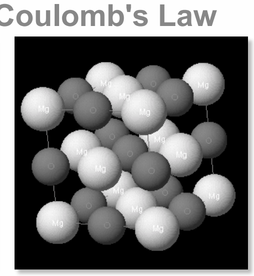
As the distance between ions increases, the attractive force _______________.

This idea is important and will come up many times in future discussions - see *handout*

MAR

Importance of Coulomb's Law


NaCl, Na^+ and Cl^-, m.p. 804 °C

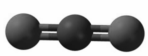
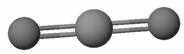


MgO, Mg^{2+} and O^{2-} m.p. 2800 °C

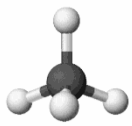
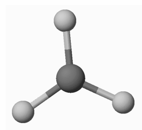
MAR

Molecular (Covalent) Compounds

Compounds without Ions

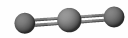
NO_2 nitrogen dioxide **CO_2 carbon dioxide**

CH_4 methane **BCl_3 boron trichloride**

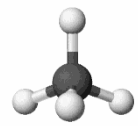
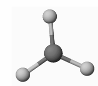
MAR

Naming Molecular (Covalent) Compounds

NO_2 nitrogen dioxide **CO_2 Carbon dioxide**

Covalent compounds formed from **two or more** nonmetals**; use** Greek prefixes

CH_4 methane **BCl_3 boron trichloride**

Ionic compounds generally involve a metal and nonmetal (NaCl) - ***do not use Greek prefixes with metals!***

MAR

Greek Prefixes

1	***mono***	**6**	***hexa***
2	***di***	**7**	***hepta***
3	***tri***	**8**	***octa***
4	***tetra***	**9**	***nona***
5	***penta***	**10**	***deca***

MAR

Three Types of Compound Naming

Fixed charge metal + nonmetal (ionic) ***No Greek prefixes or Roman numbers***

Al_2O_3 - aluminum oxide

Variable charge metal + nonmetal (ionic) ***Use Roman numbers***

Fe_2O_3 - iron(III) oxide

Watch variable charge: FeO = iron(II) oxide, etc.

Nonmetal + nonmetal (covalent) ***Use Greek prefixes***

P_2O_3 - diphosphorus trioxide

Also P_2O_5, = diphosphorus pentoxide, etc.

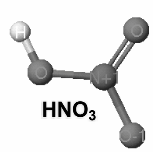
MAR

ACIDS

Acids create *hydrogen ions* in water, acidic H^+ listed *first* in compound

Some acids include:

HCl	**hydrochloric**
HNO_3	**nitric**
$HClO_4$	**perchloric**
H_2SO_4	**sulfuric**
HBrO	**hypobromous**



MAR

more in Chapter Four and Nomenclature Lab

BASES

Bases are metal hydroxides, creating OH^- in water

Some bases include:

NaOH	**sodium hydroxide**
KOH	**potassium hydroxide**
LiOH	**lithium hydroxide**
$Ca(OH)_2$	**calcium hydroxide**
$Fe(OH)_3$	**iron(III) hydroxide**

MAR

more in Chapter Four and Nomenclature Lab

Hydrated Compounds

When prepared in water and isolated as solids, many ionic compounds have water molecules trapped in the lattice.

"Waters of hydration" **result in** *beautiful* **colors**

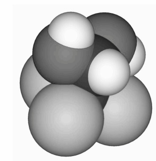
$CuSO_4 \cdot 5\ H_2O_{(s)}$ + heat
→ $CuSO_{4(s)}$ + $5\ H_2O_{(g)}$

MAR

Hydrated Compounds

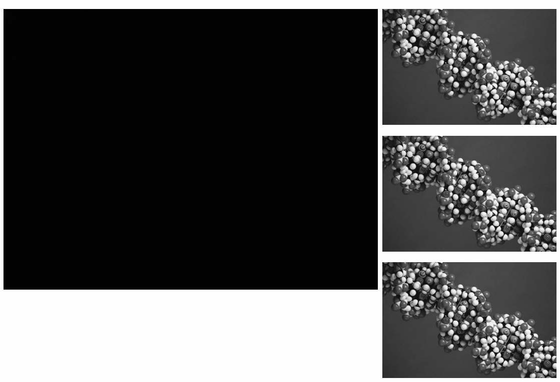
Nomenclature: **use Greek prefix + "hydrate" after regular name**

$CuSO_4 \cdot 5\ H_2O$ = copper(II) sulfate pentahydrate
$MgSO_4 \cdot 7\ H_2O$ = magnesium sulfate heptahydrate
$NiCl_2 \cdot 6\ H_2O$ = nickel(II) chloride hexahydrate
$CuSO_4$ ***without water*** **called** "anhydrous" **copper(II) sulfate**



MAR

ATOMS OVERWHELMING YOU?

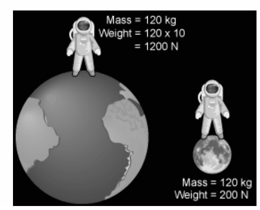


MAR

MOLECULAR WEIGHT AND MOLAR MASS

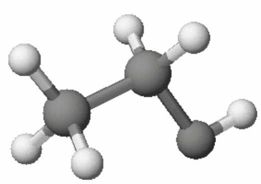
Molecular weight **is the sum of the atomic weights of all atoms in the molecule.**

Molar mass **= molecular weight in grams**



MAR

What is the molar mass of ethanol, C_2H_6O?



1 mol contains

2 mol C (12.01 g C/1 mol) = 24.02 g C

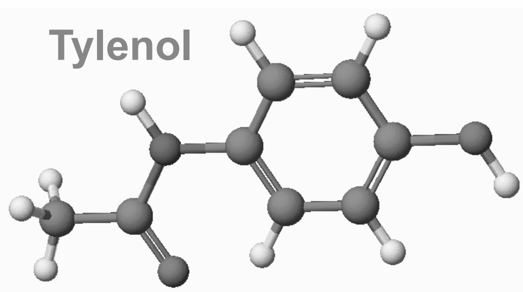
6 mol H (1.01 g H/1 mol) = 6.06 g H

1 mol O (16.00 g O/1 mol) = 16.00 g O

TOTAL = molar mass = 46.08 g/mol

MAR

Tylenol

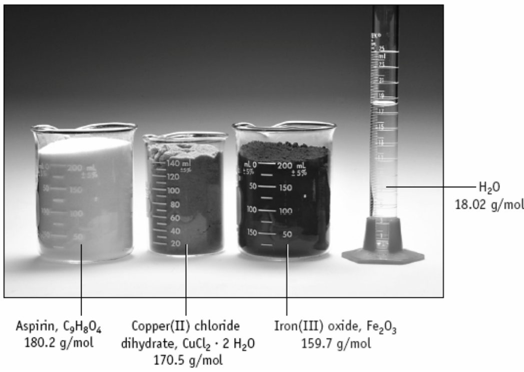


Formula = $C_8H_9NO_2$
Molar mass = 151.16 g/mol

MAR

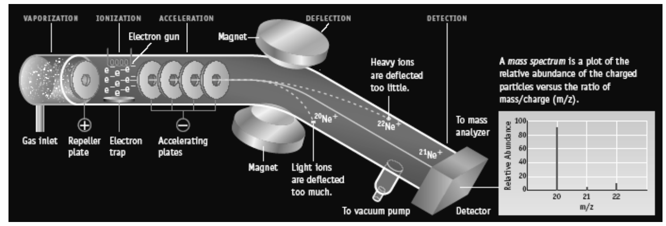
Try to use at least four sig figs for molar mass

Molar Mass



MAR

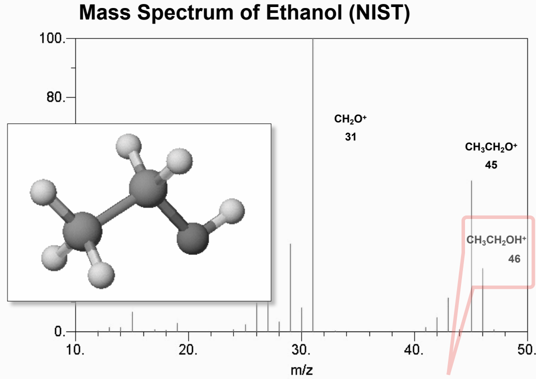
How to Determine a Formula?



Mass spectrometer

MAR

Mass Spectrum of Ethanol (NIST)



Molar mass = 46.08 g mol-1

MAR

How many moles of alcohol (ethanol) are present in a "standard" can of beer if there are 21.3 g of C_2H_6O?

(a) Molar mass of C_2H_6O = 46.08 g/mol

(b) Calc. moles of alcohol

$$21.3\text{ g} \bullet \frac{1\text{ mol}}{46.08\text{ g}} = 0.462\text{ mol}$$

0.46223958...

MAR

How many molecules of alcohol (ethanol) are present in a "standard" can of beer if there are 21.3 g of C_2H_6O?

We know there are 0.462 mol of C_2H_6O.

$$0.462\text{ mol} \bullet \frac{6.022 \times 10^{23}\text{ molecules}}{1\text{ mol}}$$

= 2.78 x 10^{23} molecules

2.78216...E23

MAR

How many atoms of C are present in a "standard" can of beer if there are 21.3 g of C_2H_6O?

We know there are 2.78 x 10^{23} molecules.

Each molecule contains 2 C atoms.

Therefore, the number of C atoms is

$$2.78 \times 10^{23}\text{ molecules} \bullet \frac{2\text{ C atoms}}{1\text{ molecule}}$$

= 5.56 x 10^{23} C atoms

5.56E23

MAR

*2.78 *10^{23} * 9 = **2.50 x 10^{24} atoms** in the 21.3 g of ethanol!*

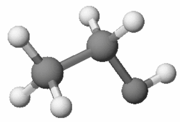
Empirical and Molecular Formulas

A pure compound always consists of the same elements combined in the same proportions by weight.

Therefore, we can express molecular composition as PERCENT BY WEIGHT

Ethanol, C_2H_6O

52.13% C, 13.15% H, 34.72% O

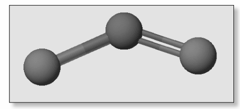


MAR

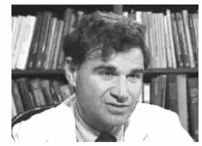
Percent Composition

Consider the nitrogen-oxygen family of compounds:

NO_2, nitrogen dioxide, and NO, nitrogen monoxide (or *nitric oxide*)



Structure of NO_2



Chemistry of NO, nitrogen monoxide

MAR

Percent Composition

Consider NO_2, Molar mass = ?

What is the weight percent of N and of O?

*To find the **weight percent** of an element in a compound:*

$$\text{Wt. \% X} = \frac{\text{g of X in compound}}{\text{molar mass of compound}} \bullet 100\%$$

In water (H_2O):

$$\text{Wt. \% O} = \frac{16.00 \text{ g O}}{18.02 \text{ g } H_2O} \bullet 100\% = 88.79\ \%$$

%H = 100 - 88.79 = 11.21%

MAR

Percent Composition

Consider NO_2, Molar mass = ?

What is the weight percent of N and of O?

$$\text{Wt. \% N} = \frac{14.01 \text{ g N}}{46.01 \text{ g } NO_2} \bullet 100\% = 30.45\ \%$$

$$\text{Wt. \% O} = \frac{2(16.00 \text{ g O})}{46.01 \text{ g } NO_2} \bullet 100\% = 69.55\ \%$$

***Test yourself:* What are the weight percentages of N and O in N_2O_4?**

MAR

Determining Formulas

In chemical analysis we first determine the % by weight of each element in a given amount of pure compound and derive the **EMPIRICAL** or **SIMPLEST** formula**.**

Weight percentages lead to empirical formulas *(but **not** molecular formulas!)*

PROBLEM: A compound of B and H is 81.10% B. What is its empirical formula?

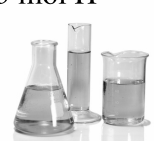
MAR

*A compound of B and H is 81.10% B. What is its **empirical** formula?*

Calculate the number of moles of each element in 100.0 g of sample.

$$81.10 \text{ g B} \bullet \frac{1 \text{ mol}}{10.81 \text{ g}} = 7.502 \text{ mol B}$$

$$18.90 \text{ g H} \bullet \frac{1 \text{ mol}}{1.008 \text{ g}} = 18.75 \text{ mol H}$$



MAR

A compound of B and H is 81.10% B. What is its ***empirical*** *formula?*

Take the ratio of moles of B and H. Always divide by the smaller number.

$$\frac{18.75 \text{ mol H}}{7.502 \text{ mol B}} = \frac{2.499 \text{ mol H}}{1.000 \text{ mol B}} = \frac{2.5 \text{ mol H}}{1.0 \text{ mol B}}$$

But we need a whole number ratio.

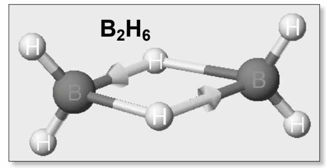
2.5 mol H/1.0 mol B = 5 mol H to 2 mol B

EMPIRICAL FORMULA = B_2H_5

MAR

The compound has an *empirical* formula of B_2H_5. What is its molecular formula?

Is the molecular formula B_2H_5, B_4H_{10}, B_6H_{15}, B_8H_{20}, etc.?



B_2H_6 is one example of this class of compounds.

MAR

The compound has an empirical formula (EF) of B_2H_5. What is its molecular formula?

To solve, **need the molar mass of the compound using a** ***mass spectrometer (a separate experiment)***

Next, **determine molar mass of the empirical formula**

Compare **molar mass of the compound to the molar mass of the empirical formula to get a whole number ratio of empirical formula units in the molecular formula**

MAR

The compound has an empirical formula (EF) of B_2H_5. What is its molecular formula?

Example:

A compound has an empirical formula of **CH_2** and a molar mass of **28.1 g mol^{-1}**. Find the **molecular formula**.

Molar mass compound (28.1 g mol^{-1}) given via outside experiment.

Molar mass empirical formula (CH_2) =

12.01 + 2*1.01 = 14.03 g mol^{-1}

Now compare molar mass compound to molar mass of empirical formula:

$$\frac{28.1 \text{ g/mol}}{14.03 \text{ g/mol of } CH_2} = \frac{2 \text{ units of } CH_2}{1 \text{ mol}}$$

Molecular formula = $(CH_2)_2$ = C_2H_4

MAR

The compound has an empirical formula (EF) of B_2H_5. What is its molecular formula?

In the boron problem,

Molar mass of ***compound (from mass spectrometer, a separate experiment)*** **= 53.3 g/mol**

Molar mass of ***empirical formula*** **(B_2H_5) = 26.67 g/mol**

*(2*10.81 + 5*1.01 = 26.67 g/mol of EF)*

Now find ratio of these masses.

$$\frac{53.3 \text{ g/mol}}{26.67 \text{ g/mol of } B_2H_5} = \frac{2 \text{ units of } B_2H_5}{1 \text{ mol}}$$

Molecular formula = $(B_2H_5)_2$ = B_4H_{10}

MAR

Determining a Molecular Formula: *Overview*

First, **convert percent by mass element values into moles** ***(assume 100 g)*****, then compare the moles to get the empirical formula (EF)**

$$\frac{18.75 \text{ mol H}}{7.502 \text{ mol B}} = \frac{2.499 \text{ mol H}}{1.000 \text{ mol B}} = \frac{2.5 \text{ mol H}}{1.0 \text{ mol B}}$$

2.5 mol H/1.0 mol B = 5 H to 2 B = B_2H_5

Next, **find the molar mass (MM) of the compound, then compare MM of compound to MM of EF**

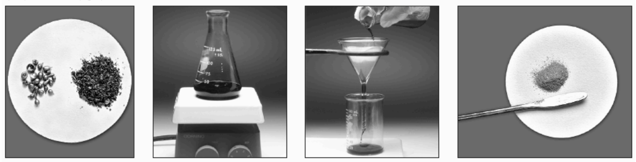
$$\frac{53.3 \text{ g/mol}}{26.67 \text{ g/mol of } B_2H_5} = \frac{2 \text{ units of } B_2H_5}{1 \text{ mol}}$$

Molecular formula = $(B_2H_5)_2$ = B_4H_{10}

MAR

Determine the formula of a compound of Sn and I using the following data.

(a) Weighed samples of tin (left) and iodine (right). (b) The tin and iodine are heated in a solvent. (c) The hot reaction mixture is filtered to recover unreacted tin. (d) When the solvent cools, solid, orange tin iodide forms and is isolated.



Mass of Sn in the beginning = 1.056 g
Mass of iodine (I_2) used = 1.947 g
Mass of Sn remaining = 0.601 g

MAR

Tin and Iodine Compound

Find the mass of Sn that combined with 1.947 g I_2.

Mass of Sn initially = 1.056 g
Mass of Sn recovered = 0.601 g
Mass of Sn used = 0.455 g
Find moles of Sn used:

$$0.455 \text{ g Sn} \bullet \frac{1 \text{ mol}}{118.7 \text{ g}} = 3.83 \times 10^{-3} \text{ mol Sn}$$

MAR

Tin and Iodine Compound

Now find the number of moles of I_2 that combined with 3.83×10^{-3} mol Sn

Mass of I_2 used = **1.947 g**

$$1.947 \text{ g } I_2 \bullet \frac{1 \text{ mol}}{253.81 \text{ g}} = 7.671 \times 10^{-3} \text{ mol } I_2$$

But we need **mol of I** for formula, *not I_2*, so convert:

$$7.671 \times 10^{-3} \text{ mol } I_2 \bullet \frac{2 \text{ mol I}}{1 \text{ mol } I_2} = 1.534 \times 10^{-2} \text{ mol I}$$

***So* 1.534×10^{-2} mol of iodine atoms were used in this reaction**

MAR

Tin and Iodine Compound

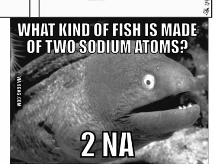
Now find the ratio of number of moles of moles of I and Sn that combined.

$$\frac{1.534 \times 10^{-2} \text{ mol I}}{3.83 \times 10^{-3} \text{ mol Sn}} = \frac{4.01 \text{ mol I}}{1.00 \text{ mol Sn}}$$

Empirical formula is SnI_4
tin(IV) iodide

MAR

End of Chapter 2 Part 2



See also:
- *Chapter Two Part 2 Study Guide*
- *Chapter Two Part 2 Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*

MAR

Important Equations, Constants, and Handouts from this Chapter:

A mole = 6.022×10^{23}

- **be able to find the molar mass of any compound using the periodic table**
- **be able to convert grams of a compound into moles and/or molecules**
- **understand how to calculate empirical formula (EF) and molecular formula (MF) using the molar mass and mass percentages**

Nomenclature: Greek prefixes, Roman numbers, nonmetal + nonmetal, fixed charge metal + nonmetal, variable charge metal + nonmetal, polyatomic ions, acids, bases, hydrated compounds, the 7 diatomics, cations, anions, covalent, ionic, the "stairs", Coulomb's Law

MAR

End of Chapter Problems: ***Test Yourself***

See practice problem set #3 and self quizzes for ***nomenclature*** *examples and practice*

1. Determine the molar mass for aluminum chloride, iron(III) oxide and phosphorus tribromide.
2. How many grams in 0.0255 mol of propanol (C_3H_7OH)? How many molecules? How many atoms of C?
3. Calculate the weight percent of lead in PbS, lead(II) sulfide. What mass of lead (in grams) is present in 10.0 g of PbS?
4. Succinic acid has an empirical formula is $C_2H_3O_2$ and a molar mass is 118.1 g/mol. What is its molecular formula?
5. A new compound containing xenon and fluorine was isolated by shining sunlight on a mixture of Xe (0.526 g) and F_2 gas. If you isolate 0.678 g of the new compound, what is its empirical formula?
6. Direct reaction of iodine (I_2) and chlorine (Cl_2) produces an iodine chloride, I_xCl_y, a bright yellow solid. If you completely used up 0.678 g of iodine and produced 1.246 g of I_xCl_y, what is the **empirical formula** of the compound? A later experiment showed that the molar mass of I_xCl_y was 467 g/mol. What is the **molecular formula** of the compound?

MAR

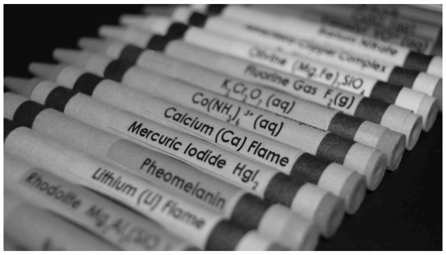
End of Chapter Problems: ***Answers***

1. 133 g/mol, 160. g/mol, 271 g/mol
2. 1.53 g C_3H_7OH, 1.54 x 10^{22} molecules, 4.62 x 10^{22} atoms C
3. 86.59%, 8.66 g Pb
4. $C_4H_6O_4$
5. XeF_2
6. ICl_3 , I_2Cl_6

Be sure to view practice problem set #3 and self quizzes for ***nomenclature*** *examples and practice*

MAR

Chemistry 221 Exam I Review
Chapters 1, 2 and 3.1 - 3.2



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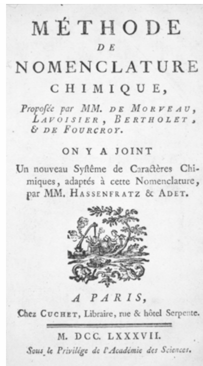
Chemistry 221
Professor Michael Russell

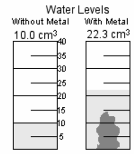


1787 Nomenclature book by Lavoisier

MAR

A piece of metal with a mass of 33.2 g is immersed in 10.0 mL of water in a graduated cylinder. Determine the identity of the metal.

	Density
Cu	8.96 g/cm^3
Mg	1.74 g/cm^3
Fe	7.87 g/cm^3
Ag	10.5 g/cm^3
Al	2.70 g/cm^3



A. Cu, 8.96 g/cm^3
B. Mg, 1.74 g/cm^3
C. Fe, 7.87 g/cm^3
D. Ag, 10.5 g/cm^3
E. Al, 2.70 g/cm^3

MAR

You are given temperature readings at three locations on Earth: 29 ˚C, 45 ˚F, and 256 K. What is the order of increasing temperature?

A. 29 ˚C < 45 ˚F < 256 K
B. 45 ˚F < 29 ˚C < 256 K
C. 256 K < 29 ˚C < 45 ˚F
D. 256 K < 45 ˚F < 29 ˚C
E. 45 ˚F < 256 K < 29 ˚C

MAR

Place the following in order of increasing size:
215 mm, 9 cm, 2.3 m, and 0.125 m

A. 215 mm < 9 cm < 2.3 m < 0.125 m
B. 215 mm < 9 cm < 0.125 m < 2.3 m
C. 9 cm < 215 mm < 0.125 m < 2.3 m
D. 9 cm < 0.125 m < 215 mm < 2.3 m
E. 0.125 m < 9 cm < 215 mm < 2.3 m

MAR

Which of the following is NOT an isotope of element X (Z = 9)?

A. $^{19}_{9}X$
B. $^{20}_{10}X$
C. $^{18}_{9}X$
D. $^{21}_{9}X$
E. $^{22}_{9}X$

MAR

Which statement describes the composition of a neutral atom of iron-58?

A. 26 neutrons, 32 protons, and 26 electrons
B. 32 neutrons, 26 protons, and 26 electrons
C. 26 neutrons, 26 protons, and 32 electrons
D. 26 neutrons, 26 protons, and 26 electrons
E. Not enough information

MAR

An element (E) has several naturally occurring isotopes, with the following abundances:
^{72}E, 54.5% ^{73}E, 15.6% ^{74}E, 29.9%

The most reasonable atomic weight for this element would be

A. 72.1
B. 72.8
C. 73.4
D. 73.8
E. 74.0

MAR

Which ion in the following list is NOT likely to form?

A. Na^{+}
B. Mg^{3+}
C. Al^{3+}
D. Fe^{2+}
E. Zn^{2+}

MAR

When the ion Sr^{2+} forms,

A. the Sr atom loses 1 electron and now has the same number of electrons as Kr
B. the Sr atom loses 1 electron and now has the same number of electrons as Xe
C. the Sr atom loses 2 electrons and now has the same number of electrons as Kr
D. the Sr atom gains 2 electrons and now has the same number of electrons as Kr
E. the Sr atom loses 3 electrons and now has the same number of electrons as Kr

MAR

Which compound formula and name in the list is NOT correct?

A. $CaSO_4$, calcium sulfate
B. $NaNO_3$, sodium nitrate
C. MgI_2, magnesium iodide
D. NH_4PO_4, ammonium phosphate
E. $Ca(ClO)_2$, calcium hypochlorite

MAR

Which compound in the list is NOT ionic?

A. LiCl, lithium chloride
B. SO_2, sulfur dioxide
C. AlF_3, aluminum fluoride
D. $Ba(NO_3)_2$, barium nitrate
E. $NaHCO_3$, sodium hydrogen carbonate

MAR

Sodium oxalate has the formula $Na_2C_2O_4$. Based on this information, the formula for iron(III) oxalate is

A. FeC_2O_4
B. $Fe(C_2O_4)_2$
C. $Fe(C_2O_4)_3$
D. $Fe_2(C_2O_4)_3$
E. $Fe_3(C_2O_4)_2$

MAR

All of the following statements concerning ionic compounds are correct EXCEPT

A. as the ion charges increase, the attraction between the ions increases.
B. ionic compounds form extended 3-dimensional networks called crystal lattices.
C. ionic crystals tend to be rigid, and they cleave along planes.
D. positive and negative ions are attracted to each other by electrostatic forces.
E. the electrostatic forces are weaker in CaO than in NaCl.

MAR

You have 0.25 mol of each of the following elements. Which one has the largest mass?

A. Fe
B. Al
C. Zn
D. Ca
E. C

MAR

Calculate the average mass of one chromium atom.

A. 8.634×10^{-23} g
B. 6.626×10^{-34} g
C. 6.022×10^{-23} g
D. 51.996 g
E. too small to calculate accurately

MAR

Which answer best represents the percent composition of the compound CO_2?

A. 50.0% C & 50.0% O
B. 12.0% C & 88.0% O
C. 27.3% C & 72.7% O
D. 12.0% C & 32.0% O
E. 64.0% C & 36.0% O

MAR

Nitrogen and oxygen form a series of oxides with the general formula N_xO_y. One of them has 46.67% N. The empirical formula for this oxide is

A. N_2O
B. NO
C. NO_2
D. N_2O_3
E. N_2O_5

MAR

Combining 6.54 g of Zn with oxygen gives a white powder, Zn_xO_y, with a mass of 8.14 g (all of the Zn reacts.) The empirical formula is:

A. ZnO
B. Zn_2O
C. ZnO_2
D. Zn_2O_3
E. Zn_3O_4

MAR

Glyceraldehyde has an empirical formula of CH_2O and a molar mass is 90.08 g/mol. The ***molecular formula*** is:

A. CH_2O
B. $C_2H_4O_2$
C. $C_3H_6O_3$
D. $C_4H_8O_4$
E. $C_4H_4O_4$

MAR

A compound with P and F is 75.41% F with a molar mass of 251.94 g/mol. What is the molecular formula?

A. P_2F_2
B. $P_{10}F_2$
C. PF_5
D. P_2F_5
E. P_2F_{10}

MAR

Which compound below represents iodous acid?

A. HI
B. HIO
C. HIO_2
D. HIO_3
E. HIO_4

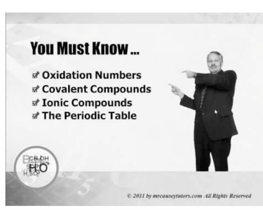
MAR

What is the molar mass of nickel(II) nitrate hexahydrate?

A. 139 g/mol
B. 201 g/mol
C. 228 g/mol
D. 291 g/mol
E. 539 g/mol

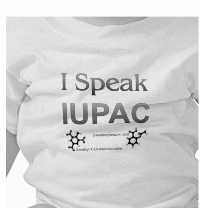
MAR

End of Review - *good luck with your studying!*



Need more practice?
- *Practice Problem Sets (online)*
- *Concept Guides (Companion and online)*
- *Chapter Guides (online)*
- *End of Chapter Problems in Textbook (every other question has answer at end)*

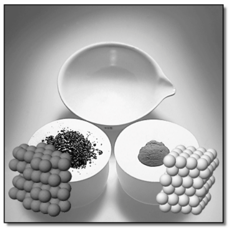
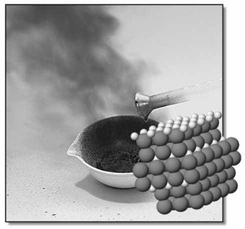
Good luck with your studying!



MAR

Chemical Reactions

Chapter 4 Part 1

Reactants: Zn + I_2 ⟶ Product: ZnI_2

Chemistry 221

Professor Michael Russell

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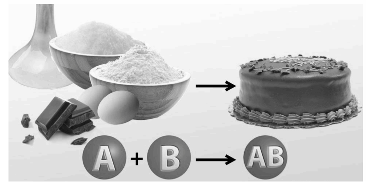
MAR

Chemistry as Cooking! - the Chemical Reaction

"Recipe" and technique leads to successful creations

Must know amounts to add, how much will be produced

Haphazard additions can be disastrous!

Chemical Equations

Depict the kind of reactants and products and their relative amounts in a reaction.

$4\ Al(s) + 3\ O_2(g)$ ---> $2\ Al_2O_3(s)$

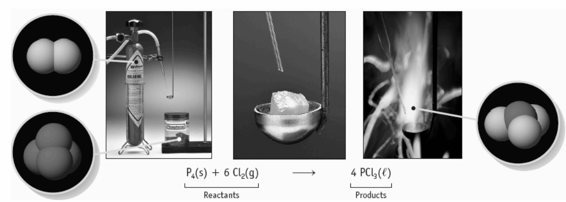
The numbers in the front are called

stoichiometric coefficients

The letters (s), (g), (aq) and (l) are the physical states of compounds.

MAR

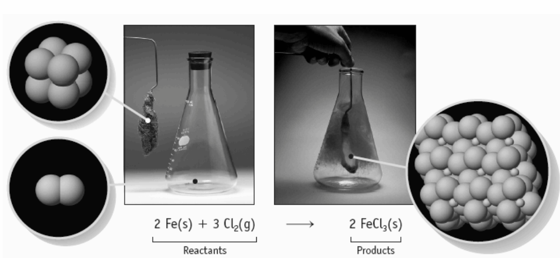
Reaction of Phosphorus with Cl_2



Notice the stoichiometric coefficients and the physical states of the reactants and products.

MAR

Reaction of Iron with Cl_2



Evidence of a chemical reaction:

heat change, precipitate formation, gas evolution, color change

MAR

Chemical Equations

$4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$

This equation means:

4 Al atoms + 3 O_2 molecules
---give--->
2 molecules of Al_2O_3

or

4 moles of Al + 3 moles of O_2
---give--->
2 moles of Al_2O_3

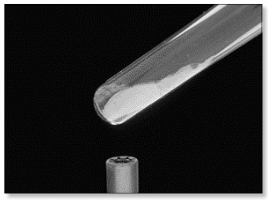
MAR

Chemical Equations

Because the same atoms are present in a reaction at the beginning and at the end, the amount of matter in a system does not change.

The Law of the Conservation of Matter

Also known as the Law of Mass Action



MAR

Chemical Equations / Lavoisier

Because of the principle of the conservation of matter,

an equation must be balanced.

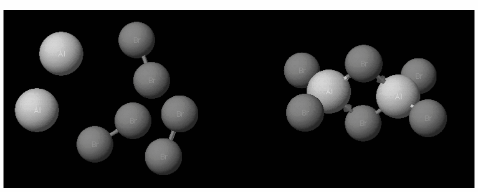
It must have the same number of atoms of the same kind on both sides.

Lavoisier, 1788

MAR

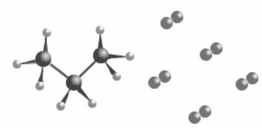
Balancing Equations

2 Al(s) + 3 Br_2(liq) ---> Al_2Br_6(s)



MAR

Balancing Equations



C_3H_8(g) + 5 O_2(g) ---->
3 CO_2(g) + 4 H_2O(g)

2 B_4H_{10}(g) + 11 O_2(g) ---->
4 B_2O_3(g) + 10 H_2O(g)

MAR

Balancing Equations - *Hints*

Balance those atoms which occur in only one compound on each side last *(i.e. O_2 in previous examples)*

Balance the remaining atoms first

Reduce coefficients to smallest whole integers

Check your answer *if uncertain*

Helpful but optional: Check that charges are balanced

MAR

STOICHIOMETRY

Stoichiometry is the study of the quantitative aspects of chemical reactions.

Stoichiometry rests on the principle of the conservation of matter.

MAR

Stoichiometry

The balanced chemical equation

$4\,Al(s) + 3\,O_2(g) \longrightarrow 2\,Al_2O_3(s)$

implies *all* of the following ratios:

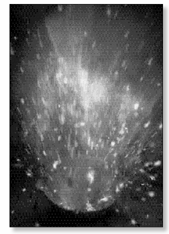
$$\frac{4 \text{ mol Al}}{3 \text{ mol } O_2} \quad \frac{4 \text{ mol Al}}{2 \text{ mol } Al_2O_3} \quad \frac{3 \text{ mol } O_2}{2 \text{ mol } Al_2O_3}$$

$$\frac{3 \text{ mol } O_2}{4 \text{ mol Al}} \quad \frac{2 \text{ mol } Al_2O_3}{4 \text{ mol Al}} \quad \frac{2 \text{ mol } Al_2O_3}{3 \text{ mol } O_2}$$

These are nothing more than "conversion units" in dimensional analysis!

MAR

PROBLEM: If 454 g of NH_4NO_3 decomposes, how much N_2O and H_2O are formed? What is the theoretical yield of products?



STEP 1

Write the balanced chemical equation

$NH_4NO_3 \longrightarrow N_2O + 2\,H_2O$

MAR

454 g *of* $NH_4NO_3 \longrightarrow N_2O + 2\,H_2O$

STEP 2 Convert mass reactant (454 g) --> moles

$$454 \text{ g} \bullet \frac{1 \text{ mol}}{80.04 \text{ g}} = 5.68 \text{ mol } NH_4NO_3$$

80.04 g/mol = molar mass of NH_4NO_3

MAR

454 g *of* $NH_4NO_3 \longrightarrow N_2O + 2\,H_2O$

STEP 3 Convert moles reactant --> moles product

Relate moles NH_4NO_3 to moles product expected.

1 mol NH_4NO_3 --> 2 mol H_2O

Express as a STOICHIOMETRIC FACTOR:

$$\frac{2 \text{ mol } H_2O \text{ produced}}{1 \text{ mol } NH_4NO_3 \text{ used}}$$

MAR

454 g *of* $NH_4NO_3 \longrightarrow N_2O + 2\,H_2O$

STEP 3 Convert moles reactant (5.68 mol) --> moles product

$$5.68 \text{ mol } NH_4NO_3 \bullet \frac{2 \text{ mol } H_2O \text{ produced}}{1 \text{ mol } NH_4NO_3 \text{ used}}$$

= 11.4 mol H_2O produced

How many moles of N_2O produced?

Answer = **5.68 mol N_2O**

MAR

454 g *of* $NH_4NO_3 \longrightarrow N_2O + 2\,H_2O$

STEP 4 Convert moles product (11.4 mol) --> mass product

This is called the THEORETICAL YIELD

$$11.4 \text{ mol } H_2O \bullet \frac{18.02 \text{ g}}{1 \text{ mol}} = 204 \text{ g } H_2O$$

ALWAYS FOLLOW THESE STEPS IN SOLVING STOICHIOMETRY PROBLEMS!

MAR

454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

STEP 5 How much N_2O is formed?

Total mass of reactants =
total mass of products

454 g NH_4NO_3 = ___ g N_2O + 204 g H_2O

mass of N_2O = 250. g ***law of mass action!***

could also turn mol NH_4NO_3 into mol N_2O, then grams of N_2O:

5.68 mol N_2O * 44.01 g/mol = 250. g

MAR

454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

Compound	NH_4NO_3	N_2O	H_2O
Initial (g)	**454 g**	**0**	**0**
Initial (mol)	**5.68mol**	**0**	**0**
Change (mol)	**-5.68**	**+5.68**	**+2(5.68)**
Final (mol)	**0**	**5.68**	**11.4**
Final (g)	**0**	**250.**	**204**

Mass is conserved!

MAR

454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

STEP 6 Calculate the percent yield

We predicted a yield of 250. g of N_2O. If you isolated only 131 g of N_2O, what is the percent yield of N_2O?

This compares the **theoretical yield** *(250. g) and* **actual yield** *(131 g) of N_2O.*

MAR

454 g *of* NH_4NO_3 --> N_2O + 2 H_2O

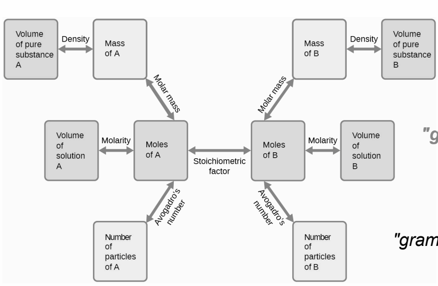
STEP 6 Calculate the percent yield

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \bullet 100\%$$

$$\%\ \text{yield} = \frac{131\ \text{g}}{250.\ \text{g}} \bullet 100\% = 52.4\%$$

MAR

GENERAL PLAN FOR STOICHIOMETRY CALCULATIONS



We use the "grams - moles - moles - grams" conversion often in chemistry (mass is grams)

Sometimes I call this the "grams - moles - moles - grams" dance lol

Molarity in next chapter - See <u>Stoichiometry Guide</u>

MAR

PROBLEM: Using 5.00 g of H_2O_2, what mass of O_2 and of H_2O can be obtained?

2 H_2O_2(liq) ---> 2 H_2O(g) + O_2(g)

Reaction is catalyzed by MnO_2

MAR

PROBLEM: Using 5.00 g of H_2O_2, what mass of O_2 and of H_2O can be obtained?

$2\ H_2O_2(liq) \longrightarrow 2\ H_2O(g) + O_2(g)$

Reaction is catalyzed by MnO_2

Step 1: moles of H_2O_2

Step 2: use STOICHIOMETRIC FACTOR to calculate moles of O_2

Step 3: mass of O_2 (2.35 g)

Step 4: mass of H_2O (2.65 g)

Try this problem yourself!

MAR

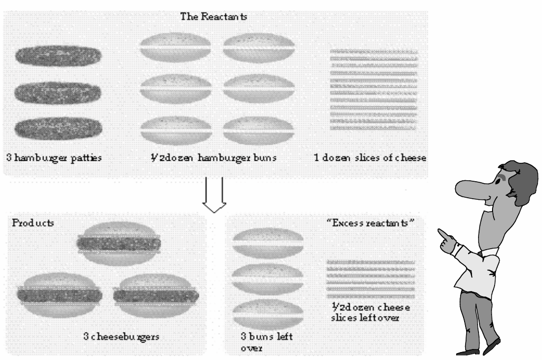
Reactions Involving a LIMITING REACTANT

In a given reaction, there is not enough of one reagent to use up the other reagent completely.

The reagent in short supply LIMITS the quantity of product that can be formed.

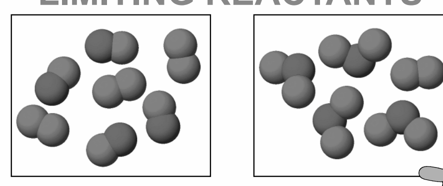
MAR

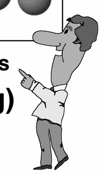
LIMITING REACTANTS



MAR

LIMITING REACTANTS





Reactants | Products

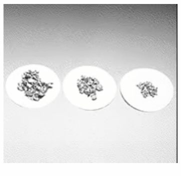
$2\ NO(g) + O_2\ (g) \longrightarrow 2\ NO_2(g)$

Limiting reactant = ___________

Excess reactant = ____________

MAR

LIMITING REACTANTS



React solid Zn with 0.100 mol HCl (aq)

$Zn_{(s)} + 2\ HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2(g)}$

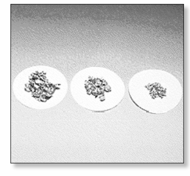
Left: Balloon inflates fully, some Zn left
* More than enough Zn to use up the 0.100 mol HCl

Center: Balloon inflates fully, no Zn left
* Right amount of each (HCl and Zn)

Right: Balloon does not inflate fully, no Zn left.
* Not enough Zn to use up 0.100 mol HCl

MAR

LIMITING REACTANTS



React solid Zn with 0.100 mol HCl (aq)

$Zn_{(s)} + 2\ HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2(g)}$

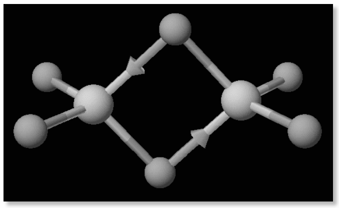
0.100 mol HCl [1 mol Zn/2 mol HCl] = 0.0500 mol Zn

	Left	Center	Right
mass Zn (g)	7.00	3.27	1.31
mol Zn	0.107	0.050	0.020
mol HCl	0.100	0.100	0.100
mol HCl/mol Zn	0.93	2.00	5.00
Lim Reactant	LR = HCl	no LR	LR = Zn

MAR

Reaction to be Studied:

2 Al + 3 Cl_2 ---> Al_2Cl_6



PROBLEM: Mix 5.40 g of Al with 8.10 g of Cl_2. How many grams of Al_2Cl_6 can form?

MAR

Step 1 of the Limiting Reactant problem: **Compare actual mole ratio of reactants to theoretical mole ratio.**

2 Al + 3 Cl_2 ---> Al_2Cl_6

Reactants must be in the mole ratio

$$\frac{\text{mol Cl}_2}{\text{mol Al}} = \frac{3}{2}$$

MAR

Deciding on the Limiting Reactant

2 Al + 3 Cl_2 ---> Al_2Cl_6

If $\boxed{\dfrac{\text{mol Cl}_2}{\text{mol Al}} > \dfrac{3}{2}}$

then there is not enough Al to use up all the Cl_2, and the limiting reagent is **Al**

MAR

Deciding on the Limiting Reactant

2 Al + 3 Cl_2 ---> Al_2Cl_6

If $\boxed{\dfrac{\text{mol Cl}_2}{\text{mol Al}} < \dfrac{3}{2}}$

then there is not enough Cl_2 to use up all the Al, and the limiting reagent is **Cl_2**

MAR

Step 2 of the Limiting Reactant problem:
Calculate moles of each reactant

We have 5.40 g of Al and 8.10 g of Cl_2. How much Al_2Cl_6 can form?

$$5.40 \text{ g Al} \bullet \frac{1 \text{ mol}}{27.0 \text{ g}} = 0.200 \text{ mol Al}$$

$$8.10 \text{ g Cl}_2 \bullet \frac{1 \text{ mol}}{70.9 \text{ g}} = 0.114 \text{ mol Cl}_2$$

MAR ***2 Al + 3 Cl_2 ---> Al_2Cl_6***

Step 3 of the Limiting Reactant problem:
Compare moles to find limiting reactant

$$\frac{\text{mol Cl}_2}{\text{mol Al}} = \frac{0.114 \text{ mol}}{0.200 \text{ mol}} = 0.570$$

This should be 3/2 or 1.5/1 if reactants are present in the exact stoichiometric ratio.

$\boxed{\text{Limiting reagent is } \underline{\text{Cl}_2}}$

MAR ***2 Al + 3 Cl_2 ---> Al_2Cl_6***

Mix 5.40 g of Al with 8.10 g of Cl_2. What mass of Al_2Cl_6 can form?

Limiting reactant = Cl_2
Base all calculations on Cl_2

grams Cl_2 → moles Cl_2 → $\frac{1 \text{ mol } Al_2Cl_6}{3 \text{ mol } Cl_2}$ → moles Al_2Cl_6 → grams Al_2Cl_6

MAR

2 Al + 3 Cl_2 ---> Al_2Cl_6

CALCULATIONS: calculate mass of Al_2Cl_6 expected using limiting reactant.

Step 1: Calculate moles of Al_2Cl_6 expected using chlorine:

$$0.114 \text{ mol } Cl_2 \bullet \frac{1 \text{ mol } Al_2Cl_6}{3 \text{ mol } Cl_2} = 0.0380 \text{ mol } Al_2Cl_6$$

Step 2: Calculate mass of Al_2Cl_6 expected based on chlorine:

$$0.0380 \text{ mol } Al_2Cl_6 \bullet \frac{266.4 \text{ g } Al_2Cl_6}{\text{mol}} = \boxed{10.1 \text{ g } Al_2Cl_6}$$

MAR

2 Al + 3 Cl_2 ---> Al_2Cl_6

Alternate Limiting Reactant Method

Calculate theoretical yield of product based on *both* reactants.

Smaller theoretical yield comes from limiting reactant, greater yield from excess reactant.

$$8.10 \text{ g } Cl_2 \bullet \frac{1 \text{ mol}}{70.9 \text{ g}} \bullet \frac{1 \text{ mol } Al_2Cl_6}{3 \text{ mol } Cl_2} \bullet \frac{266.4 \text{ g}}{1 \text{ mol}} = 10.1 \text{ g } Al_2Cl_6$$

$$5.40 \text{ g Al} \bullet \frac{1 \text{ mol}}{27.0 \text{ g}} \bullet \frac{1 \text{ mol } Al_2Cl_6}{2 \text{ mol Al}} \bullet \frac{266.4 \text{ g}}{1 \text{ mol}} = 26.6 \text{ g } Al_2Cl_6$$

10.1 g < 26.6 g, *so:* limiting reactant = Cl_2, theoretical yield = 10.1 g, excess reactant = Al

MAR

2 Al + 3 Cl_2 ---> Al_2Cl_6

How much of which reactant will remain when reaction is complete?

Cl_2 was the limiting reactant. Therefore, Al was present in excess. But by how much?

***First* find how much Al was required based on limiting reactant (Cl_2).**

***Then* find how much Al is in excess.**

MAR

2 Al + 3 Cl_2 ---> Al_2Cl_6

Calculating Excess Al

2 Al + 3 Cl_2 ⟶ products

0.200 mol   **0.114 mol = LR**

$$8.10 \text{ g } Cl_2 \bullet \frac{1 \text{ mol}}{70.9 \text{ g}} \bullet \frac{2 \text{ mol Al}}{3 \text{ mol } Cl_2} \bullet \frac{26.98 \text{ g}}{1 \text{ mol}} = 2.05 \text{ g Al}$$

Excess Al = Al available - Al required
= 5.40 g - 2.05 g
= 3.35 g Al unused in reaction

MAR

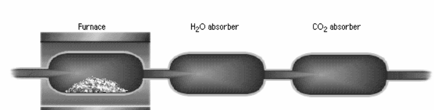
2 Al + 3 Cl_2 ---> Al_2Cl_6

Using Stoichiometry to Determine a Formula

Hydrocarbons, C_xH_y, can be burned in oxygen to give CO_2 and H_2O (combustion reaction).

The CO_2 and H_2O can be collected to determine the empirical formula of the hydrocarbon.

C_xH_y + O_2 ---> CO_2 + H_2O

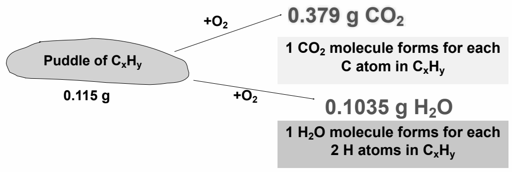


MAR

Using Stoichiometry to Determine a Formula

What is the empirical formula of a hydrocarbon, C_xH_y, if burning 0.115 g produces 0.379 g CO_2 and 0.1035 g H_2O?

C_xH_y + some O_2 ---> 0.379 g CO_2 + 0.1035 g H_2O



MAR

Using Stoichiometry to Determine a Formula

C_xH_y + some oxygen --->
0.379 g CO_2 + 0.1035 g H_2O

***First*, recognize that all C in CO_2 and all H in H_2O comes from C_xH_y.**

1. Calculate amount of C in CO_2

8.61×10^{-3} mol CO_2 --> 8.61×10^{-3} mol C

1 mol C per 1 mol CO_2

2. Calculate amount of H in H_2O

5.744×10^{-3} mol H_2O -- >1.149×10^{-2} mol H

2 mol H per 1 mol water!

MAR

Using Stoichiometry to Determine a Formula

C_xH_y + some oxygen --->
0.379 g CO_2 + 0.1035 g H_2O

Now find <u>ratio</u> of mol H/mol C to find values of x and y in C_xH_y.

1.149×10^{-2} mol H/ 8.61×10^{-3} mol C

= 1.33 mol H / 1.00 mol C

= 4 mol H / 3 mol C

Empirical formula = C_3H_4

MAR

Formulas with C, H and O

Caproic acid, the substance responsible for "dirty gym socks" smell, contains C, H *and* O.

Combustion analysis of 0.450 g caproic acid gives 0.418 g H_2O and 1.023 g CO_2, and the molar mass was found to be 116.2 g mol^{-1}.

What is the molecular formula of caproic acid?

$C_xH_yO_z$ + some oxygen ---> 1.023 g CO_2 + 0.418 g H_2O

***Careful:* oxygen comes from caproic acid and O_2, need special technique**

MAR

Formulas with C, H and O

Combustion analysis of 0.450 g caproic acid gives 0.418 g H_2O and 1.023 g CO_2, and the molar mass is 116.2 g mol^{-1}. What is the molecular formula?

Start with "regular" approach for mol H & mol C:

0.418 g H_2O * (mol/18.02 g) * (2 mol H/mol H_2O) = 0.0464 mol H

0.0464 mol H * (1.01 g/mol H) = 0.0469 g H

1.023 g CO_2 * (mol/44.01 g) * (1 mol C/mol CO_2) = 0.02324 mol C

0.02324 mol C * (12.01 g/mol C) = 0.2791 g C

***Why did we convert to grams?* Law of Mass Action!**

MAR

Formulas with C, H and O

0.450 g caproic acid: 0.418 g H_2O (0.0464 mol H, 0.0469 g H) and 1.023 g CO_2 (0.02324 mol C, 0.2791 g C), molar mass = 116.2 g/mol. What is the molecular formula?

Realize that 0.450 g of caproic acid equals all the g C, g H *and g O* in the complex.

Converting mol H and mol C to grams, then subtracting from 0.450 g, gives g O in caproic acid:

$$\underbrace{0.450\text{ g}}_{\text{caproic acid}} - \underbrace{0.0469\text{ g}}_{\text{g of H in acid}} - \underbrace{0.2791\text{ g}}_{\text{g of C in acid}} = \underbrace{0.124\text{ g O}}_{\text{g of O in acid}}$$

0.124 g O * (mol O / 16.00 g) = 0.00775 mol O

MAR

Formulas with C, H and O

0.450 g caproic acid: 0.418 g H_2O (0.0464 mol H) and 1.023 g CO_2 (0.02324 mol C), molar mass = 116.2 g/mol, 0.00775 mol O. What is the molecular formula?

Now compare moles:

$C_{0.02324}H_{0.0464}O_{0.00775}$ *gives* C_3H_6O = *empirical formula*

C_3H_6O has a molar mass of 58.1 g/mol, which is half of the 116.2 g/mol value

Molecular Formula = $(C_3H_6O)_2$, or

$C_6H_{12}O_2$

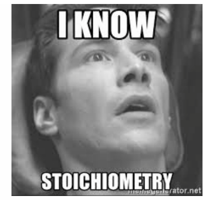
You can now find empirical formulas based on combustion analysis (this chapter) and elemental percentages (previous chapter)!

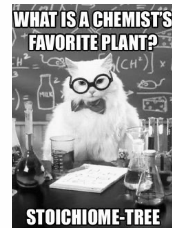
MAR

End of Chapter 4 Part 1

See also:

- *Chapter Four Part 1 Study Guide*
- *Chapter Four Part 1 Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*





MAR

Important Equations, Constants, and Handouts from this Chapter:

- **be able to find the theoretical yield, actual yield, percent yield**
- **be able to determine the limiting reactant, excess reactant, excess reactant remaining at end of reaction**
- **understand how to calculate empirical formula (EF) and molecular formula (MF) using organic compounds containing oxygen**

Balancing Equations:
Reactants, Products, states of matter (s, l, g, aq), stoichiometric coefficients, Law of Conservation of Matter ("mass action")

MAR

End of Chapter Problems: ***Test Yourself***

See practice problem set #4 and self quizzes for ***balancing chemical equations*** *examples and practice*

1. What mass of Br_2, in grams, is required for complete reaction with 2.56 g of Al? What mass of white, solid Al_2Br_6 is expected? The equation: 2 Al(s) + 3 Br_2(l) → Al_2Br_6(s)
2. Aluminum chloride is made by treating aluminum with chlorine: **2 Al(s) + 3 Cl_2(g) → 2 $AlCl_3$(s)** If you begin with 2.70 g of Al and 4.05 g of Cl_2, which reactant is limiting? What mass of $AlCl_3$ can be produced? What mass of the excess reactant remains when the reaction is completed?
3. $Cu(NH_3)_4SO_4$ is made via: **$CuSO_4$(aq) + 4 NH_3(aq) → $Cu(NH_3)_4SO_4$(aq)** If you use 10.0 g of $CuSO_4$ and excess NH_3, what is the theoretical yield of $Cu(NH_3)_4SO_4$? If you isolate 12.6 g of $Cu(NH_3)_4SO_4$, what is the percent yield of $Cu(NH_3)_4SO_4$?
4. An unknown compound has the formula $C_xH_yO_z$. You burn 0.0956 g of the compound and isolate 0.1356 g of CO_2 and 0.0833 g of H_2O. What is the empirical formula of the compound? If the molar mass is 62.1 g/mol, what is the molecular formula?

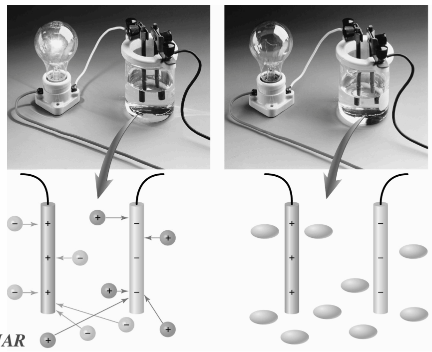
MAR

End of Chapter Problems: ***Answers***

1. 22.7 g Br_2 , 25.3 g Al_2Br_6
2. Chlorine is limiting; 5.09 g $AlCl_3$; 1.67 g Al remains
3. 14.3 g $Cu(NH_3)_4SO_4$, 88.3%
4. EF = CH_3O, MF = $C_2H_6O_2$

Be sure to view practice problem set #4 and self quizzes for ***balancing chemical equations*** *examples and practice*

MAR



Solution Stoichiometry and Chemical Reactions

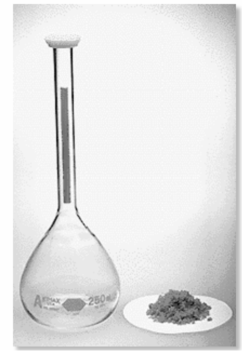
Chapter 3 & Chapter 4, or "Chapter 4 Part II"

Chemistry 221
Professor Michael Russell

Last update: 4/29/24

MAR

Terminology



In SOLUTION **we need to define the -**

- SOLVENT

the component whose physical state is preserved when solution forms

- SOLUTE

the other solution component

- **Compounds are *soluble* when they dissolve, *insoluble* when they stay as solids**

MAR

Water Solubility of Ionic Compounds

SOLUBLE COMPOUNDS	EXCEPTIONS
Almost all salts of Na^+, K^+, NH_4^+	
Salts of nitrate, NO_3^-; chlorate, ClO_3^-; perchlorate, ClO_4^-; acetate, $CH_3CO_2^-$	
Almost all salts of Cl^-, Br^-, I^-	Halides of Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds containing F^-	Fluorides of Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
Salts of sulfate, SO_4^{2-}	Sulfates of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

INSOLUBLE COMPOUNDS	EXCEPTIONS
Most salts of carbonate, CO_3^{2-}; phosphate, PO_4^{3-}; oxalate, $C_2O_4^{2-}$; chromate, CrO_4^{2-}	Salts of NH_4^+ and the alkali metal cations
Most metal sulfides, S^{2-}	
Most metal hydroxides and oxides	

If one ion from the "Soluble Compd." list is present in a compound, the compound is water soluble.

$Ba(NO_3)_2(aq)$? soluble

$BaCl_2(aq)$? soluble

$BaSO_4(aq)$? insoluble *should write $BaSO_4(s)$!*

Use this solubility guide in CH 221-223!

MAR

WATER SOLUBILITY OF IONIC COMPOUNDS

Not all ionic compounds dissolve in water. Some are INSOLUBLE.

Many ions, however, make compounds SOLUBLE ***all* of the time.**

Examples: Na^+, K^+, Li^+, NH_4^+, NO_3^-, ClO_3^-, ClO_4^-, $CH_3CO_2^-$, **and *most*** SO_4^{2-}, Cl^-, Br^- **and** I^- **compounds.**

SOLUBLE COMPOUNDS	EXCEPTIONS
Almost all salts of Na^+, K^+, NH_4^+	
Salts of nitrate, NO_3^-; chlorate, ClO_3^-; perchlorate, ClO_4^-; acetate, $CH_3CO_2^-$	
Almost all salts of Cl^-, Br^-, I^-	Halides of Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds containing F^-	Fluorides of Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
Salts of sulfate, SO_4^{2-}	Sulfates of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

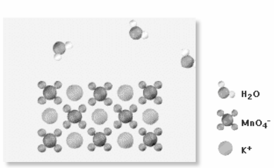
INSOLUBLE COMPOUNDS	EXCEPTIONS
Most salts of carbonate, CO_3^{2-}; phosphate, PO_4^{3-}; oxalate, $C_2O_4^{2-}$; chromate, CrO_4^{2-}	Salts of NH_4^+ and the alkali metal cations
Most metal sulfides, S^{2-}	
Most metal hydroxides and oxides	

MAR

Ionic Compounds in Aqueous Solution

Many reactions involve soluble ionic compounds, especially reactions in water - aqueous solutions.

$KMnO_4$ in water



$K^+(aq)$ + $MnO_4^-(aq)$

MAR

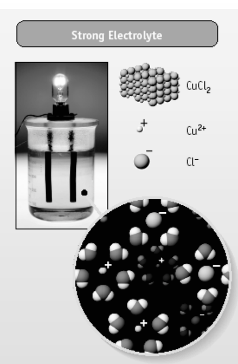
Aqueous Solutions

How do we know ions are present in aqueous solutions?

The solutions conduct electricity! The ions are called ELECTROLYTES

HCl, $KMnO_4$, $MgCl_2$, and NaCl are strong electrolytes. **They dissociate completely (*or nearly so*) into ions.**

$KMnO_4(aq) \longrightarrow K^+(aq) + MnO_4^-(aq)$

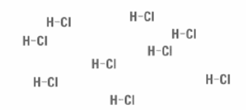


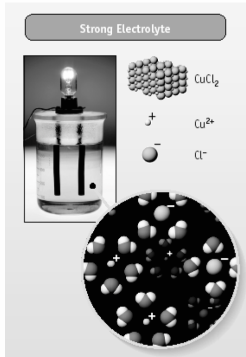
A strong electrolyte conducts electricity. $CuCl_2$ is completely dissociated into Cu^{2+} and Cl^- ions.

MAR

Aqueous Solutions

HCl, $MgCl_2$, and NaCl are strong electrolytes.
They dissociate completely (or nearly so) into ions.





A strong electrolyte conducts electricity. $CuCl_2$ is completely dissociated into Cu^{2+} and Cl^- ions.

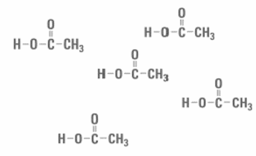
MAR

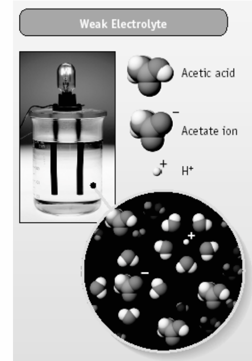
Aqueous Solutions

Acetic acid ionizes only to a small extent, so it is a weak electrolyte

$CH_3CO_2H(aq)$ --->

$CH_3CO_2^-(aq)$ + $H^+(aq)$



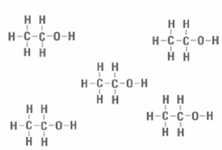


A weak electrolyte conducts electricity poorly because so few ions are present in solution.

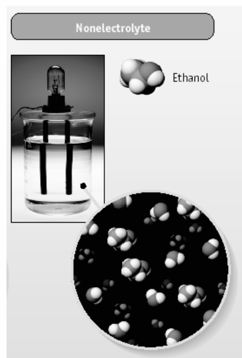
MAR

Aqueous Solutions

Some compounds (sugar, ethanol, acetone, etc.) dissolve in water but do not conduct electricity. They are called nonelectrolytes.



See "Dissolve, Dissociate and Electrolyte" Guide



A nonelectrolyte does not conduct electricity because no ions are present in solution.

MAR

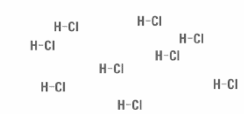
Acids

An acid -------> H^+ in water

Some strong acids include:

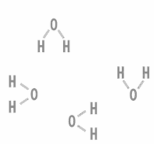
HCl	**hydrochloric**
HNO_3	**nitric**
$HClO_4$	**perchloric**
H_2SO_4	**sulfuric**

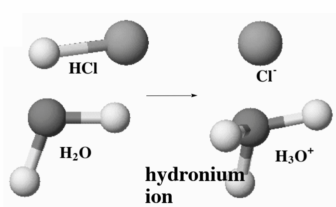
All strong acids are strong electrolytes



MAR

The Nature of Acids





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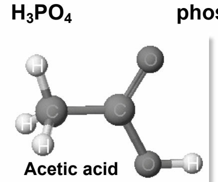
Weak Acids

All weak acids are weak electrolytes

CH_3CO_2H	**acetic acid**
H_2CO_3	**carbonic acid**
H_3PO_4	**phosphoric acid**



MAR

BASES

Base ---> OH^- in water

Bases are often metal hydroxides
$NaOH(aq)$ ---> $Na^+(aq)$ + $OH^-(aq)$

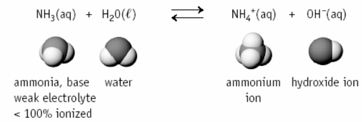
NaOH is a strong base

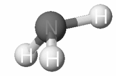
NaOH NaOH NaOH NaOH NaOH NaOH

All strong bases are strong electrolytes

MAR

Ammonia, NH_3
An Important Weak Base





All weak bases are weak electrolytes

MAR

Common Acids and Bases

Strong Acids (Strong Electrolytes)		Strong Bases (Strong Electrolytes)	
HCl	Hydrochloric acid	LiOH	Lithium hydroxide
HBr	Hydrobromic acid	NaOH	Sodium hydroxide
HI	Hydroiodic acid	KOH	Potassium hydroxide
HNO_3	Nitric acid		
$HClO_4$	Perchloric acid		
H_2SO_4	Sulfuric acid		
Weak Acids (Weak Electrolytes)*		**Weak Base (Weak Electrolyte)**	
H_3PO_4	Phosphoric acid	NH_3	Ammonia
H_2CO_3	Carbonic acid		
CH_3CO_2H	Acetic acid		
$H_2C_2O_4$	Oxalic acid		
$C_4H_6O_6$	Tartaric acid		
$C_6H_8O_7$	Citric acid		
$C_9H_8O_4$	Aspirin		

*These are representative of hundreds of weak acids.

Know the strong acids & bases!

MAR

Net Ionic Equations

$Mg(s)$ + 2 $HCl(aq)$ --> $H_2(g)$ + $MgCl_2(aq)$

We really should write:

$Mg(s)$ + 2 $H^+(aq)$ + 2 $Cl^-(aq)$ --->
$H_2(g)$ + $Mg^{2+}(aq)$ + 2 $Cl^-(aq)$

MAR

Net Ionic Equations

$Mg(s)$ + 2 $HCl(aq)$ → $H_2(g)$ + $MgCl_2(aq)$

Aqueous solutes (HCl, $MgCl_2$) dissociate; we *really* should write:

$Mg(s)$ + 2 $H^+(aq)$ + 2 $Cl^-(aq)$ →
$H_2(g)$ + $Mg^{2+}(aq)$ + 2 $Cl^-(aq)$

We leave the spectator ions (Cl^-) out in writing the NET IONIC EQUATION:
$Mg(s)$ + 2 $H^+(aq)$ → $H_2(g)$ + $Mg^{2+}(aq)$

See Net Ionic Reactions Handout

MAR

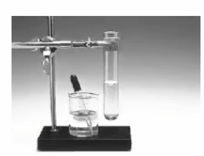
Net Ionic Equations

$K_2CrO_4(aq)$ + $Pb(NO_3)_2(aq)$ -->
$PbCrO_4(s)$ + 2 $KNO_3(aq)$

NET IONIC EQUATION
$Pb^{2+}(aq)$ + $CrO_4^{2-}(aq)$ ---> $PbCrO_4(s)$

K^+ and NO_3^- are spectators

See Net Ionic Reactions Handout



MAR

CHEMICAL REACTIONS IN WATER

We will look at
EXCHANGE REACTIONS

AX + BY ⟶ AY + BX

The anions exchange places between cations

$Pb(NO_3)_2(aq) + 2\ KI(aq)$
$\text{---->}\ PbI_2(s) + 2\ KNO_3(aq)$

Exchange reactions often called Double Displacement Reactions

MAR

Precipitation Reactions

The "driving force" is the formation of an insoluble compound - a precipitate.

$Fe(NO_3)_3(aq) + 3\ NaOH(aq)$
----->
$3\ NaNO_3(aq) + Fe(OH)_3(s)$

Net ionic equation

$Fe^{3+}(aq) + 3\ OH^-(aq)\ \text{--->}\ Fe(OH)_3(s)$

MAR

See "Five Types of Reactions" Handout

Acid-Base Reactions

Acids react readily with bases. The "driving force" is the formation of water.

$NaOH(aq) + HCl(aq) \rightarrow$
$NaCl(aq) + H_2O(liq)$

Net ionic equation:

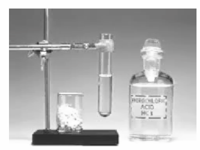
$OH^-(aq) + H^+(aq) \rightarrow$
$H_2O(liq)$

This applies to ALL reactions of STRONG acids and bases.

Acid-base reactions often called "neutralizations", water and "salt" created

MAR

See "Five Types of Reactions" Handout



Gas-Forming Reactions

$CaCO_3(s) + 2\ HCl(aq)\ \text{--->}$
$CaCl_2(aq) + H_2CO_3(aq)$

Carbonic acid is unstable and forms CO_2 & H_2O

$H_2CO_3(aq)\ \text{--->}\ CO_2(g) + \text{water}$

Another gas forming species:

$NH_4OH(aq)\ \text{--->}\ NH_3(g) + \text{water}$

MAR

See "Five Types of Reactions" Handout

Combustion Reactions

A special example of a gas-forming reaction

Used in quantitative chemistry; high temperatures

Reactants: *oxygen (O_2) and "something organic" (C, H, sometimes O or N)*

Products: *water and carbon dioxide (also NO_2 if N present)*

Examples:

$C_2H_{4(g)} + 3\ O_{2(g)} \rightarrow 2\ H_2O_{(g)} + 2\ CO_{2(g)}$

$4\ C_6H_5NO_2 + 29\ O_{2(g)} \rightarrow 10\ H_2O_{(g)} + 24\ CO_{2(g)} + 4\ NO_{2(g)}$

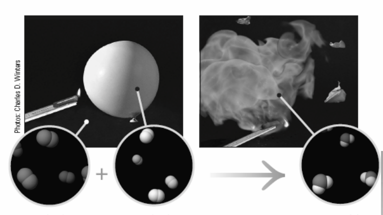
MAR

See "Five Types of Reactions" Handout

Oxidation-Reduction Reactions

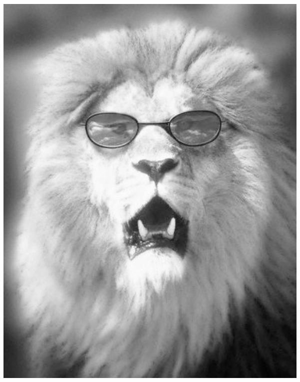
REDOX = reduction & oxidation

$2\ H_2(g) + O_2(g)\ \text{--->}\ 2\ H_2O(liq)$



MAR

See "Five Types of Reactions" Handout



LEO
says
GER

MAR

LEO **says** GER

Lose	Gain
Electrons	Electrons
Oxidized	Reduced

$Zn(s) \rightarrow Zn^{2+} + 2e^-$ ***Oxidized***

$Cu^{2+} + 2e^- \rightarrow Cu(s)$ ***Reduced***

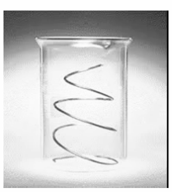
Can also use ***"OIL RIG":***

OIL = "Oxidation is Losing" *(electrons)*

RIG = "Reduction is Gaining" *(electrons)*

MAR

REDOX REACTIONS



In all reactions: if something has been oxidized then something has also been reduced:

$Cu(s) + 2\ Ag^+(aq) \rightarrow$
$Cu^{2+}(aq) + 2\ Ag(s)$

Redox reactions incredibly useful (fuels, batteries, much more)

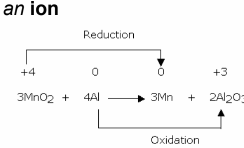
Oxidation numbers help visualize electron transfer pathways

MAR

Use **oxidation number rules** *to determine redox activity:*

- **Atoms in free element have ox. no. = 0**
  $Zn(s)$, $O_2(g)$, $Br_2(liq)$
- *In simple ions,* **ox. no. = charge on ion**
  -1 for Cl^-, +2 for Mg^{2+}
- **In compounds, F is *always* -1, O is -2** *(except peroxides (O = -1) and with F) and* **H is +1** *(except hydrides (H = -1))*
- **Sum of oxidation numbers = 0** *for a* **compound or equals the overall charge** *for an* **ion**

OXIDATION NUMBERS

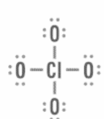


MAR

OXIDATION NUMBERS

Determining oxidation numbers takes practice

HF
H: +1
F: -1



ClO_4^-
Cl: +7
O: -2

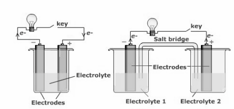
MAR

Recognizing a Redox Reaction

$2\ Al(s) + 3\ Cu^{2+}(aq) \text{ ---> } 2\ Al^{3+}(aq) + 3\ Cu(s)$

$Al(s) \text{ --> } Al^{3+}(aq) + 3\ e^-$

- **Ox. no. of Al *increases* as e^- are donated by the metal; Al is OXIDIZED (or the REDUCING AGENT)**



$Cu^{2+}(aq) + 2\ e^- \text{ --> } Cu(s)$

- **Ox. no. of Cu *decreases* as e^- are accepted by the ion; Cu^{2+} is REDUCED (or the OXIDIZING AGENT)**

MAR

Examples of Redox Reactions

NO = reducing agent
O_2 = oxidizing agent
$2\,NO + O_2 \rightarrow 2\,NO_2$

Fe = reducing agent
Cl_2 = oxidizing agent
$2\,Fe + 3\,Cl_2 \rightarrow 2\,FeCl_3$

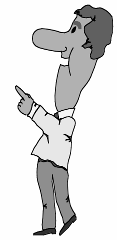
reducing agent = oxidized
oxidizing agent = reduced

MAR

Concentration (Molarity) of Solute

The amount of solute in a solution is given by its concentration

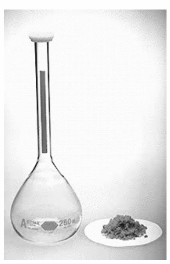
$$\text{Molarity (M)} = \frac{\text{moles solute}}{\text{liters of solution}}$$



Concentration (M) = […]
"3.6 M" means a concentration of 3.6 molarity
"concentration" and molarity often the same

MAR

PROBLEM: Dissolve 5.00 g of $NiCl_2 \cdot 6\,H_2O$ in enough water to make 250. mL of solution. Calculate molarity.



MAR

PROBLEM: Dissolve 5.00 g of $NiCl_2 \cdot 6\,H_2O$ in enough water to make 250. mL of solution. Calculate molarity.

Step 1: **Calculate moles of $NiCl_2 \cdot 6H_2O$**

$$5.00\text{ g} \cdot \frac{1\text{ mol}}{237.7\text{ g}} = 0.0210\text{ mol}$$

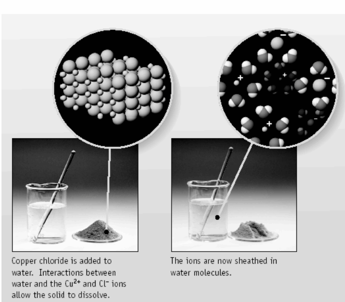
Step 2: **Calculate molarity**

$$\frac{0.0210\text{ mol}}{0.250\text{ L}} = 0.0841\text{ M}$$

$[NiCl_2 \cdot 6\,H_2O] = 0.0841$ M

MAR

The Nature of a $CuCl_2$ Solution Ion Concentrations



$CuCl_2$(aq) -->

Cu^{2+}(aq) + 2 Cl^-(aq)

If $[CuCl_2]$ = 0.30 M, then

$[Cu^{2+}]$ = 0.30 M

$[Cl^-]$ = 2 x 0.30 M = 0.60 M

MAR

USING MOLARITY

What mass of oxalic acid, $H_2C_2O_4$, is required to make 250. mL of a 0.0500 M solution?

$$\text{moles} = M \cdot V$$

Step 1: **Calculate moles of acid required.**

(0.0500 mol/L)(0.250 L) = 0.0125 mol

Step 2: **Calculate mass of acid required.**

(0.0125 mol)(90.00 g/mol) = 1.13 g

MAR

Preparing Solutions

Weigh out a solid solute and dissolve in a given quantity of solvent

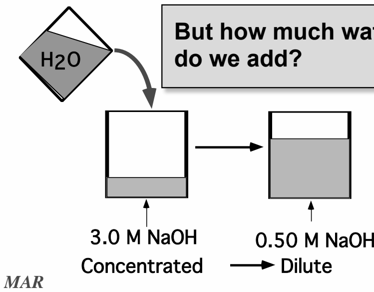
or

Dilute a concentrated solution to give one that is less concentrated.

MAR

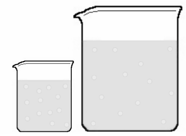
You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?

But how much water do we add?



The important point:

moles of NaOH in ORIGINAL solution = moles of NaOH in FINAL solution



MAR

You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?

Moles of NaOH in original solution =

M * V =

(3.0 mol/L)(0.0500 L) = 0.15 mol NaOH

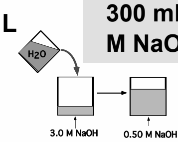
Therefore, moles of NaOH in final solution must also = 0.15 mol NaOH

(0.15 mol NaOH)(1 L/0.50 mol) = 0.30 L

or 300 mL = volume of final solution

Conclusion:

add 250 mL of water to 50.0 mL of 3.0 M NaOH to make 300 mL of 0.50 M NaOH.



MAR

Preparing Solutions by Dilution

A shortcut

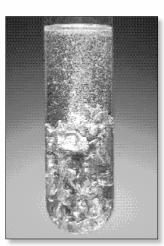
$$M_{initial} * V_{initial} = M_{final} * V_{final}$$

Often abbreviated: $M_iV_i = M_fV_f$ ***or*** $C_1V_1 = C_2V_2$

MAR

SOLUTION STOICHIOMETRY

Zinc reacts with acids to produce H_2 gas. What volume of 2.50 M HCl is needed to convert 10.0 g of Zn?



MAR

$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Zinc reacts with acids to produce H_2 gas. What volume of 2.50 M HCl is needed to convert 10.0 g of Zn?

Step 1: Calculate moles of Zn

$$10.0 \text{ g Zn} \bullet \frac{1.00 \text{ mol Zn}}{65.39 \text{ g Zn}} = 0.153 \text{ mol Zn}$$

Step 2: Use the stoichiometric factor

$$0.153 \text{ mol Zn} \bullet \frac{2 \text{ mol HCl}}{1 \text{ mol Zn}} = 0.306 \text{ mol HCl}$$

Step 3: Calculate volume of HCl required

$$0.306 \text{ mol HCl} \bullet \frac{1.00 \text{ L}}{2.50 \text{ mol}} = 0.122 \text{ L HCl}$$

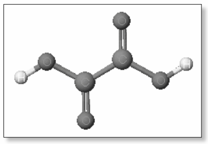
MAR

$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

ACID-BASE REACTIONS
Titrations

$H_2C_2O_4(aq)$ + 2 $NaOH(aq)$ --->
acid base
$Na_2C_2O_4(aq)$ + 2 $H_2O(liq)$

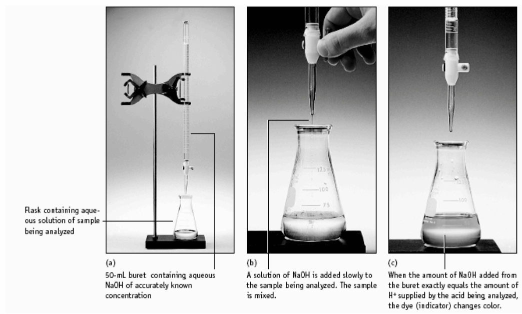
Carry out this reaction using a TITRATION.



Oxalic acid, $H_2C_2O_4$

MAR

Setup for titrating an acid with a base



MAR

LAB PROBLEM #1: Standardize a solution of NaOH - i.e., accurately determine its concentration.

1.065 g of $H_2C_2O_4$ (oxalic acid) requires 35.62 mL of NaOH for titration to an equivalence point. What is the concentration of the NaOH?

MAR *$H_2C_2O_4(aq)$ + 2 $NaOH(aq)$ ---> $Na_2C_2O_4(aq)$ + 2 $H_2O(liq)$*

1.065 g of $H_2C_2O_4$ (oxalic acid) requires 35.62 mL of NaOH for titration to an equivalence point. What is the concentration of the NaOH?

Step 1: **Calculate moles of $H_2C_2O_4$**

$$1.065 \text{ g} * \frac{1 \text{ mol}}{90.04 \text{ g}} = 0.01183 \text{ mol}$$

Step 2: **Calculate moles of NaOH req'd**

$$0.01183 \text{ mol acid} \bullet \frac{2 \text{ mol NaOH}}{1 \text{ mol acid}} = 0.02366 \text{ mol NaOH}$$

MAR *$H_2C_2O_4(aq)$ + 2 $NaOH(aq)$ ---> $Na_2C_2O_4(aq)$ + 2 $H_2O(liq)$*

1.065 g of $H_2C_2O_4$ (oxalic acid) requires 35.62 mL of NaOH for titration to an equivalence point. What is the concentration of the NaOH?

Step 1: **Calculate moles of $H_2C_2O_4$**
= 0.01183 mol acid

Step 2: **Calculate moles of NaOH req'd**
= 0.02366 mol NaOH

Step 3: **Calculate concentration of NaOH**

$$\frac{0.02366 \text{ mol NaOH}}{0.03562 \text{ L}} = 0.6642 \text{ M}$$

[NaOH] = 0.6642 M

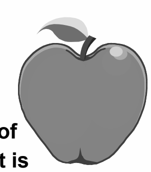
MAR *$H_2C_2O_4(aq)$ + 2 $NaOH(aq)$ ---> $Na_2C_2O_4(aq)$ + 2 $H_2O(liq)$*

LAB PROBLEM #2:
Use standardized NaOH to determine the amount of an acid in an unknown.

Apples contain malic acid, $C_4H_6O_5$.

$C_4H_6O_5(aq)$ + 2 $NaOH(aq)$ --->
$Na_2C_4H_4O_5(aq)$ + 2 $H_2O(liq)$

76.80 g of apple requires 34.56 mL of 0.6642 M NaOH for titration. What is weight % of malic acid?



MAR

76.80 g of apple requires 34.56 mL of 0.6642 M NaOH for titration. What is weight % of malic acid?

Step 1: **Calculate moles of NaOH used.**

M * V = (0.6642 M)(0.03456 L)

= 0.02295 mol NaOH

Step 2: **Calculate moles of acid titrated.**

$$0.02295 \text{ mol NaOH} \bullet \frac{1 \text{ mol acid}}{2 \text{ mol NaOH}}$$

= 0.01148 mol acid

MAR

$C_4H_6O_5(aq) + 2\ NaOH(aq) \longrightarrow Na_2C_4H_4O_5(aq) + 2\ H_2O(liq)$

76.80 g of apple requires 34.56 mL of 0.6642 M NaOH for titration. What is weight % of malic acid?

Step 1: **moles of NaOH = 0.02295**

Step 2: **moles of acid titrated = 0.01148**

Step 3: **Calculate mass of acid titrated.**

$$0.01148 \text{ mol acid} \bullet \frac{134.1 \text{ g}}{\text{mol}} = 1.539 \text{ g}$$

Step 4: **Calculate % malic acid.**

(1.539 g acid / 76.80 g apple) *100 =

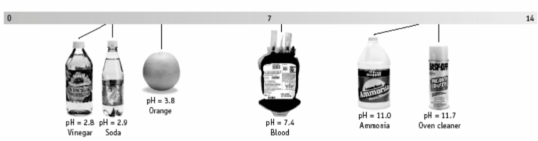
2.004 %

MAR

$C_4H_6O_5(aq) + 2\ NaOH(aq) \longrightarrow Na_2C_4H_4O_5(aq) + 2\ H_2O(liq)$

pH, a Concentration Scale

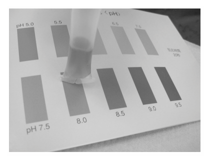
pH: a way to express acidity - the concentration (M) of H^+ in solution.



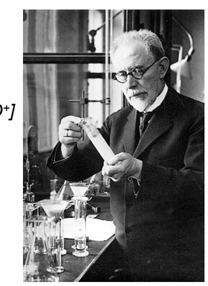
Low pH: high $[H^+]$ **High pH: low $[H^+]$**

Acidic solution	pH < 7
Neutral	pH = 7
Basic solution	pH > 7

MAR



The pH Scale



Søren Sørensen, *creator of the pH scale*

pH = - log $[H^+]$

$[H^+]$ often written as $[H_3O^+]$

In a neutral solution,

$[H^+]$ = $[OH^-]$ = 1.00 x 10^{-7} M at 25 °C

pH = - log $[H^+]$

= -log (1.00 x 10^{-7})

= - (-7) = 7

MAR

$[H^+]$ and pH

If the $[H^+]$ of soda is 1.6 x 10^{-3} M, the pH is ____?

Because pH = - log $[H^+]$

then

pH= - log (1.6 x 10^{-3})

pH = - (-2.80)

pH = 2.80

MAR

pH and $[H^+]$

If the pH of Coke is 3.12, it is ___________.

Because pH = - log $[H^+]$ then

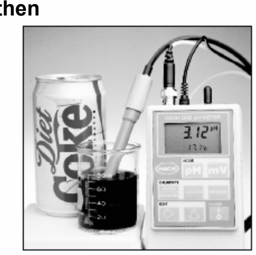
log $[H^+]$ = - pH

Take antilog and get

$[H^+]$ = 10^{-pH}

$[H^+]$ = $10^{-3.12}$

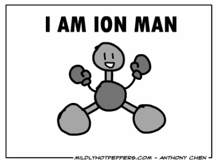
$[H^+]$ = 7.6 x 10^{-4} M



MAR

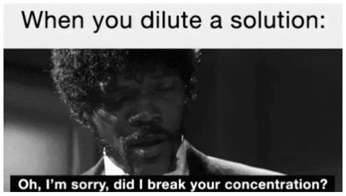
more on acids, bases and pH in CH 223...

End of Chapter Four Part 2



See also:

- *Chapter Four Part 2 Study Guide*
- *Chapter Four Part 2 Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*



MAR

Important Equations, Constants, and Handouts from this Chapter:

- Know how the **solubility guide** works
- Know what makes an acid acidic (and bases basic) and strong or weak; know how to use the pH scale
- Know how to write and determine net ionic equations and find spectator ions
- Know how to use molarity with solution stoichiometry problems
- Molarity (M) = mol of solute per Liter of solution
- $M_1V_1 = M_2V_2$

Solutions: **Solute, solvent, aqueous, electrolyte (strong, weak, non), solubility (use the Net Ionics solubility table), precipitation, types of reactions, molarity (M)**

Know the five types of reactions: precipitation, acid-base, gas forming, combustion and redox. Know how to determine if something has been oxidized or reduced (and the **oxidizing agent** and **reducing agent**)

MAR

End of Chapter Problems: ***Test Yourself***

1. Predict whether these compounds would be labeled as insoluble or soluble: HCl, NaCl, AgCl
2. Predict the products of this precipitation reaction and write the net ionic equation: $NiCl_2(aq) + (NH_4)_2S(aq) \rightarrow$? List any spectator ions.
3. In the following reaction, decide which reactant is oxidized and which is reduced. Designate the oxidizing agent and the reducing agent. **$Si(s) + 2\ Cl_2(g) \rightarrow SiCl_4(l)$**
4. Identify the ions and their concentration that exist in this aqueous solution: **0.25 M $(NH_4)_2SO_4$**
5. What volume of 0.109 M HNO_3, in milliliters, is required to react completely with 2.50 g of $Ba(OH)_2$? **$2\ HNO_3(aq) + Ba(OH)_2(s) \rightarrow 2\ H_2O(l) + Ba(NO_3)_2(aq)$**
6. A table wine has a pH of 3.40. What is the hydrogen ion concentration of the wine? Is it acidic or basic?
7. If 50.0 mL of 0.0135 M $BaCl_2$ is diluted to a total of 400. mL, what is the new concentration of $BaCl_2$?

MAR

End of Chapter Problems: ***Answers***

1. Soluble: HCl(aq), NaCl(aq). Insoluble: AgCl(s)
2. $NiCl_2(aq) + (NH_4)_2S(aq) \rightarrow NiS(s) + 2\ NH_4Cl(aq)$
   $Ni^{2+}(aq) + S^{2-}(aq) \rightarrow NiS(s)$ *Spectator ions:* NH_4^{+1} and Cl^{-1}
3. Si is oxidized and is the reducing agent; Cl_2 is reduced and is the oxidizing agent
4. 0.50 M NH_4^{+1}; 0.25 M SO_4^{2-}
5. 268 mL
6. acidic; $[H^+] = 4.0 \times 10^{-4}$ M
7. 0.00169 M

MAR

Energy & Chemical Reactions:
Thermochemistry
or
Thermodynamics

Chapter Five

Chemistry 221
Professor Michael Russell

Last update: 4/29/24

MAR

Energy & Chemistry

Burning peanuts supplies sufficient energy to boil water

Burning sugar (sugar reacts with $KClO_3$, a strong oxidizing agent)

Knowing if reactions give energy off or need energy to occur is important

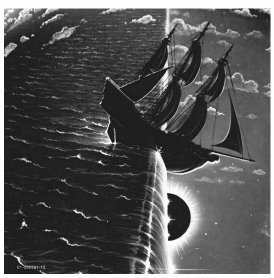
Thermochemistry or Thermodynamics: the study of energy exchange

MAR

CHEMICAL REACTIVITY

What drives chemical reactions? How do they occur, and how fast?

The first question is answered by THERMODYNAMICS, and the second question is answered by KINETICS.



MAR

CHEMICAL REACTIVITY

THERMODYNAMICS dictates if the reaction will occur or not.

Sand will not decompose into silicon and oxygen

Paper will combine with oxygen to burn

MAR

CHEMICAL REACTIVITY

KINETICS dictates how fast the reaction will occur.

***Example:* diamond into graphite is thermodynamically favored, but the kinetics of the reaction is too slow to be of useful**

MAR

CHEMICAL REACTIVITY

We have already seen a number of "driving forces" for reactions that are PRODUCT-FAVORED.

- **formation of a precipitate** *(precipitation)*
- **gas formation** *(gas forming)*
- **H_2O formation** *(acid-base)*
- **electron transfer** *(redox) i.e. in a battery*

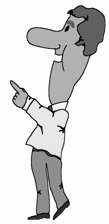
In general, reactions that transfer energy to their surroundings are product-favored.

MAR

All of thermodynamics depends on the law of CONSERVATION OF ENERGY:

- The total energy is unchanged in a chemical reaction.
- Energy expressed as heat and/or work. Chemists usually focus on heat (q) instead of work (w).
- First law of thermodynamics:

$$\Delta E = q + w = 0$$



Energy & Chemistry

MAR

UNITS OF ENERGY

1 calorie (cal) = heat required to raise temperature of **1.00 g of H_2O by 1.0 ^oC.**

1000 cal = 1 kilocalorie = 1 kcal

1 kcal = 1 Calorie (a *"food calorie"*)

But in our science classes we will use the unit called the JOULE (J) *(= kg m^2 / s^2)*

1 cal = 4.184 joules

1 Cal = 1000 cal = 4184 J

also: 1000 J = 1 kJ, etc.

James Joule 1818-1889

MAR

Memorize 4.184!

Potential Energy

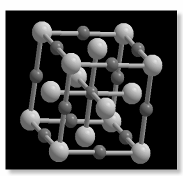
Two types of energy: **potential energy** *and* **kinetic** *energy*

Potential energy - energy due to position or composition, "stored energy"

In chemistry, **positive and negative particles (ions) attract one another - electrostatic energy.**

Also: **gravitational, nuclear**

Bungee jumper



NaCl - composed of Na^+ and Cl^- ions.

MAR

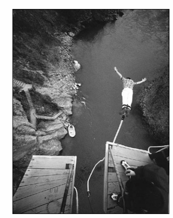
Kinetic Energy

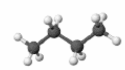
Kinetic energy: energy of motion. *Three forms of kinetic energy:*

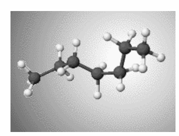
Translational ***- from physics, KE = $^1/_2$ mv^2 = 0.5(mass) (velocity)^2***

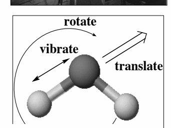
Vibrational

Rotational





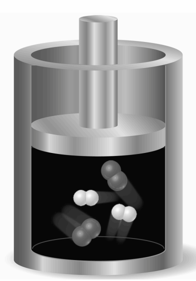




Kinetic energy depends on temperature

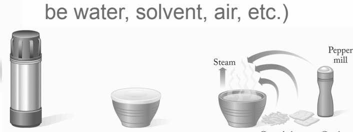
MAR

System *and* Surroundings



Piston with H_2 and O_2

- The **system** includes the molecules we want to study (here, the hydrogen and oxygen molecules).
- The **surroundings** are everything else (here, the cylinder and piston; could be water, solvent, air, etc.)



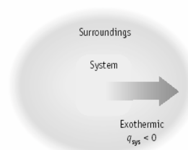
(a) **Isolated system:** A thermos bottle containing hot soup with the lid screwed on tightly (b) **Closed system:** A cup of hot soup with a lid (c) **Open system:** An open cup of hot soup

MAR

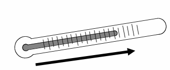
Directionality of Heat Transfer

Heat always transfers from the hot object to the cooler object.

EXOthermic: heat transfers from SYSTEM to SURROUNDINGS.



Exothermic: energy transferred from system to surroundings



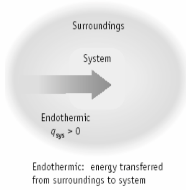
T(system) goes down
T(surr) goes up -
this is what we usually measure

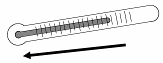
MAR

Directionality of Heat Transfer

Heat always transfers from the hot object to the cooler object.

ENDOthermic: heat transfers from SURROUNDINGS to the SYSTEM.





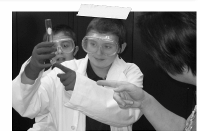
T(system) goes up
T(surr) goes down - *this is what we usually measure*

MAR

Enthalpy (ΔH)

- **Enthalpy** (ΔH) is heat (q) transferred at constant pressure (i.e. $\Delta H = q_P$)
- $\Delta H > 0$, ***Endothermic***; $\Delta H < 0$, ***Exothermic***
- Add subscripts to indicate ΔH for specific process i.e., ΔH_{vap}, ΔH_{rxn}, ΔH_f

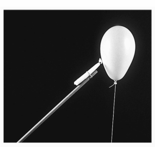
$\Delta E = q + w$ *and*
$\Delta E = \Delta H + w$ *(P const)*
w is usually small, so
$\Delta H \approx \Delta E$

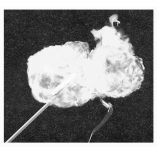


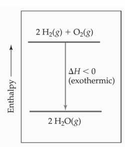
MAR

Enthalpy of Reaction

The quantity, ΔH, is called the enthalpy of reaction (ΔH_{rxn}), or the heat of reaction and is usually measured in kJ/mol





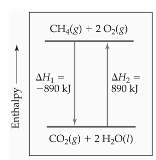


MAR

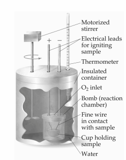
The Truth About Enthalpy

ΔH for a reaction in the forward direction is equal in size, but opposite in sign, to ΔH for the reverse reaction.

ΔH for a reaction depends on the state of the products and the state of the reactants.



We cannot know the exact enthalpy of the reactants and products, but we measure ΔH through **calorimetry**, the measurement of heat flow.

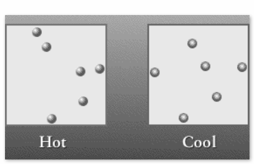


Bomb calorimeter

MAR

Calorimetry

Calorimetry is the measure of heat (energy) transfer

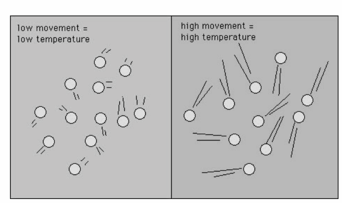


Heat energy is associated with molecular motions
(Kinetic Molecular Theory)

Heat transfers until thermal equilibrium is established

MAR

Heat & Matter *- no change in state*



When matter absorbs heat (q, in J), its temperature (ΔT) will rise depending on its mass (m, in g) and specific heat capacity (C):

$$q = mC\Delta T$$

MAR

$q = mC\Delta T$

Heat transferred with no phase change

q = heat *(or enthalpy)* **"lost" or "gained"**

m = **sample** mass *(usually grams)*

C = specific heat capacity

ΔT = **change in** temperature, **K** *or* **°C**

ΔT = *final* **T** - *initial* **T**

q **is** positive **when heat flows in (**cold surroundings**)** *- endothermic*

q **is** negative **when heat flows out (**hot surroundings**)** *- exothermic*

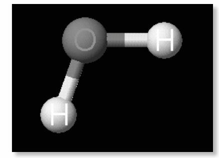
MAR

Specific Heat Capacity

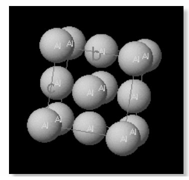
The heat (J) required to raise 1 g of an object by 1 ˚C (or K)

Memorize 4.184 for water

Water



Aluminum

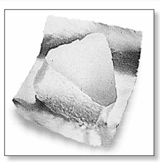


Substance	Spec. Heat (J/g•K)
$H_2O(l)$	4.184
Ethylene glycol(l)	2.39
Al(s)	0.902
glass(s)	0.84

MAR

Specific Heat Capacity

If 25.0 g of Al cool from 310. °C to 37 °C, how many joules of heat energy are lost by the Al?



$$q = mC\Delta T$$

$C_{Al} = 0.902\ J\ g^{-1}\ K^{-1}$

MAR

Specific Heat Capacity

If 25.0 g of Al cool from 310. °C to 37 °C, how many joules of heat energy are lost by the Al? $C_{Al} = 0.902\ J\ g^{-1}\ K^{-1}$

heat gain/lost = q = (mass)(sp. ht.)(ΔT)

where $\Delta T = T_{final} - T_{initial}$

q = (25.0 g)(0.902 J/g•˚C)(37 - 310.)˚C

q = - 6160 J

Notice also that ΔT can be either K or ˚C - the difference in temperatures is the same!

MAR

Specific Heat Capacity

A piece of iron (88.5 g) at 77.8 °C is placed in 244 g of water at 18.8 °C. What is the final temperature of the mixture?

By the law of conservation of energy:

$$q_{hot} + q_{cold} = 0, \text{ or}$$

$$m_{Fe}C_{Fe}\Delta T_{Fe} + m_{water}C_{water}\Delta T_{water} = 0$$

$$C_{Fe} = 0.449\ J\ g^{-1}\ K^{-1}$$

C_{water} = 4.184 J g^{-1} K^{-1} *Memorize*

Final Temperature (warm) same for Fe & H_2O

MAR

Specific Heat Capacity

Fe (88.5 g, 0.449 J/g•K, 77.8 °C) and **water** (244 g, 4.184 J/g•K, 18.8 °C); final temperature?

$$q_{hot} + q_{cold} = 0$$

$$m_{Fe}C_{Fe}\Delta T_{Fe} + m_{water}C_{water}\Delta T_{water} = 0$$

$$88.5 * 0.449 * (T_f - 77.8) + 244 * 4.184 * (T_f - 18.8) = 0$$

$$39.7T_f - 3090 + 1020T_f - 19200 = 0$$

$$1060T_f = 22300$$

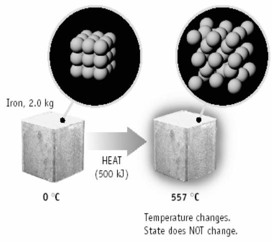
$$T_f = 21.0\ °C$$

Efficient method to determine approximate final temperature of mixture

MAR

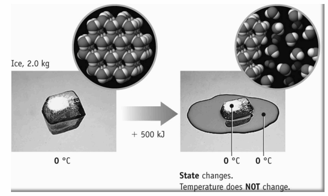
Heat Transfer No Change in State

q transferred = (mass)(sp. ht.)(ΔT)



MAR

Heat & Changes of State

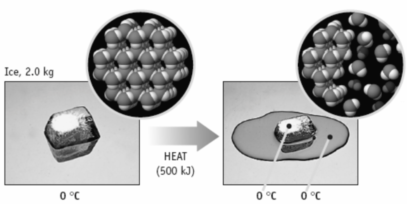


When matter absorbs heat, its temperature will rise until it undergoes a *Phase Change (solid to liquid, liquid to gas, solid to gas)*

The matter will continue to absorb energy, however during the phase change its temperature remains constant (no ΔT): Phase changes are "*Isothermal*" processes.

MAR

Heat Transfer with Change of State



Changes of state involve energy (at constant T)

Ice + 333 J/g (heat of fusion) -----> Liquid water

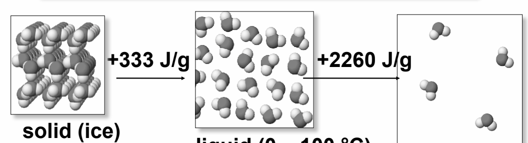
q = (heat of "something")("mass")

MAR

Heat & Changes of State

What quantity of heat is required to melt 500. g of ice at 0.0 °C and heat the water to steam at 100.0 °C?

Heat of fusion (ΔH_{fus}) for ice = 333 J/g
Specific heat of water = 4.184 J/g·K
Heat of vaporization (ΔH_{vap}) = 2260 J/g



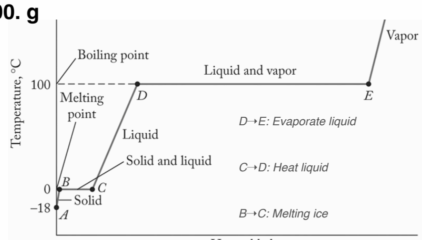
MAR

What quantity of heat is required to melt 500. g of ice at 0.0 °C and heat the water to steam at 100.0 °C?

Strategy: break this problem into **three calculations:**

- Melt ice *(phase change, ice - liquid, ΔH_{fus})*
- Heat liquid **(0 °C to 100 °C)**
- Evaporate liquid *(Phase change, liquid - steam, ΔH_{vap})*
- Mass is constant at **500. g**

Heat & Changes of State



MAR

Heat & Changes of State

What quantity of heat is required to melt 500. g of ice at 0.0 °C and heat the water to steam at 100.0 °C?

1. To <u>melt ice</u> at 0.0°C *(use ΔH_{fus})*

q = (500. g)(333 J/g) = 1.67×10^5 J

2. To <u>heat water</u> from 0.0 °C to 100.0 °C

q = (500. g)(4.184 J/g·K)(100. - 0.)K = 2.09×10^5 J

3. To <u>evaporate water</u> at 100. °C *(use ΔH_{vap})*

q = (500. g)(2260 J/g) = 1.13×10^6 J

4. Total heat energy = 1.51×10^6 J = 1510 kJ

Heat Transfer covered more in Chapter 10 of CH 222

MAR

Standard Enthalpy Values

Most ΔH values are labeled ΔH^o

Measured (^{o}) under standard conditions

P = 1 bar *(approx. 1 atm)*

Concentration = 1 mol/L

T = usually 25 ^{o}C

with all species in *standard states*

i.e., C = graphite and O_2 = gas

ΔH^0_x

Δ = change in
H = heat energy
0 = standard conditions
x = type of change

MAR

Thermochemical Equations

Thermochemical equations are regular chemical equations with an energy term.

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g) \quad \Delta H^o = -802$ kJ

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g) + 802$ kJ

Energy is a **product** just like CO_2 or H_2O! *(Exothermic)*

This provides new conversion factors!
From the equation:

$$\frac{+802 \text{ kJ of Energy Released}}{1 \text{ mol } CH_4(g) \text{ consumed}} \qquad \frac{+802 \text{ kJ of Energy Released}}{2 \text{ mol } H_2O\ (g) \text{ produced}}$$

MAR

Example: How many kJ of energy are released when 128.5 g of methane, CH_4(g), are combusted?

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g) \quad \Delta H^o = -802$ kJ

g → mols → J
↑ *molar mass* ↑ *Reaction enthalpy*

$$128.5\text{g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.04 \text{ g}} \times \frac{1 \text{ mol rxn}}{1 \text{ mol } CH_4} \times \frac{-802 \text{ kJ}}{1 \text{mol rxn}} = -6.43 \times 10^3 \text{ kJ}$$

MAR

Enthalpy Values

ΔH values depend on how the reaction is written and on phases of reactants and products

$H_2(g) + 1/2\ O_2(g) \text{ --> } H_2O(g)$
$\Delta H^\circ_1 = -242$ kJ

$2\ H_2(g) + O_2(g) \text{ --> } 2\ H_2O(g)$
$\Delta H^\circ_2 = -484$ kJ $= 2^* \Delta H^\circ_1$

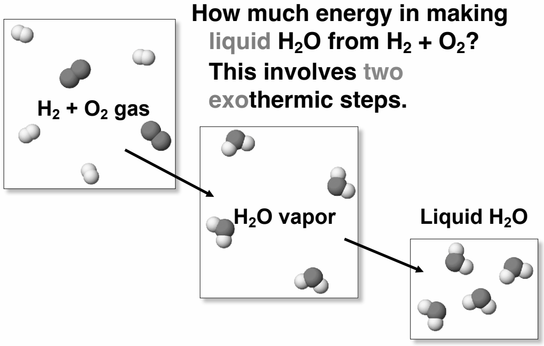
$H_2O(g) \text{ ---> } H_2(g) + 1/2\ O_2(g)$
$\Delta H^\circ_3 = +242$ kJ $= -(\Delta H^\circ_1)$

$H_2(g) + 1/2\ O_2(g) \text{ --> } H_2O(\text{liquid})$
$\Delta H^\circ_4 = -286$ kJ $\neq \Delta H^\circ_1$

MAR

USING ENTHALPY

How much energy in making liquid H_2O from $H_2 + O_2$? This involves two exothermic steps.



MAR

USING ENTHALPY

Making H_2O from H_2 involves two steps.

$H_2(g) + 1/2\ O_2(g) \text{ ---> } \cancel{H_2O(g)} + 242$ kJ

$\cancel{H_2O(g)} \text{ ---> } H_2O(liq) + 44$ kJ

$H_2(g) + 1/2\ O_2(g) \text{ --> } H_2O(liq) + 286$ kJ

Example of HESS'S LAW-

If a rxn. is the sum of 2 or more others, the net ΔH is the sum of the ΔH's of the other rxns.

Mr. Hess

MAR

Hess's Law Problem

Example: Determine the $\Delta H°$ for the reaction:

$3\ H_2(g) + N_2(g) \rightarrow 2\ NH_3(g)$ $\Delta H°_3 = ???$

Given the following:

(1) $2\ H_2(g) + N_2(g) \rightarrow N_2H_4(g)$ $\Delta H°_1 = +95.4$ kJ

(2) $N_2H_4(g) + H_2(g) \rightarrow 2NH_3(g)$ $\Delta H°_2 = –187.6$ kJ

Desired equation has no N_2H_4, so try to remove by adding equations (1) & (2) together:

$$2\ H_2(g) + N_2(g) + \cancel{N_2H_4(g)} + H_2(g) \rightarrow \cancel{N_2H_4(g)} + 2NH_3(g)$$

Combine like terms, reduce to:

$$3\ H_2(g) + N_2(g) \rightarrow 2\ NH_3(g) \quad \Delta H°_3$$

And get:

$$\Delta H°_3 = \Delta H°_1 + \Delta H°_2 = +95.4 \text{ kJ} + (–187.6 \text{ kJ}) = –92.2 \text{ kJ}$$

MAR

Standard Enthalpy Values

Compound	ΔH_f (kJ/mol)	Compound	ΔH_f (kJ/mo
AgBr(s)	-99.5	$C_2H_2(g)$	+226.7
AgCl(s)	-127.0	$C_2H_4(g)$	+52.3
AgI(s)	-62.4	$C_2H_6(g)$	-84.7
$Ag_2O(s)$	-30.6	$C_3H_8(g)$	-103.8
$Ag_2S(s)$	-31.8	n-$C_4H_{10}(g)$	-124.7
$Al_2O_3(s)$	-1669.8	n-$C_5H_{12}(l)$	-173.1
$BaCl_2(s)$	-860.1	$C_2H_5OH(l)$	-277.6

NIST (Nat'l Institute for Standards and Technology) gives values of

$\Delta H_f°$ = standard molar enthalpy of formation

- the enthalpy change when 1 mol of compound is formed from elements under standard conditions.

These values are available in your text

All elements in standard states have $\Delta H_f° = 0$ (graphite but not diamond, $O_2(g)$ but not O_3 or $O_2(l)$, etc.)

MAR

Question: What is the standard molar enthalpy of formation equation for potassium permanganate?

elements compound

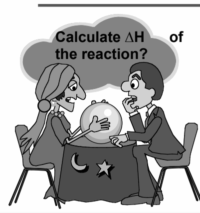
$$K(s) + Mn(s) + 2\ O_2(g) \longrightarrow KMnO_4(s)$$

- salts, metals: solids at *standard state conditions*. oxygen is a gas
- balance for one mole of the product; reactants elements in their standard states

$KMnO_4(aq)$ and $KMnO_4(s)$

MAR

Using Standard Enthalpy Values



In general, when ALL enthalpies of formation are known:

$$\Delta H°_{rxn} = \Sigma\ n\Delta H_f°\ (\text{products}) - \Sigma\ n\Delta H_f°\ (\text{reactants})$$

Σ = summation sign*, or "add up all of the"*

n = stoichiometric coefficients

"Energy gained" - "Energy spent"

MAR

Using Standard Enthalpy Values

Example: **Find $\Delta H°_{rxn}$ for**

$CaCO_3(s) \dashrightarrow CaO(s) + CO_2(g)$

using:

$$\Delta H°_{rxn} = \Sigma\ n\Delta H_f°\ (\text{prod}) - \Sigma\ n\Delta H_f°\ (\text{react})$$

Answer:

$\Delta H°_{rxn} = \{\Delta H_f°(CaO) + \Delta H_f°(CO_2)\} - \{\Delta H_f°(CaCO_3)\}$

$\Delta H°_{rxn} = \{-635.1 + -393.5\} - \{-1206.2\}$

$\Delta H°_{rxn} = +177.6$ kJ

All stoichiometries (n) are one in this example

MAR

Using Standard Enthalpy Values

Calculate the heat of combustion of methanol, i.e., $\Delta H°_{rxn}$ for

$CH_3OH(g) + 3/2\ O_2(g) \dashrightarrow CO_2(g) + 2\ H_2O(g)$

$$\Delta H°_{rxn} = \Sigma\ n\Delta H_f°\ (\text{prod}) - \Sigma\ n\Delta H_f°\ (\text{react})$$

As before, look up $\Delta H_f°$ values for reactants and products in your text

Elements in standard states have $\Delta H_f° = 0$

MAR

Using Standard Enthalpy Values

$CH_3OH(g) + 3/2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$

$\Delta H^o_{rxn} = \Sigma\ \Delta H_f^o\ (prod) - \Sigma\ \Delta H_f^o\ (react)$

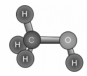
$\Delta H^o_{rxn} = \{\Delta H_f^o\ (CO_2) + 2\ \Delta H_f^o\ (H_2O)\}$

$- \{3/2\ \Delta H_f^o\ (O_2) + \Delta H_f^o\ (CH_3OH)\}$

$= \{(-393.5\ kJ) + 2\ (-241.8\ kJ)\}$

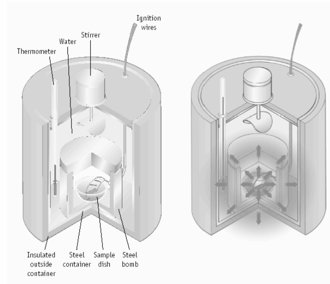
$- \{0 + (-201.5\ kJ)\}$

$\Delta H^o_{rxn} = -675.6\ kJ$ **per mol of methanol**



MAR

CALORIMETRY
Measuring Heats of Reaction

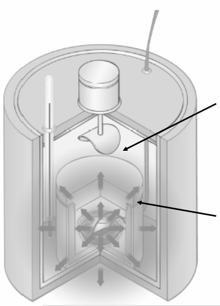


Constant Volume "Bomb" Calorimeter

- **Burn combustible sample.**
- **Measure heat evolved in a reaction.**
- **Derive ΔE for reaction.**

MAR

Calorimetry



Some heat from reaction warms water
q_{water} **= (water mass)(sp. ht.)(ΔT)**

Some heat from reaction warms "bomb"
q_{bomb} **= (heat capacity, J/K)(ΔT)**

Total heat evolved = $q_{total} = q_{water} + q_{bomb}$

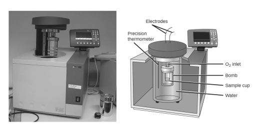
MAR

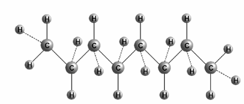
Measuring Heats of Reaction: CALORIMETRY

Calculate heat of combustion of octane:

$C_8H_{18} + {}^{25}/_2\ O_2 \rightarrow 8\ CO_2 + 9\ H_2O$

- **Burn 1.00 g of octane**
- **Temp rises from 25.00 to 33.20 ^oC**
- **Calorimeter contains 1200. g water**
- **Heat capacity of bomb = 837 J/K**





MAR

Measuring Heats of Reaction: CALORIMETRY

Step 1: *Calculate* heat transferred from reaction to water.

q = (1200. g)(4.184 J/g•K)(8.20 K) = 41,200 J

Step 2: *Calculate* heat transferred from reaction to bomb

q = (bomb heat capacity)(ΔT)

= (837 J/K)(8.20 K) = 6860 **J**

Step 3: *Calculate* total heat ***absorbed***

41,200 J + 6860 J = 48,100 J * kJ/10^3 J = 48.1 kJ

Step 4: Total heat ***evolved*** **= -(**Total heat ***absorbed*****) =** - 48.1 kJ

Heat of combustion *for 1.00 g of octane* **=** (-48.1 kJ / 1.00 g) * (114.23 g / mol C_8H_{18}) **=** $-5.49*10^3$ **kJ/mol**

MAR

octane is C_8H_{18}, 114.23 g/mol

Quick & Dirty Thermochemistry!

memorize 4.184 J/g°C *heat capacity liquid water*

q = mCΔT *heat transfer no change in phase*

q = "heat"x"mass" *heat transfer with change in phase*

be able to use Hess's Law

understand Heat of Formation, ΔH_f^o *(1 mol product, reactants are elements in standard states)*

understand how to use:

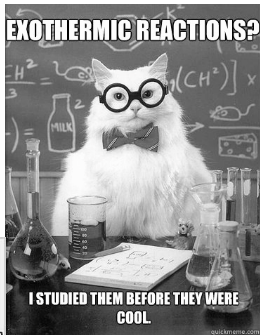
$\Delta H^o_{rxn} = \Sigma\ n\Delta H_f^o\ (prod) - \Sigma\ n\Delta H_f^o\ (react)$

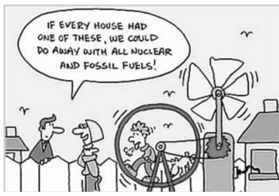
MAR

End of Chapter Five

See also:

- *Chapter Five Study Guide*
- *Chapter Five Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*





MAR

Important Equations, Constants, and Handouts from this Chapter:

Thermodynamics:

- **endothermic & exothermic**
- **product & reactant favored**
- **Law of Conservation of Energy**
- **units of energy (J, kJ, cal, etc.)**
- **system & surroundings**
- **heat & work**
- **enthalpy**
- **calorimetry**
- **heat of reaction and heat of formation**
- **standard enthalpy conditions**
- **Hess's Law**

$\Delta E = q + w = 0$

$q = mC\Delta T$ **(no phase change)**

q = (heat of "something") ("mass") (phase change)

memorize $C_{water(liquid)} = 4.184\ J\ g^{-1}\ K^{-1}$

$$\Delta H^o_{rxn} = \Sigma\, n\Delta H_f^o \text{ (products)} - \Sigma\, n\Delta H_f^o \text{ (reactants)}$$

MAR

End of Chapter Problems: ***Test Yourself***

1. The initial temperature of a 344 g sample of iron is 18.2 °C. If the sample absorbs 2.25 kJ of heat, what is its final temperature? C_{Fe} = 0.449 J/g K
2. One beaker contains 156 g of water at 22 °C and a second beaker contains 85.2 g of water at 95 °C. The water in the two beakers is mixed. What is the final water temperature?
3. What quantity of heat is required to vaporize 125 g of benzene, C_6H_6, at its boiling point, 80.1 °C? The heat of vaporization of benzene is 30.8 kJ/mol.
4. The enthalpy changes for the following reactions can be measured:
   $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$ ΔH° = -802.4 kJ
   $CH_3OH(g) + 3/2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$ ΔH° = -676 kJ

Use these values and Hess's law to determine the enthalpy change for the reaction: $CH_4(g) + 1/2\ O_2(g) \rightarrow CH_3OH(g)$

5. Write a balanced chemical equation for the formation of $Li_2CO_3(s)$ from the elements in their standard states. Find the value of $\Delta H_f°$ for $Li_2CO_3(s)$ in a table of values.
6. $Ca(OH)_2$ reacts slowly with CO_2 to give $CaCO_3$: $\mathbf{Ca(OH)_2(s) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(g)}$ Calculate the standard enthalpy change for this reaction.

MAR

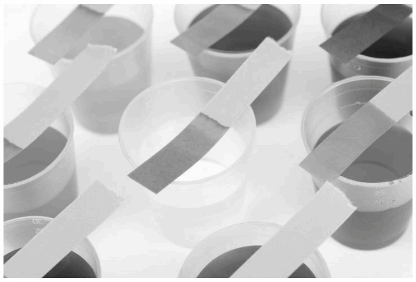
End of Chapter Problems: ***Answers***

1. 306.0 K (32.8 ºC)
2. 321 K (48 ºC)
3. 49.3 kJ
4. -126 kJ
5. $2\ Li(s) + C(s) + 3/2\ O_2(g) \rightarrow Li_2CO_3(s)$ ΔH_f^o = –1216.04 kJ
6. ΔH^o_{rxn} = –83.1 kJ

MAR

Chemistry 221 Exam II Review

Chapters 3, 4 and 5



Chemistry 221
Professor Michael Russell

Last update: 4/29/24

MAR

____ $H_2S(g)$ + ____ $SO_2(g)$ → ____ $S(s)$ + ____ $H_2O(g)$

Which statement regarding this reaction is true?

A. 3 moles of S are produced per mole of H_2S.
B. 1 mole of SO_2 is consumed per mole of H_2S.
C. 1 mole of H_2O is produced per mole of H_2S.
D. The total number of moles of products is always equal to the total number of moles of reactants used.
E. None of these statements are true.

MAR

What is the balanced equation for the combustion of butane, C_4H_{10}?

A. $C_4H_{10}(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
B. $2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$
C. $C_4H_{10}(g) + 13\ O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)$
D. $C_4H_{10}(g) + 9\ O_2(g) \rightarrow 4\ CO_2(g) + 10\ H_2O(g)$

MAR

In the reaction of 2.0 mol of CCl_4 with an excess of HF, 1.7 mol of CCl_2F_2 is obtained.

$$CCl_4(l) + 2\ HF(g) \rightarrow CCl_2F_2(l) + 2\ HCl(g)$$

Which statement is true here?

A. The theoretical yield for CCl_2F_2 is 1.7 mol.
B. The actual yield for CCl_2F_2 is 1.0 mol.
C. The percent yield for the reaction is 85%.
D. Theoretical yield cannot be determined unless the exact amount of HF used is known.
E. Infinite diversity in infinite combinations (IDIC)

MAR

Burning sulfur in an atmosphere of fluorine produces the very stable compound SF_6.

$$S_8(s) + 24\ F_2(g) \rightarrow 8\ SF_6(g)$$

If you wish to produce 2.50 moles of SF_6, you will need to use:

A. 0.313 moles of S_8 and 7.50 moles of F_2.
B. 1.00 moles of S_8 and 24.0 moles of F_2.
C. 0.125 moles of S_8 and 3.00 moles of F_2.
D. 8.00 moles of S_8 and 24.0 moles of F_2.
E. More information is required to answer this question.

MAR

Ammonia is prepared by the reaction:

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

If 10.0 mol of N_2 are mixed with 25.0 mol of H_2, the amount of NH_3 produced will be:

A. 20.0 mol NH_3
B. 16.7 mol NH_3
C. 37.5 mol NH_3
D. 25.0 mol NH_3
E. 35.0 mol NH_3

MAR

A compound with C, H and O is found through combustion analysis of a 0.255 g sample to give 0.561 g CO_2 and 0.306 g H_2O; it also has a molar mass of 60.1 g/mol. What is the molecular formula?

A. CH_3CO_2H
B. $C_4H_9O_3$
C. C_3H_6O
D. $C_9H_7O_3$
E. C_3H_8O

MAR

Which of the following is the only insoluble salt in water?

A. NH_4NO_3
B. NaOH
C. PbI_2
D. K_2CO_3
E. LiCl

MAR

Which of the compounds below is *not* an acid in aqueous solution?

A. CH_3CO_2H
B. H_3PO_4
C. NH_3
D. HCl
E. $HClO_4$

MAR

Which equation below best represents the balanced, net ionic equation for the reaction of magnesium carbonate with nitric acid?

A. $MgCO_3(s) + 2\ HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
B. $MgCO_3(s) + 2\ H^+(aq) \rightarrow Mg^{2+}(aq) + CO_2(g) + H_2O(l)$
C. $Mg^{2+}(aq) + 2\ NO_3^-(aq) \rightarrow Mg(NO_3)_2(s)$
D. $MgCO_3(s) + 2\ HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2CO_3(aq)$
E. More information is required to answer this question.

MAR

Which equation below best represents the balanced net ionic equation for the reaction of potassium hydroxide and iron(II) chloride to give iron(II) hydroxide and potassium chloride?

A. $2\ KOH(aq) + FeCl_2(aq) \rightarrow Fe(OH)_2(s) + 2\ KCl(aq)$
B. $2\ KOH(aq) + FeCl_2(aq) \rightarrow Fe(OH)_2(aq) + 2\ KCl(aq)$
C. $2\ OH^-(aq) + Fe^{2+}(aq) \rightarrow Fe(OH)_2(s)$
D. $K^+(aq) + Cl^-(aq) \rightarrow KCl(aq)$
E. More information is required to answer this question.

MAR

Which of the following statements is correct regarding the reaction of Zn with VO_2^+?

$$Zn(s) + 4\ H^+(aq) + 2\ VO_2^+(aq) \rightarrow Zn^{2+}(aq) + 2\ VO^{2+}(aq) + 2\ H_2O(l)$$

A. Zn is oxidized and VO_2^+ is the reducing agent.
B. Zn is reduced and VO_2^+ is the reducing agent.
C. Zn is oxidized and VO_2^+ is the oxidizing agent.
D. Zn is reduced and VO_2^+ is the oxidizing agent.
E. This is not a redox reaction.

MAR

Assume you dissolve 6.73 g Na_2CO_3 in enough water to make 250. mL of solution. (Molar mass of Na_2CO_3 = 106 g/mol.) What is the concentration of the sodium carbonate?

A. 26.9 M
B. 0.0635 M
C. 0.254 M
D. 0.762 M
E. 42 M

MAR

60.0 mL of 0.25 M HCl are added to a 500. mL volumetric flask; water is added to the mark on the flask. What is the concentration of HCl in the diluted solution?

A. 0.015 M
B. 0.025 M
C. 0.030 M
D. 0.060 M
E. 0.050 M

MAR

What is the pH of dilute nitric acid with a concentration of 0.030 M?

A. 0.030
B. 1.52
C. 1.82
D. 2.50
E. 3.00

MAR

What mass of Na_2CO_3 (molar mass = 106.0 g/mol) is required for complete reaction with 25.0 mL of 0.155 M HNO_3?

$Na_2CO_3(aq) + 2\ HNO_3(aq) \rightarrow 2\ NaNO_3(aq) + CO_2(g) + H_2O(l)$

A. 0.410 g
B. 205 g
C. 0.205 g
D. 0.122 g
E. 37 kg

MAR

A piece of copper (5.00 g) is heated for 2.0 seconds, and 100. J of heat energy is transferred to the copper. The temperature increases from 20.0 °C to 71.9 °C. Calculate the specific heat capacity of copper.

A. 0.278 J/g•K
B. 0.385 J/g•K
C. 1.93 J/g•K
D. 2.60 J/g•K
E. -0.977 J/g•K

MAR

When 108 grams of water at 22.5 °C are mixed with 65.1 grams of water at an unknown temperature, the final temperature of the mixture is 47.9 °C. What was the initial temperature of the other sample of water?

A. 8.9 °C
B. 79.7 °C
C. 67.0 °C
D. 90.0 °C
E. 274 °C

MAR

The standard molar enthalpy of combustion for propane is -2044 kilojoules.

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$

What is the standard enthalpy change for the combustion of 3.000 mol of propane (C_3H_8)?

A. -6132 kJ
B. -2044 kJ
C. -4088 kJ
D. +2044 kJ
E. +6132 kJ

MAR

Calculate the enthalpy for the reaction

$SiH_4(g) + 2\ O_2(g) \rightarrow SiO_2(g) + 2\ H_2O(g)$

using these values:

$\Delta H^\circ_f[SiH_4(g)]$ = +34.3 kJ/mol;
$\Delta H^\circ_f[SiO_2(g)]$ = -910.9 kJ/mol; and
$\Delta H^\circ_f[H_2O(g)]$ = -241.8 kJ/mol

A. -1187.0 kJ/rxn
B. -1428.8 kJ/rxn
C. -1360.2 kJ/rxn
D. -2218.7 kJ/rxn
E. Not enough information

MAR

Which equation below defines the standard molar enthalpy of formation of gaseous methanol, CH_3OH?

A. $CH_4(g) + ½\ O_2(g) \rightarrow CH_3OH(g)$
B. $C(s) + 2\ H_2(g) + ½\ O_2(g) \rightarrow CH_3OH(g)$
C. $CO(g) + 2\ H_2(g) \rightarrow CH_3OH(g)$
D. $H_2O(g) + C(s) + H_2(g) \rightarrow CH_3OH(g)$
E. You'll go blind if you drink methanol! Who cares! :)

MAR

Calculate the standard molar enthalpy of formation for $FeCl_2(s)$ using the following:

$½\ Cl_2(g) + FeCl_2(s) \rightarrow FeCl_3(s)$ $\quad \Delta H^\circ_r$ = -57.7 kJ/rxn
$Fe(s) + {}^3/_2\ Cl_2(g) \rightarrow FeCl_3(s)$ $\quad \Delta H^\circ_f$ = -399.5 kJ/rxn

A. -57.7 kJ/mol
B. -341.8 kJ/mol
C. -284.1 kJ/mol
D. -457.2 kJ/mol
E. 42 kJ/mol

MAR

End of Review - *good luck with your studying!*

Need more practice?
- *Practice Problem Sets (online)*
- *Concept Guides (Companion and online)*
- *Chapter Guides (online)*
- *End of Chapter Problems in Textbook (every other question has answer at end)*

Good luck with your studying!

MAR

$H\Psi = E\Psi$

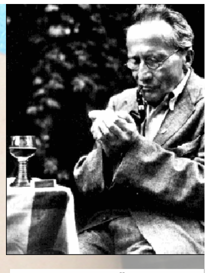
The Structure of Atoms

Chapter 6 Part 1

Chemistry 221

Professor Michael Russell

MAR *Last update: 4/29/24*

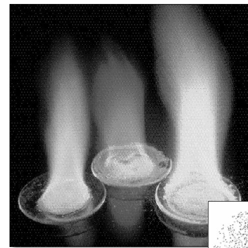


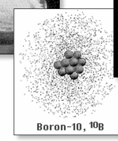
$$\Psi(1s) = 2a_0^{-1.5}e^{\frac{-r}{a_0}}$$

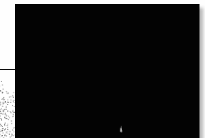
$$\Psi(2s) = \frac{1}{\sqrt{8}}a_0^{-1.5}(2 - \frac{r}{a_0})e^{\frac{-r}{2a_0}}$$

$$\Psi(2p) = \frac{1}{\sqrt{24}}a_0^{-1.5}(\frac{r}{a_0})e^{\frac{-r}{2a_0}}$$

Atomic Structure







Exploring color.... on the atomic level

MAR

ELECTROMAGNETIC RADIATION

MAR

ELECTROMAGNETIC RADIATION

EM theory developed by James Maxwell using **partial differentials** *(calculus!)* in 1860s

Most subatomic particles behave as PARTICLES **and yet often obey the physics of** WAVES. **(!)**

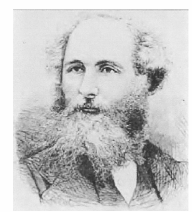
Define properties of WAVES:

Wavelength, λ

Frequency, ν

Node

Amplitude



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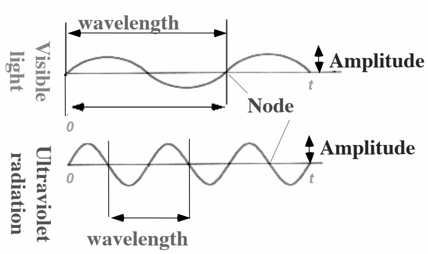
ELECTROMAGNETIC RADIATION



Frequency = 1.5 waves / *t*

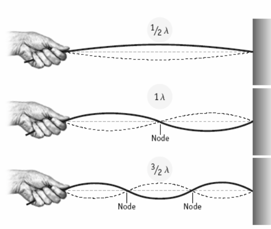
Frequency = 3.0 waves / *t*

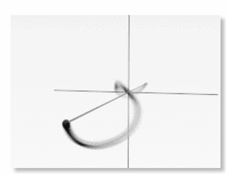
Note:

Wavelength ↑, frequency ↓, *and*
Wavelength ↓, frequency ↑

MAR

ELECTROMAGNETIC RADIATION





A node in a standing wave

MAR

ELECTROMAGNETIC RADIATION

Waves have a frequency

Use the Greek letter "nu", ν, for frequency, and units are "cycles per sec"

All radiation: $\nu \cdot \lambda = c$

where c = velocity of light = 2.998 x 10^8 m/sec

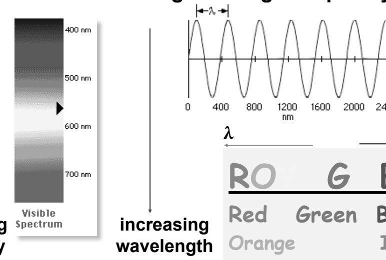
Note: **long wavelength --> small frequency**
short wavelength --> high frequency

Memorize 2.998 x 10^8 m/sec! Always use this value for c!

MAR

Visible Light / EM Radiation

Note: **long wavelength --> small frequency**
short wavelength --> high frequency



increasing frequency (υ)

increasing wavelength (λ)

λ υ

RO G BIV

Red Green Blue
Orange Indigo
Violet

MAR

ELECTROMAGNETIC RADIATION

Red light has λ = 700. nm. Calculate the frequency.

$$700.\ \text{nm} \bullet \frac{1 \times 10^{-9}\ \text{m}}{1\ \text{nm}} = 7.00 \times 10^{-7}\ \text{m}$$

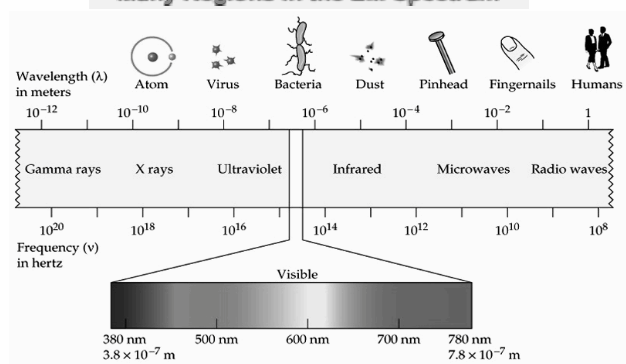
Recall: $\nu = c / \lambda$

$$\text{Freq} = \frac{2.998 \times 10^{8}\ \text{m/s}}{7.00 \times 10^{-7}\ \text{m}} = 4.28 \times 10^{14}\ \text{sec}^{-1}$$

Frequency = 4.28 * 10^{14} s^{-1} *or*
4.28 * 10^{14} Hz

MAR

Many Regions in the EM Spectrum



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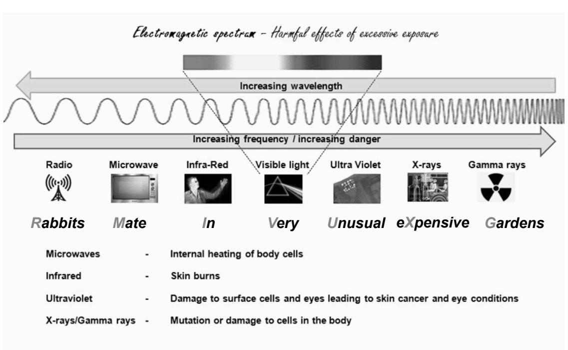

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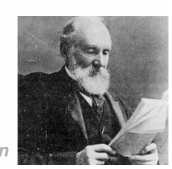
Lord Kelvin's 1900 "Clouds" Speech

In **1900**, **Lord Kelvin** stated that current thermodynamic understanding explained all energy phenomenon except for two not yet understood "clouds":

- **the failure of the Michelson-Morley experiment** *(which led to special relativity)*
- **the inability to understand black body radiation** *(which led to quantum theory)*

"Pride goeth before a fall"

Great scientists make mistakes as well

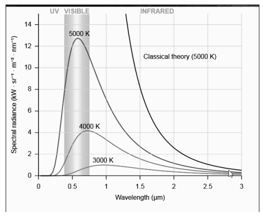


Lord Kelvin

MAR

Ultraviolet Catastrophe

At high temperatures, solids emit **red**, **blue**, even white light when heated. Energy of light emitted relative to temperature.
At very high temperature, intensity of light reaches maximum in ultraviolet region, then decreases.
Classical physics predicted no maximum intensity - *catastrophe!*



MAR

Quantization of Energy

Max Planck proposed that an object can gain or lose energy by absorbing or emitting radiant energy in QUANTA
Proposed that the energy of radiation proportional to the frequency:

$$E = h \bullet v$$

where h = Planck's constant = 6.626 x 10^{-34} J•s *memorize*

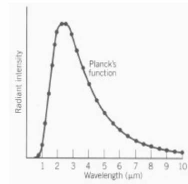


FIGURE 3.13 Planck's function fits the observed data perfectly.

Memorize 6.626 x 10^{-34} J•sec! Always use this value for h!

MAR

Quantization of Energy

$$E = h \bullet v = hc / \lambda$$

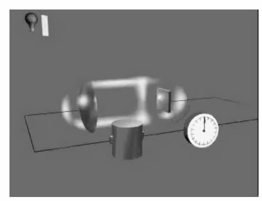
Light with large λ (small ν) has a small E.
Light with a short λ (large ν) has a large E.

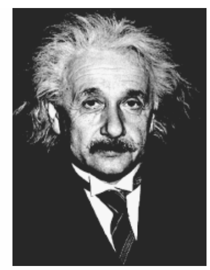
h = 6.626 x 10^{-34} J•s

MAR

Photoelectric Effect

Light applied to metal; electrons emitted as long as threshold frequency maintained.
Elimination of light halts the process





Albert Einstein (1879-1955) explained phenomenon Received Nobel Prize

MAR

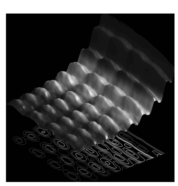
Photoelectric Effect

Photoelectric effect experiment shows *particle* nature of light.

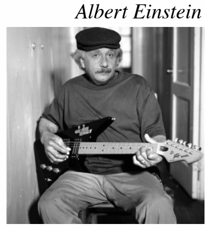
Classical physics said E of ejected e^- should *increase* as light intensity increases - not observed!

No e^- observed until light of a minimum E (or ν) is used.

Light said to display "wave-particle duality" - it behaves like a wave in some experiments *(diffraction, interference)* but as a particle in others *(photoelectric effect)!!!*



Albert Einstein



MAR

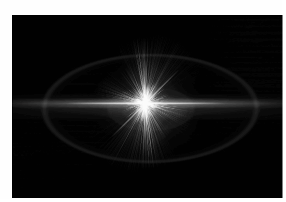
Photoelectric Effect

Experimental observations understood if light consists of particles called PHOTONS with discrete energy.

***PROBLEM:* Calculate the energy of 1.00 mol of photons of red light (λ = 700. nm)**

From earlier:

λ = 700. nm
ν = 4.28 x 10^{14} sec^{-1}



MAR

Energy of Radiation

PROBLEM: Calculate the energy of 1.00 mol of photons of red light.

λ = 700. nm

ν = 4.28×10^{14} sec^{-1}

E = h•ν

= $(6.626 \times 10^{-34}$ J•s$)(4.28 \times 10^{14}$ sec$^{-1})$

= 2.84×10^{-19} J per *photon*

MAR

Energy of Radiation

Energy of 1.00 mol of photons of red light.

E = h • ν

= $(6.626 \times 10^{-34}$ J•s$)(4.28 \times 10^{14}$ sec$^{-1})$

= 2.84×10^{-19} J per photon

E per mol =

$(2.84 \times 10^{-19}$ J/ph$)(6.022 \times 10^{23}$ ph/mol$)$

= 171,000 J/mol * (kJ / 1000 J)

= 171 kJ/mol

This is within the range of energies that can break bonds.

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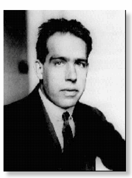
Atomic Line Spectra and Niels Bohr



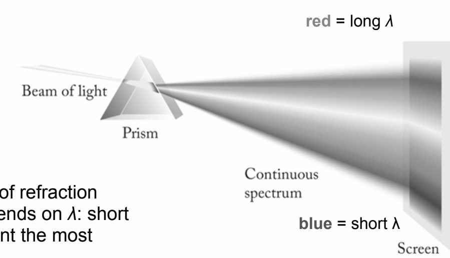
Niels Bohr

(1885-1962)

Bohr's greatest contribution to science was in building a simple model of the atom. His model was based on an understanding of the SHARP LINE SPECTRA of excited atoms.

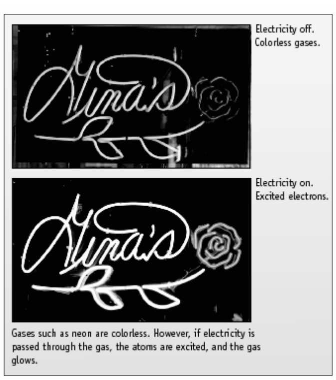
MAR

Refractive Spectrum of White Light

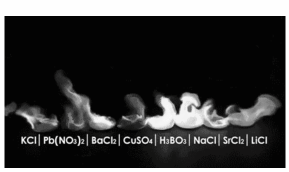


Angle of refraction depends on λ: short λ bent the most

MAR



Excited Gases & Atomic Structure



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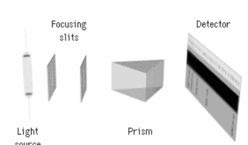
Line Spectra of Excited Atoms

Excited atoms emit light of only certain wavelengths

The wavelengths of emitted light depend on the element.



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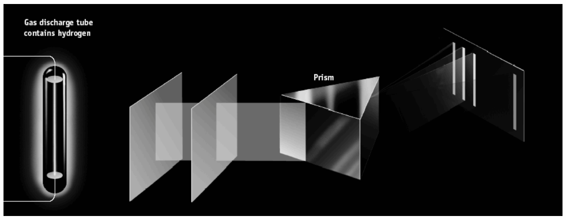
Spectrum of Excited Hydrogen Gas



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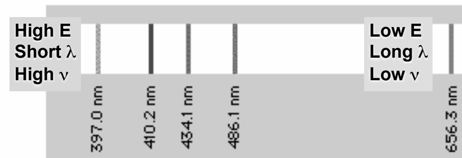
Line Emission Spectra of Excited Atoms



Visible lines in H atom spectrum are called the BALMER series.

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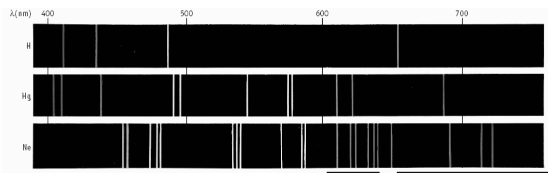
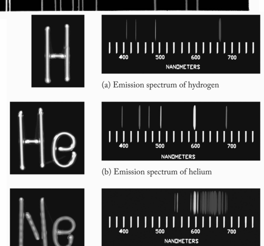


Line Spectra of Other Elements



(a) Emission spectrum of hydrogen

(b) Emission spectrum of helium

(c) Emission spectrum of neon

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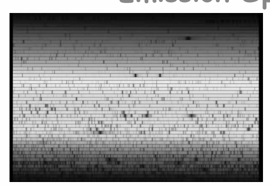
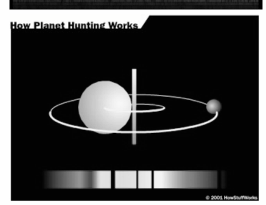
Emission Spectra in Astronomy





Composition of stars and stellar objects determined through emission spectrographs

Astronomers must account for red and blue shifts (the "Doppler effect") of moving objects in emission spectra

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The Electric Pickle

Excited atoms can emit light.

Here the solution in a pickle is excited electrically. The Na^+ ions in the pickle juice give off light characteristic of that element.

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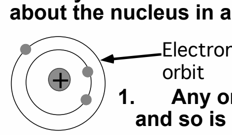
Atomic Spectra and Bohr

One view of atomic structure in early 20th century was that an electron (e-) traveled about the nucleus in an orbit.



1. Any orbit should be possible and so is any energy.
2. But a charged particle moving in an electric field should emit energy.

End result should be destruction - according to classical physics!

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Atomic Spectra and Bohr

Bohr said classical view is wrong.

Need a new theory - now called QUANTUM or WAVE MECHANICS.

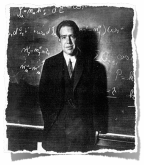
e- can only exist in certain discrete orbits - called stationary states.

e- is restricted to QUANTIZED energy states.

Energy of state = - Rhc/n^2

n = quantum no. = 1, 2, 3, 4, ….

(R = Rydberg constant, 1.097*10^7 m^{-1})



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Atomic Spectra and Bohr

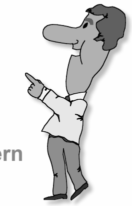
Energy of quantized state = - Rhc/n^2

n = quantum no. = 1, 2, 3, 4, ….

R = Rydberg constant = 1.097*10^7 m^{-1}

h = Planck's constant = 6.626*10^{-34} J s

c = speed of light = 2.998*10^8 m s^{-1}



But note - same eqns. come from modern wave mechanics approach.

Results can be used to explain atomic spectra.

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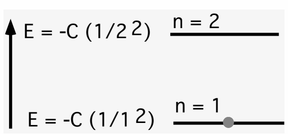
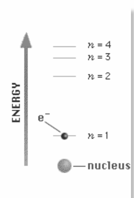
Atomic Spectra and Bohr

If e-'s are in quantized energy states, then ΔE of states can have only certain values. This explain sharp line spectra.





Note that C = Rhc

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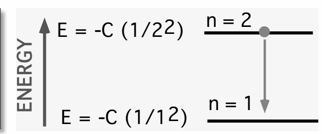
Energy of quantized state = - Rhc/n^2

Atomic Spectra and Bohr



Calculate ΔE for 1 mol of e- "falling" from high energy level (n = 2) to low energy level (n = 1). ***(L = Avogadro's number)***

$\Delta E = E_{final} - E_{initial} = -RhcL[(1/n_f^2) - (1/n_i^2)]$

$= -RhcL[(1/1^2) - (1/2^2)] = -RhcL[(1) - (1/4)]$

ΔE = -(3/4)RhcL = - 984 kJ/mol

Note that the process is exothermic!

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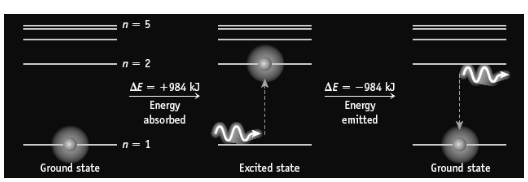
R = 1.097 x 10^7 m^{-1}



ΔE = -(3/4)Rhc = -984 kJ/mol

What is the ν and λ of the emitted light photon?

ν = -984*10^3 J/mol / (h * 6.022*10^{23} mol^{-1})

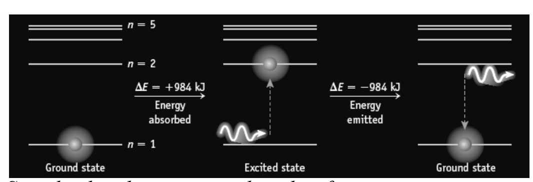
ν = 2.47*10^{15} s^{-1} ***(always positive!)***

and $\lambda = c/\nu$ = 122 nm

This is exactly in agreement with experiment!

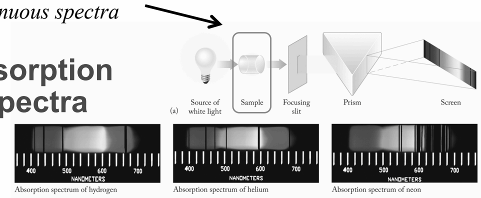
Remember: absorption = endothermic, emission = exothermic

MAR



Sample absorbs some wavelengths of light, causing dark lines in the continuous spectra

Absorption Spectra



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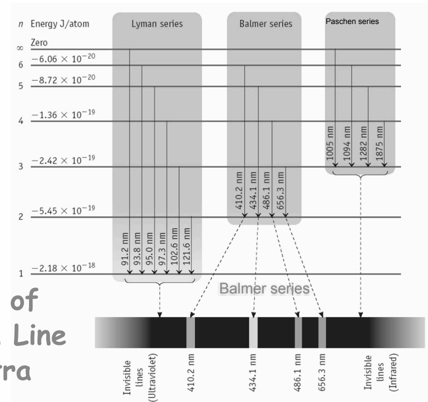

Origin of Emission Line Spectra

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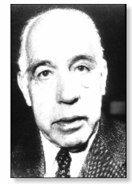
Atomic Line Spectra and Niels Bohr



Niels Bohr
(1885-1962)

Bohr's theory was a great accomplishment.

Received Nobel Prize, 1922

Problems with theory -

- **theory only successful for H & He^{+1}**
- **introduced quantum idea artificially.**

So, we go on to QUANTUM or WAVE MECHANICS

MAR

Quantum or Wave Mechanics

L. de Broglie (1892-1987)

de Broglie (1924) proposed that *all* moving objects have wave properties.

For light: $E = mc^2$

$E = h\nu = hc / \lambda$

Therefore, $mc = h / \lambda$

and for particles

$\lambda = h / mv$

(for particles traveling less than speed of light)

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The de Broglie Wave Equation

Louis de Broglie

***Example:* Calculate the wavelength of an electron traveling at 75.0% the speed of light:**

$\lambda = h / mv$

$$\lambda = \frac{6.626 \times 10^{-34}\ \mathrm{J \cdot s}}{9.11 \times 10^{-31}\ \mathrm{kg} * 2.998 \times 10^{8}\ \frac{m}{s} * 0.750}$$

$\lambda = 3.23 \times 10^{-12}$ m

*Note: **1 J = 1** kg m² s⁻² (derived SI unit)*
Must use kg for mass in these problems

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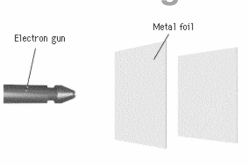
The de Broglie Wave Equation: $\lambda = h / mv$



Experimental proof of wave properties of electrons

electron with velocity = 1.90×10^8 cm/sec

$\lambda = 0.388$ nm

measurable!

Baseball (115 g) at 1000 mph

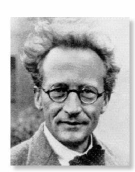
$\lambda = 1.3 \times 10^{-33}$ cm

unmeasurable, but deadly!

Electrons and light both exhibit wave-particle duality!

MAR

Quantum or Wave Mechanics



E. Schrödinger 1887-1961

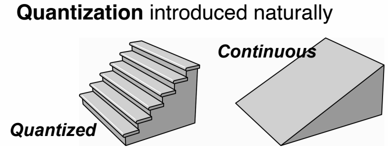
Schrödinger applied idea of e- behaving as a wave to the problem of electrons in atoms.

He developed the WAVE EQUATION

Solution to wave equation gives set of mathematical expressions called WAVE FUNCTIONS, Ψ

Ψ describes the *motion* of electron waves with location and time

Quantization introduced naturally



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WAVE FUNCTIONS, Ψ

Ψ is a function of distance and two angles.

Each Ψ corresponds to an ORBITAL, the region of space within which an electron is found.

Ψ does NOT describe the exact location of the electron.

Ψ^2 is proportional to the probability of finding an e- at a given point.

$$\hat{H}\Psi = i\hbar\frac{\partial}{\partial t}\Psi \qquad \hat{H} = -\frac{\hbar^2}{2m}\nabla^2 + V(\mathbf{r})$$

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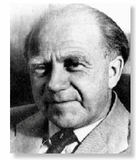
Uncertainty Principle



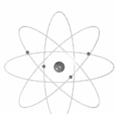
W. Heisenberg
1901-1976

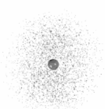
$$\Delta x \cdot m\Delta v \geq \frac{h}{4\pi}$$

- **Problem of defining nature of electrons in atoms explained by Heisenberg.**
- **Cannot simultaneously define the position and momentum (or energy) of an electron.**
- **We define e- energy exactly but accept limitation that we do not know exact position.**

MAR

Implications of Quantum Chemistry





Current model of the atom

- **Modern view of the atom involves *probability* of electron's position (uncertainty principle) while electron's *quantized energy level* known accurately.**
- **Classic physics predicts "planets around the sun" idea, but this is incorrect.**

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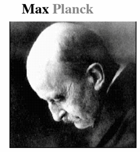
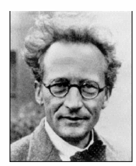
The Giants of Quantum Physics



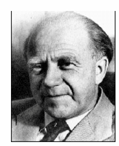
Max Planck

Niels Bohr

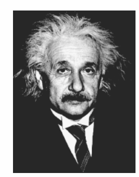
Louis Victor de Broglie



Erwin Schrödinger



Werner Heisenberg



Albert Einstein

MAR

...but great scientists are human!

Heisenberg & Bohr dining in Copenhagen, 1934
(Note the Carlsberg beer!)

They *ALL* took Chem 221 at one time - just as you are now!

'Argue for your limitations, and they are yours'

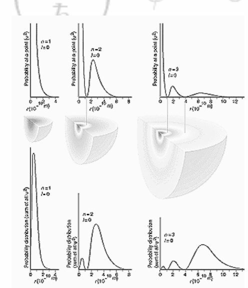
MAR

The Usefulness of Quantum Mechanics

Quantum mechanics involves math equations with calculus

We will use the results of these equations to (eventually) describe how electrons are placed in atoms and ions

First we shall look at orbitals, then shells and subshells



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Subshells & Shells

via quantum mechanics (and calculus!):

Orbitals hold electrons.

Orbitals placed in subshells which are further grouped into shells.

Each shell has a number called the PRINCIPAL QUANTUM NUMBER, n

Principal quantum number helps us to determine the energy and size of an orbital or subshell

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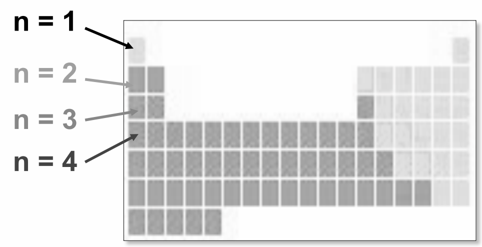
Subshells & Shells



n = Principal Quantum Number = Period Number
(for main group metals and nonmetals)

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QUANTUM NUMBERS

Via quantum mechanics, the shape, size, and energy of each orbital is a function of 4 quantum numbers:

n (principal) --> shell *(energy, size)*

l (angular) --> subshell *(shape)*

m_l (magnetic) --> designates an orbital within a subshell *(direction)*

m_s (spin) --> designates the spin direction of an electron *(Chapter 6 part 2)*

See: Quantum Numbers Handout

MAR

QUANTUM NUMBERS

Symbol	Values	Description
n (principal)	1, 2, 3, ... ∞	Orbital size and energy (*shell*)
l (angular)	0, 1, 2 .. n-1	Orbital shape or type (*subshell*)
m_l (magnetic)	-l... 0... +l	Orbital orientation

# of orbitals in subshell = 2 l + 1

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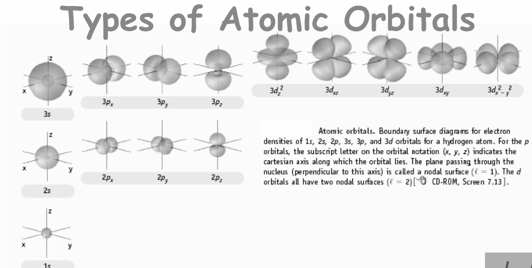
m_s covered in Chapter Six Part 2

Types of Atomic Orbitals



Atomic orbitals. Boundary surface diagrams for electron densities of 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, and 3*d* orbitals for a hydrogen atom. For the *p* orbitals, the subscript letter on the orbital notation (*x*, *y*, *z*) indicates the cartesian axis along which the orbital lies. The plane passing through the nucleus (perpendicular to this axis) is called a nodal surface (ℓ = 1). The *d* orbitals all have two nodal surfaces (ℓ = 2)[CD-ROM, Screen 7.13].

Use *nl* notation: 3s

value of n

value of l:

l	*letter*	
0	s	*stupid*
1	p	*people*
2	d	*drink*
3	f	*freakin'*
4	g	*gas*

MAR

Shells and Subshells

When n = 1, then l = 0 and m_l = 0

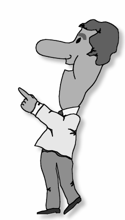
Therefore, in n = 1, there is 1 type of subshell (l)

and that subshell has a single orbital

(m_l has a single value ---> 1 orbital)

This subshell is labeled 1s ("ess")

Each shell has 1 orbital labeled s, and it is SPHERICAL in shape.



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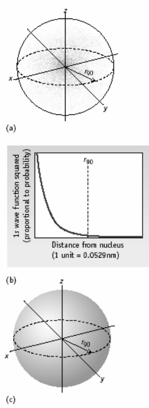
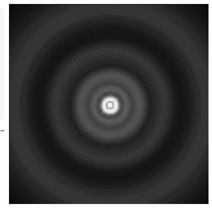
German: sphere = "sphäre", spherical = "sphärisch"

s Orbitals (l = 0)



All *s* orbitals are *spherical* in shape.

s orbitals have *n-l-1 spherical nodes*



a 6s orbital with 6-0-1 = 5 spherical nodes

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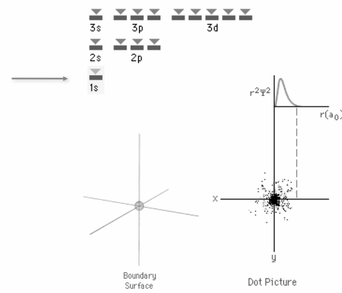
1s Orbital (n=1, l=0)



No spherical nodes in 1s

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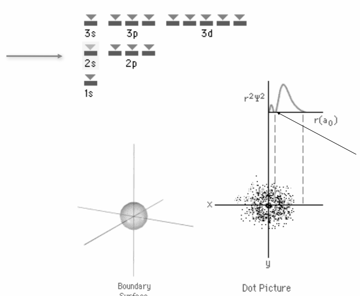
2s Orbital (n=2, l=0)



One spherical node in 2s = n-l-1 in s orbitals

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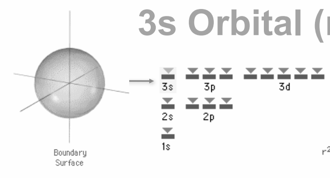
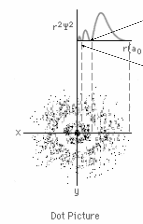
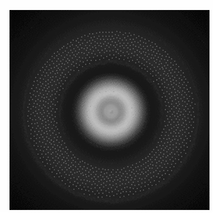
3s Orbital (n=3, l=0)







Two spherical nodes in 3s

Number of spherical nodes = n-l-1 in s orbitals

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p Orbitals (l=1)

When n = 2, then l = 0 and 1

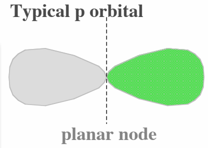
Therefore, in n = 2 shell there are *2 types of orbitals* or 2 subshells

For l = 0 m_l = 0

this is a 2s subshell

For l = 1 m_l = -1, 0, +1

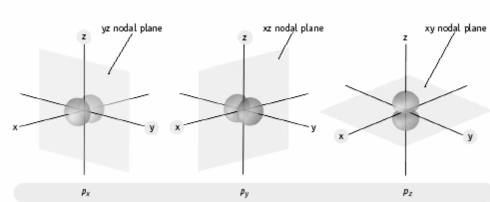
this is a 2p subshell with 3 orbitals



When l = 1, there is a PLANAR NODE through the nucleus.

MAR

p Orbitals (l=1)

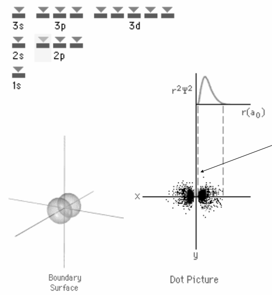


The three p orbitals lie 90° apart in space

Each p orbital has 1 planar node

MAR

$2p_x$ Orbital (n=2, l=1)

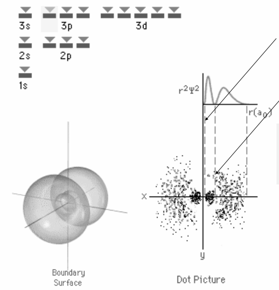


One planar node in $2p_x$

$2p_y$ and $2p_z$ similar

MAR

$3p_x$ Orbital (n=3, l=1)



One planar node in $3p_x$

One spherical node in $3p_x$

 # spherical nodes = n - l - 1

$3p_x$ *(n = 3)* larger than $2p_x$ *(n = 2)*

MAR

d Orbitals (l=2)

When n = 3, what are the values of l?

l = 0, 1, 2

and so there are 3 subshells in the shell.

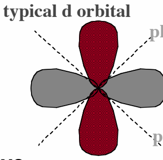
For l = 0, m_l = 0
---> 3s subshell with single orbital

For l = 1, m_l = -1, 0, +1
---> 3p subshell with 3 orbitals

For l = 2, m_l = -2, -1, 0, +1, +2
---> 3d subshell with 5 orbitals

MAR

d Orbitals (l=2)



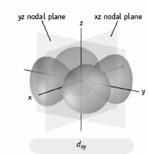
s orbitals have no planar node (l = 0) and so are spherical.

p orbitals have l = 1 and have 1 planar node ("dumbbell" shaped)

This means d orbitals (with l = 2) have 2 planar nodes

 # planar nodes = l

 # spherical nodes = n - l - 1



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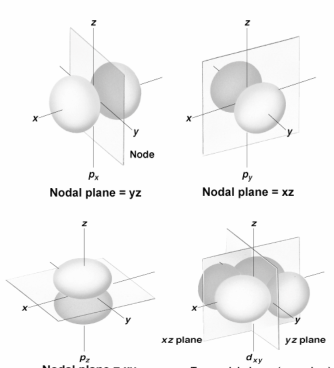
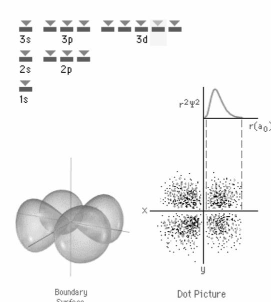


Nodal Planes in p and d Orbitals

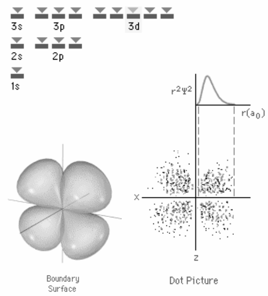
MAR

$3d_{xy}$ Orbital (n=3, l=2)



MAR

$3d_{xz}$ Orbital (n=3, l=2)



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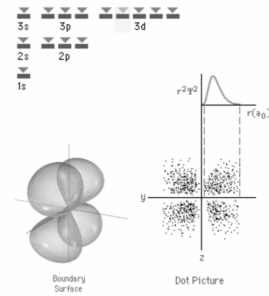
$3d_{yz}$ Orbital (n=3, l=2)



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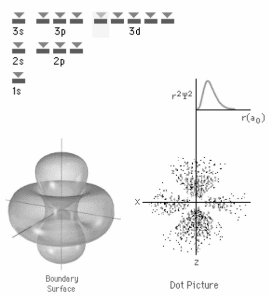
$3d_{z^2}$ Orbital (n=3, l=2)



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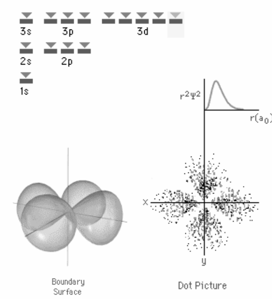
$3d_{x^2-y^2}$ Orbital (n=3, l=2)



MAR

The n = 3 shell *(example)*

l **values**			
subshell letter			
nl **notation**			
m_l **values**			
# orbitals			

***three subshells** when n=3*

***l*: 0, 1, ... (n-1)**

m_l : -l...0...+l

subshell:

	s	p	d	f
l:	0	1	2	3

 # orbitals =

***1 + 3 + 5* = 9**

Total orbitals = n^2

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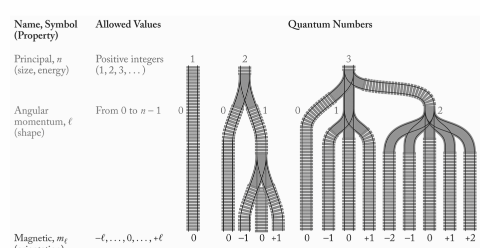
The Train Track Model



MAR

f Orbitals (l = 3)

When n = 4, l = 0, 1, 2, 3 so there are 4 subshells in the shell.

For l = 0, m_l = 0

---> 4s *subshell with single orbital*

For l = 1, m_l = -1, 0, +1

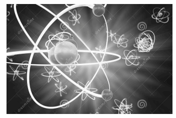
---> 4p *subshell with 3 orbitals*

For l = 2, m_l = -2, -1, 0, +1, +2

---> 4d *subshell with 5 orbitals*

For l = 3, m_l = -3, -2, -1, 0, +1, +2, +3

---> 4f subshell with 7 orbitals



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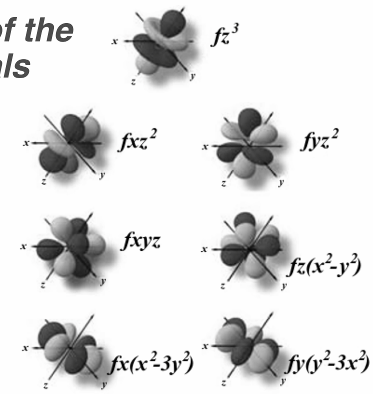
Pictures of the f Orbitals



MAR

Quick & Dirty Quantum Chemistry

memorize **c = 2.998 x 10^8 m/s**

memorize **h = 6.626 x 10^{-34} J s**

E = hν = hc/λ *ν = frequency (Hz)*

λ = h / mv *v = velocity (m/s)*

 # orbitals in a shell = **n^2**

 # orbitals in a subshell = **2l + 1**

stupid people drive freakin' gas hogs

0 1 2 3 4 5 *(l values)*

so a **4d** *subshell would have* **n = 4, l = 2**

 # planar nodes = **l**

 # spherical nodes = **n - l - 1**

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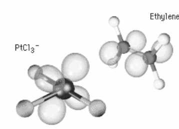
Importance of Orbitals

Knowledge of orbitals critical when understanding bonding in molecules (we'll see this in CH 222)

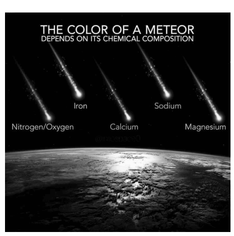


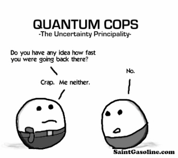
MAR

End of Chapter Six Part 1

See also:

- *Chapter Six Part 1 Study Guide*
- *Chapter Six Part 1 Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*





"More decisive? How can I be more decisive? - I live by the uncertainty principle!"

MAR

Important Equations, Constants, and Handouts from this Chapter:

- *know relationship between* **frequency, wavelength, energy, speed of light, energy per mole**
- *know the* **regions** *and* **relative energies** *within the* **electromagnetic spectrum**
- *know about* **sharp line spectra, absorbance** *and* **emission spectra**
- *know about* **wave particle duality** (*including* **de Broglie**)
- **quantum numbers:** *know the origin and meaning of* **n, l, m_l**
- *know* **"nl" notation** *(4s, 3d, etc.)*
- *know how to find* **spherical** *and* **planar nodes, number of orbitals,** *etc.*

c = 2.998 x 10^8 m/s

h = 6.626 x 10^{-34} J s

E = hν = hc/λ *(E/M)*

λ = h / mv *(particles)*

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End of Chapter Problems: *Test Yourself*

1. Place the following types of radiation in order of increasing energy per photon: **yellow light, x-rays, microwaves** and your favorite FM music **radio** station at 92.3 MHz.
2. Aluminum has an emission line at 396.15 nm. What is the frequency of this line? What is the energy of one photon with this wavelength? Of 1.00 mol of these photons?
3. A rifle bullet (mass = 1.50 g) has a velocity of 7.00 x 10^2 miles per hour. What is the wavelength associated with this bullet? (0.6214 miles = 1 km)
4. a. When n = 4, what are the possible values of ℓ?
   b. When ℓ is 2, what are the possible values of m_ℓ?
   c. For a 4*s* orbital, what are the values of n and ℓ?
5. Explain why these sets of quantum numbers are incorrect.
   a. n = 3, l = 3, m_l = 0, ms = +1/2
   b. n= 4, l = 3, m_l = -4, ms = -1/2
6. How many nodal surfaces (planar and spherical) are associated with each of the following atomic orbitals? **5f** *and* **4s**

MAR

End of Chapter Problems: *Answers*

1. radio, microwave, yellow light, x-rays
2. υ = 7.568 x 10^{14} Hz, E = 5.014 x 10^{14} J/ph, E = 3.02 x 10^5 J/mol
3. 1.41 x 10^{-33} m
4. a. 0, 1, 2 or 3. b. 0, ±1, ±2. c. n = 4, l = 0.
5. a. l cannot equal n. b. m_l can only equal ±l (+3 to -3 only)
6. **5f**: three planar and one spherical node. **4s**: zero planar and three spherical nodes.

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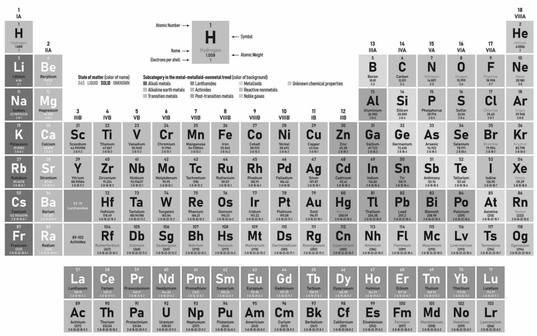
The Structure of Atoms and Periodic Trends
Chapter Six Part 2

Periodic Table of the Elements



CH 221
Professor Michael Russell

MAR *Last update: 4/29/24*

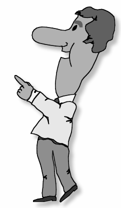
Arrangement of Electrons in Atoms

Electrons in atoms are arranged as

SHELLS (n)

SUBSHELLS (l)

ORBITALS (m_l)



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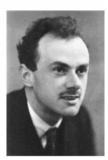
Arrangement of Electrons in Atoms

Each orbital can be assigned no more than 2 electrons!

This is tied to the existence of a 4th quantum number, the electron spin quantum number, m_s.

m_s arises naturally when relativity (Einstein) combined with quantum mechanics (Paul Dirac)



Paul Dirac

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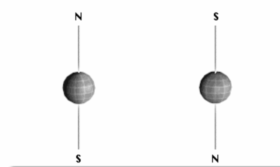

Electron Spin Quantum Number, m_s

Electron spin can be proven experimentally.
Two spin directions are given by m_s where m_s = $+^1/_2$ and $-^1/_2$.
Leads to magnetism in atoms and ions

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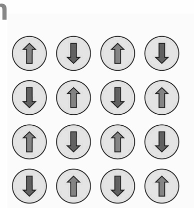
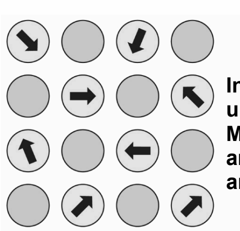
Magnetism

In diamagnetic systems, all electron spins are paired - no net magnetic moment.





In paramagnetic systems, unpaired spins present. Magnetic fields randomly arranged unless placed in an external magnetic field.

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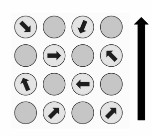
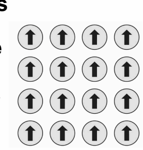
Magnetism

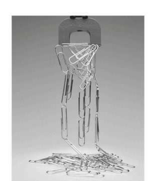
In ferromagnetic substances the orientations of magnetic fields from unpaired electrons are affected by spins from electrons around them.



H_0 - a magnetic field

When an external field is applied *and then removed*, the substance maintains the magnetic moment and becomes a permanent magnet.





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Electron Spin Quantum Number

Diamagnetic: NOT attracted to a magnetic field; spin paired.
Paramagnetic: substance is attracted to a magnetic field. Substance has unpaired electrons.

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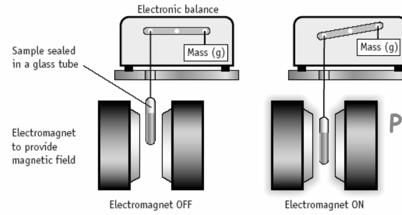
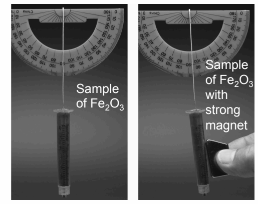


Measuring Paramagnetism

Paramagnetic: substance is attracted to a magnetic field. Substance has unpaired electrons.

Diamagnetic: NOT attracted to a magnetic field



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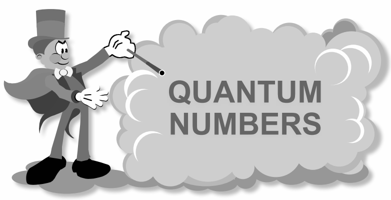

n	---> **shell**	**1, 2, 3, 4, ...**
l	---> **subshell**	**0, 1, 2, ... (n - 1)**
m_l	---> **orbital**	**-l ... 0 ... +l**
m_s	---> **electron spin**	**$+^1/_2$ and $-^1/_2$**

See: Quantum Numbers Handout

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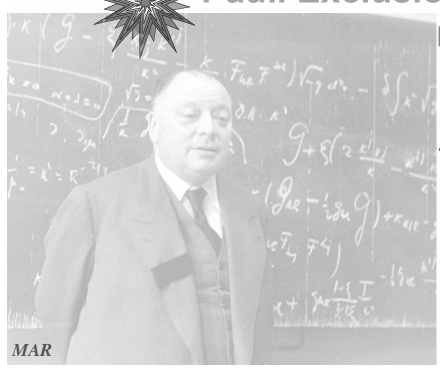
Pauli Exclusion Principle



No two electrons in the same atom can have the same set of 4 quantum numbers.

That is, each electron has a unique address which will consist of its own values of n, l, m_l and m_s.

MAR

Wolfgang Pauli

Electrons in Atoms - the Pauli Exclusion Principle

*When n = 1, then **l = 0** and **m_l = 0***
*this shell has a **single orbital (1s)** to which*
2e- can be assigned

n = 1, l = 0, m_l = 0, m_s = $+^1/_2$ - this is electron #1
n = 1, l = 0, m_l = 0, m_s = $-^1/_2$ - this is electron #2

*When n = 2, then **l = 0** (s), **1** (p)*

2s orbital	**2e-**
three 2p orbitals	**6e-**
TOTAL =	**8e-**

"No two electrons in the same atom can have the same set of 4 quantum numbers."

electron number	n	l	m_l	m_s	
1	2	0	0	½	2s
2	2	0	0	-½	2s
3	2	1	-1	½	2p
4	2	1	-1	-½	2p
5	2	1	0	½	2p
6	2	1	0	-½	2p
7	2	1	+1	½	2p
8	2	1	+1	-½	2p

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Electrons in Atoms

*When n = 3, then **l = 0** (s), **1** (p), **2** (d)*

3s orbital	**2e-**
three 3p orbitals	**6e-**
five 3d orbitals	**10e-**
TOTAL =	**18e-**

Each electron has its own set of four quantum numbers!

Wolfgang Pauli

electron number	n	l	m_l	m_s	
1	3	0	0	½	3s
2	3	0	0	-½	3s
3	3	1	-1	½	3p
4	3	1	-1	-½	3p
5	3	1	0	½	3p
6	3	1	0	-½	3p
7	3	1	+1	½	3p
8	3	1	+1	-½	3p
9	3	2	-2	½	3d
10	3	2	-2	-½	3d
11	3	2	-1	½	3d
12	3	2	-1	-½	3d
13	3	2	0	½	3d
14	3	2	0	-½	3d
15	3	2	+1	½	3d
16	3	2	+1	-½	3d
17	3	2	+2	½	3d
18	3	2	+2	-½	3d

MAR

Electrons in Atoms

electron number

	n	l	m_l	m_s	
1	4	0	0	1/2	4s
2	4	0	0	-1/2	4s
3	4	1	-1	1/2	4p
4	4	1	-1	-1/2	4p
5	4	1	0	1/2	4p
6	4	1	0	-1/2	4p
7	4	1	+1	1/2	4p
8	4	1	+1	-1/2	4p
9	4	2	-2	1/2	4d
10	4	2	-2	-1/2	4d
11	4	2	-1	1/2	4d
12	4	2	-1	-1/2	4d
13	4	2	0	1/2	4d
14	4	2	0	-1/2	4d
15	4	2	+1	1/2	4d
16	4	2	+1	-1/2	4d
17	4	2	+2	1/2	4d
18	4	2	+2	-1/2	4d

When n = 4, **l = 0**(*s*), **1**(*p*), **2**(*d*), **3**(*f*)

4s orbital	**2e-**
three 4p orbitals	**6e-**
five 4d orbitals	**10e-**
seven 4f orbitals	**14e-**
TOTAL =	**32e-**

electron number

	n	l	m_l	m_s	
19	4	3	-3	1/2	4f
20	4	3	-3	-1/2	4f
21	4	3	-2	1/2	4f
22	4	3	-2	-1/2	4f
23	4	3	-1	1/2	4f
24	4	3	-1	-1/2	4f
25	4	3	0	1/2	4f
26	4	3	0	-1/2	4f
27	4	3	+1	1/2	4f
28	4	3	+1	-1/2	4f
29	4	3	+2	1/2	4f
30	4	3	+2	-1/2	4f
31	4	3	+3	1/2	4f
32	4	3	+3	-1/2	4f

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Electron Shell (n)	Subshells Available	Orbitals Available ($2\ell + 1$)	Number of Electrons Possible in Subshell [$2(2\ell + 1)$]	Maximum Electrons Possible for nth Shell ($2n^2$)
1	*s*	1	2	2
2	*s*	1	2	8
	p	3	6	
3	*s*	1	2	18
	p	3	6	
	d	5	10	
4	*s*	1	2	32
	p	3	6	
	d	5	10	
	f	7	14	
5	*s*	1	2	50
	p	3	6	
	d	5	10	
	f	7	14	
	*g**	9	18	
6	*s*	1	2	72
	p	3	6	
	d	5	10	
	*f**	7	14	
	*g**	9	18	
	*h**	11	22	

*These orbitals are not used in the ground state of any known element.

Distribution of Electrons in Shells

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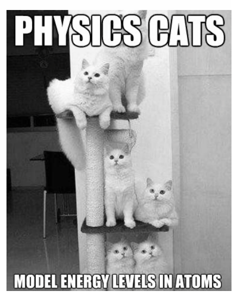
Assigning Electrons to Atoms

Electrons generally assigned to orbitals of successively higher energy.

<u>For H atoms</u>, E = - Rhc($1/n^2$). E depends only on n.

<u>For many-electron atoms</u>, energy depends on both n and l... introducing the "n + l" rule



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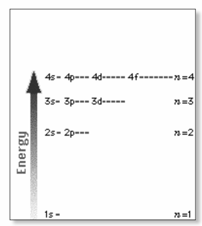
Assigning Electrons to Subshells



In H atom all subshells of same n have same energy.

In many-electron atom:

a) subshells increase in energy as value of n + l increases. (The *important* n + l rule)

b) for subshells of same n + l, subshell with lower n is lower in energy.

See <u>Electron Configurations Handout</u>

Using the n + l rule assumes zero point energy - the lowest energy state possible, or ground state

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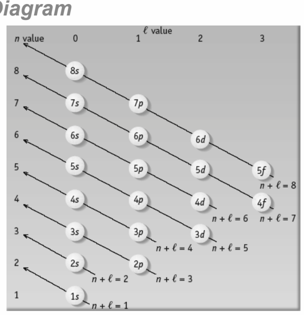
An *Aufbau* Diagram



Electron Filling Order

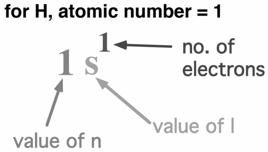
Aufbau comes from a German word meaning "building up", formulated by Bohr and Pauli in 1920s

MAR

Writing Atomic Electron Configurations

Two ways of writing configs. One is called the spectroscopic or "spdf" notation.

SPECTROSCOPIC NOTATION for H, atomic number = 1



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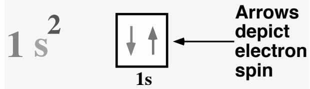
Writing Atomic Electron Configurations

Two ways of writing configs. Other is called the orbital box notation.

ORBITAL BOX NOTATION for He, atomic number = 2



One electron has n = 1, l = 0, m_l = 0, m_s = + 1/2

Other electron has n = 1, l = 0, m_l = 0, m_s = - 1/2

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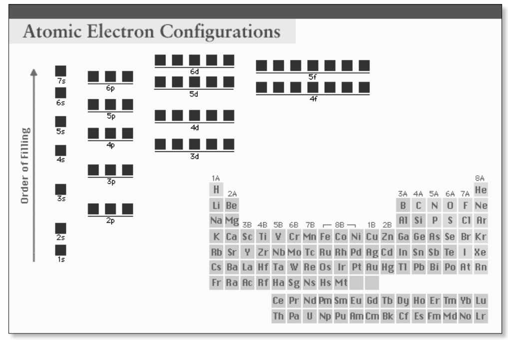

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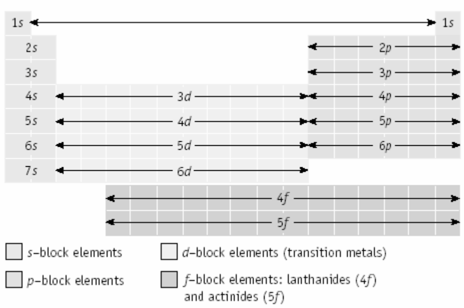
Atomic Electron Configurations Diagram

Electron Configurations and the Periodic Table



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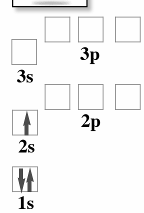
Lithium

Group 1A

Atomic number = 3

$1s^2 2s^1$ ---> 3 total electrons

paramagnetic



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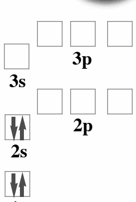
Beryllium

Group 2A

Atomic number = 4

$1s^2 2s^2$ ---> 4 total electrons

diamagnetic



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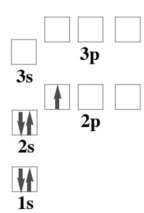
Boron

Group 3A

Atomic number = 5

$1s^2 2s^2 2p^1$ ---> 5 total electrons

paramagnetic



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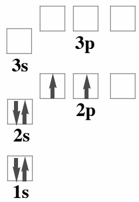
Carbon

Group 4A
Atomic number = 6
$1s^2 2s^2 2p^2$ --->
6 total electrons
paramagnetic



Here we see for the first time HUND'S RULE. When placing electrons in a set of orbitals having the same energy, we place them singly as long as possible.

Friedrich Hund

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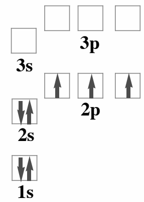
Nitrogen

Group 5A
Atomic number = 7
$1s^2 2s^2 2p^3$ --->
7 total electrons
paramagnetic



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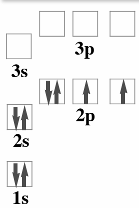
Oxygen

Group 6A
Atomic number = 8
$1s^2 2s^2 2p^4$ --->
8 total electrons
paramagnetic



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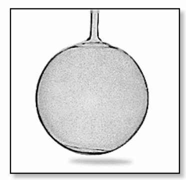
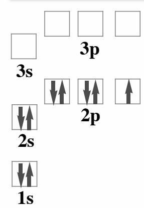


Fluorine

Group 7A
Atomic number = 9
$1s^2 2s^2 2p^5$ --->
9 total electrons
paramagnetic



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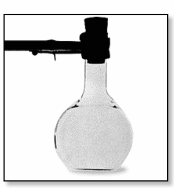
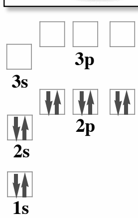


Neon

Group 8A
Atomic number = 10
$1s^2 2s^2 2p^6$ --->
10 total electrons
diamagnetic



Note that we have reached the end of the 2nd period, and the 2nd shell is full!

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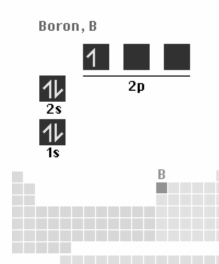
Electron Configurations of p-Block Elements



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Sodium

Group 1A
Atomic number = 11
$1s^2\,2s^2\,2p^6\,3s^1$ or
"neon core" + $3s^1$
[Ne] $3s^1$ *(uses noble gas notation)*
And: we have begun a new period!

All Group 1A elements have [core]ns^1 configurations.

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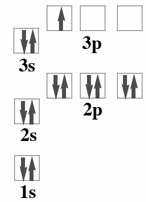
Aluminum

Group 3A
Atomic number = 13
$1s^2\,2s^2\,2p^6\,3s^2\,3p^1$ *or*
[Ne] $3s^2\,3p^1$

All* Group 3A elements have [core] $ns^2\,np^1$ configurations where n is the period number.

* *some have $(n-1)d^{10}$ as well*



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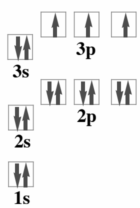
Phosphorus

Group 5A
Atomic number = 15
$1s^2\,2s^2\,2p^6\,3s^2\,3p^3$ *or*
[Ne] $3s^2\,3p^3$

All* Group 5A elements have [core] $ns^2\,np^3$ configurations where n is the period number.

* *some have $(n-1)d^{10}$ also*



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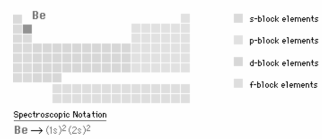
Relationship of Electron Configuration and Region of the Periodic Table



Gray = s block
Orange = p block
Green = d block
Violet = f block

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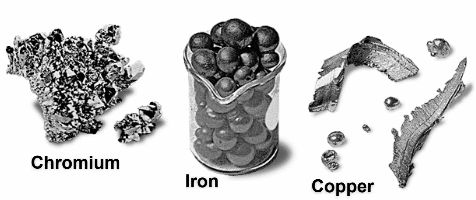
Transition Metals

All transition metals have the configuration [core]$ns^x(n-1)d^y$ and so are "d-block" elements.



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Fourth Period Electron Configurations

K	$1s^22s^22p^63s^23p^64s^1$	or [Ar]$4s^1$
Ca	$1s^22s^22p^63s^23p^64s^2$	or [Ar]$4s^2$
Sc	$1s^22s^22p^63s^23p^64s^23d^1$	or [Ar]$4s^23d^1$
Ti	$1s^22s^22p^63s^23p^64s^23d^2$	or [Ar]$4s^23d^2$
V	$1s^22s^22p^63s^23p^64s^23d^3$	or [Ar]$4s^23d^3$
Cr	$1s^22s^22p^63s^23p^64s^13d^5$	or [Ar]$4s^13d^5$
Mn	$1s^22s^22p^63s^23p^64s^23d^5$	or [Ar]$4s^23d^5$
Cu	$1s^22s^22p^63s^23p^64s^13d^{10}$	or [Ar]$4s^13d^{10}$

Note: *exceptions!*

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Electron Configuration Anomalies

Cr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ or $[Ar]4s^1 3d^5$

Cu $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$ or $[Ar]4s^1 3d^{10}$

Chromium, copper and other elements do not follow the *n* + *l* filling orders

Anomalies arise from stability associated with half-filled and completely filled d-subshells.

Know how n + l rule works, and know that anomalies exist on the periodic table

MAR

Lanthanides and Actinides

All these elements have the configuration
$[core]ns^x(n - 1)d^y(n - 2)f^z$
and so are "f-block" elements

Exceptions exist:

Cerium
[Xe] $6s^2$ $5d^1$ $4f^1$
[Xe] $6s^2$ $4f^2$ expected

Uranium
[Rn] $7s^2$ $6d^1$ $5f^3$
[Rn] $7s^2$ $5f^4$ expected

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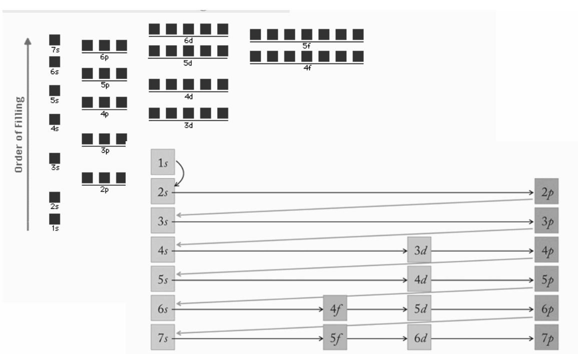

Electron Configurations Filling Order

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Transition Metals

Iron:
Zinc:
Technetium:
Niobium:
Osmium:
Meitnerium:

notice f orbitals in 6th period & beyond

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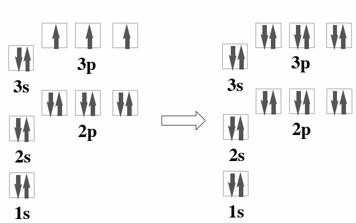
Anion Configurations

To form *anions* from elements add 1 or more e- using normal n + l rules

P [Ne] $3s^2$ $3p^3$ + 3e- → P^{3-} [Ne] $3s^2$ $3p^6$



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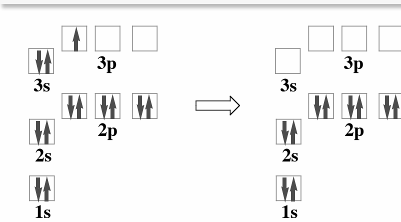
Cation Configurations

To form cations from elements remove 1 or more e- from subshell of highest n [or highest (n + l)].

Al [Ne] $3s^2$ $3p^1$ - 3e- → Al^{3+} [Ne]



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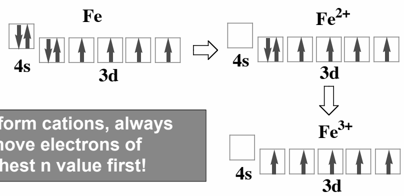
Ion Configurations

For transition metals, remove ns electrons and then (n - 1) electrons.

Fe [Ar] $4s^2$ $3d^6$ ---> Fe^{2+} [Ar] $3d^6$ ---> Fe^{3+} [Ar] $3d^5$



To form cations, always remove electrons of highest n value first!

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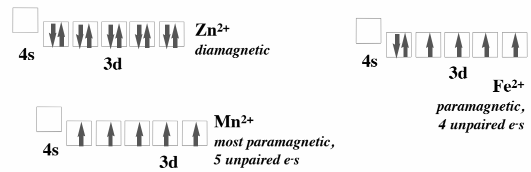
Magnetic Properties

Magnetic properties of ions assist us with charges

DIAMAGNETIC ions have no unpaired electrons. Ions with unpaired electrons are PARAMAGNETIC.

As number of unpaired electrons increases, the degree of paramagnetism also increases



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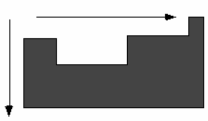
Periodic Trends

CH 221 Q&D Guide to Periodic Trends:

- **Atomic and ionic size: increase left *and* down**
- **Ionization energy *and* Electron affinity: increase right and up**
- *See Periodic Trends Handout*

Electrons held more tightly

Larger orbitals. Electrons held less tightly.



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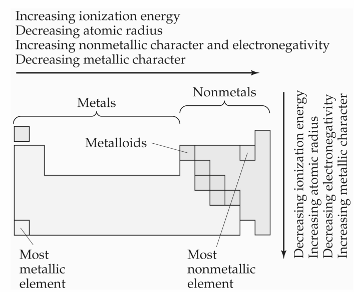
CH 221 Periodic Trends "Cheat Sheet"



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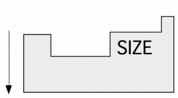
Atomic Size



Size increases as you go down a group.

Because electrons are added further from the nucleus, there is less attraction.

Size increases as you go left across a period.

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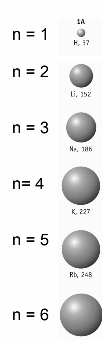
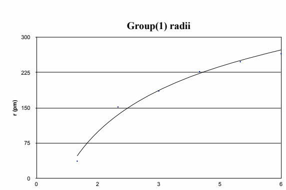


Moving down group IA, the atomic radii increase with the principle quantum number



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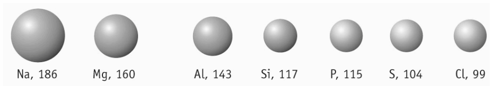
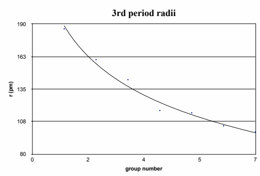


Moving **across** the 3rd period we see the atomic radii of the elements **decrease**.

Atomic **radii** *generally* **increase** going **right to left** on the periodic table.



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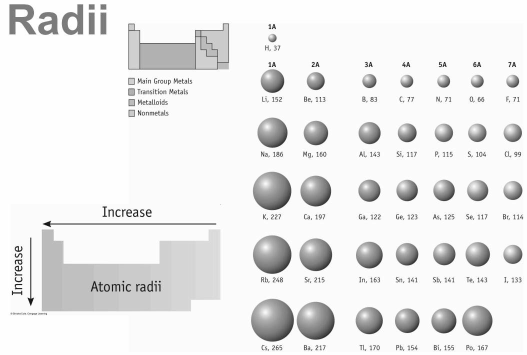
Atomic Radii



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Ion Sizes

Li,152 pm
3e and 3p

Does the size go up or down when losing an electron to form a cation?

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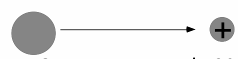
Ion Sizes



Li,152 pm
3e and 3p

Li^+, 60 pm
2e and 3 p

Forming a cation.

- **CATIONS are SMALLER than the atoms from which they come.**
- **The electron/proton attraction has increased, and so size DECREASES.**

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Ion Sizes

F,64 pm
9e and 9p

Does the size go up or down when gaining an electron to form an anion?

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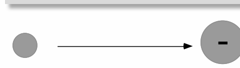
Ion Sizes



F,64 pm
9e and 9p

F^-, 136 pm
10 e and 9 p

Forming an anion.

- **ANIONS are LARGER than the atoms from which they come.**
- **The electron/proton attraction has decreased, and so size INCREASES.**
- **Trends in ion sizes are the same as atom sizes** *(but only compare cations to cations or anions to anions!)*

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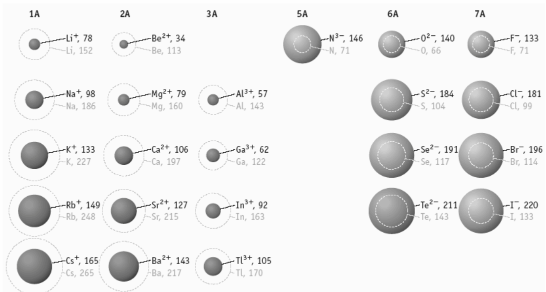
Trends in Ion Sizes



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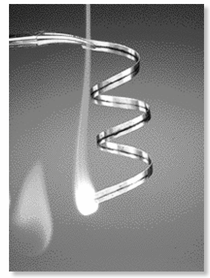
Redox Reactions



Why do metals lose electrons in their reactions?

Why does Mg form Mg^{2+} ions and not Mg^{3+}?

Why do nonmetals take on electrons?

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Ionization Energy

$Mg (g) + 738 kJ$ ---> $Mg^+ (g) + e-$

$Mg^+ (g) + 1451 kJ$ ---> $Mg^{2+} (g) + e-$

$Mg^{2+} (g) + 7733 kJ$ ---> $Mg^{3+} (g) + e-$

Mg^+ has 12 protons and only 11 electrons. Therefore, IE for $Mg^+ > Mg$.

IE = energy required to remove an electron from an atom in the gas phase.

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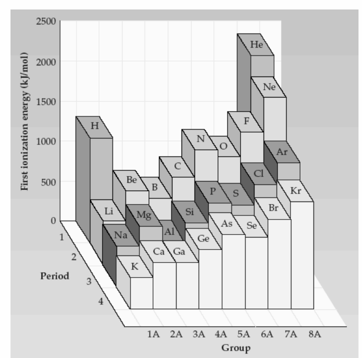
Trends in Ionization Energy



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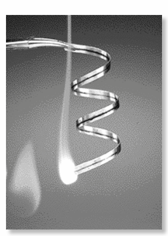
Trends in Ionization Energy

Ionization Energy increases moving right across a period and up a group on the periodic table



Metals lose electrons more easily than nonmetals.

Metals are good reducing agents.

Nonmetals lose electrons with difficulty.

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Periodic Trend in the Reactivity of Alkali Metals with Water

Lithium

Sodium

Potassium

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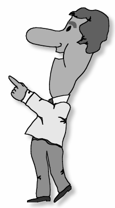
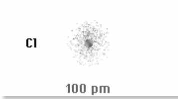
Electron Affinity

Nonmetlas tend to GAIN electrons to form anions.

Electron affinity is the energy involved when an atom gains an electron.

A(g) + e- ---> A^-(g) E.A. = ΔE





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Trends in Electron Affinity

Electron Affinity increases as you move right across a period (EA becomes more negative).

Electron Affinity increases as you move up a group (EA becomes more negative).

1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
H −72.6							He (0.0)*
Li −59.6	Be >0	B −26.7	C −122	N +7	O −141	F −328	Ne (+29)*
Na −52.9	Mg >0	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+35)*
K −48.4	Ca −2.4	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)*
Rb −46.9	Sr −5.0	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+41)*
Cs −45.5	Ba −14	Tl −19.2	Pb −35.2	Bi −91.3	Po −183.3	At −270*	Rn (+41)*

*Calculated values.

Electron Affinity values (kJ/mol)

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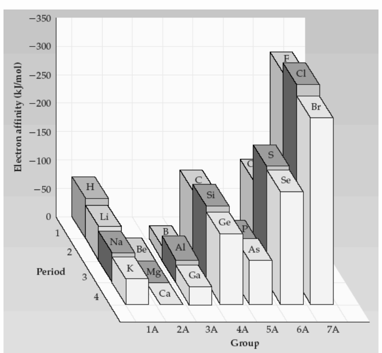
Trends in Electron Affinity



Notice:
$EA_{(F)} < EA_{(Cl)}$
unknown mechanism, electron repulsion? atom size?

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Implications of Periodic Trends

Useful in predicting reactivities, chemical formulas, etc.

Metals: **low ionization energy, give up electrons easily**
Nonmetals: **high electron affinity, love electrons from metals**

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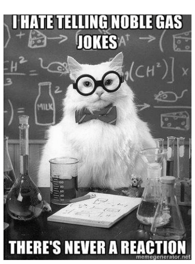
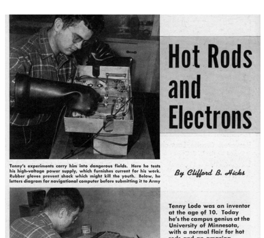


End of Chapter 6 Part 2



What Quantum Physics does to a man

See also:
- *Chapter Six Part 2 Study Guide*
- *Chapter Six Part 2 Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*

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Important Equations, Constants, and Handouts from this Chapter:

- quantum numbers: *know the origin and meaning of* **n, l, m_l, m_s**
- *understand* **paramagnetism** *and* **diamagnetism** *for atoms and ions*
- *know* "nl" notation *(4s, 3d, etc.) and the* "n + l" rule *for energy*
- *know how the* **Pauli Exclusion Theory** *and* **Hund's Rule** *apply towards electrons in orbitals; know the* **Aufbau Principle**
- *know how to* **create electron configurations** *for* **neutral atoms** *and also* **cations** *and* **anions using both orbital box** *and* **spectroscopic notation**
- *know the* **periodic trends** *for* **size**, **ion size**, **ionization energy** *and* **electron affinity**

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End of Chapter Problems: ***Test Yourself***

1. Depict the electron configuration for arsenic (As) using *spdf* notation.
2. Using orbital box diagrams and/or noble gas notation, depict the electron configurations of the following: (a) V, (b) V^{2+}, and (c) V^{5+}. Are any of the ions paramagnetic? How many unpaired electrons are in each species?
3. Arrange the following elements in order of increasing size: Al, B, C, K, and Na.
4. Name the element corresponding to each characteristic below.
   a. the element with the electron configuration $1s^22s^22p^63s^23p^3$
   b. the alkaline earth element with the smallest atomic radius
   c. the element with the largest ionization energy in Group 5A
   d. the element whose 2+ ion has the configuration $[Kr]4d^5$
   e. the element with the most negative electron affinity in Group 6A
   f. the element whose electron configuration is $[Ar]3d^{10}4s^2$

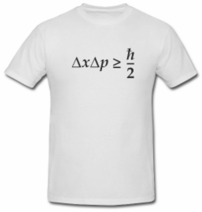
MAR

End of Chapter Problems: ***Answers***

1. $[Ar]3d^{10}4s^24p^3$ ***or*** $1s^22s^22p^63s^23p^63d^{10}4s^24p^3$
2. V: $[Ar]4s^23d^3$ (paramagnetic, 3 unpaired electrons); V^{2+}: $[Ar]3d^3$ (paramagnetic, 3 unpaired electrons); $V5^+$: [Ar] (diamagnetic, 0 unpaired electrons);
3. C < B < Al < Na < K
4. a. P b. Be c. N d. Tc e. O f. Zn

MAR

Chemistry 221 Final Exam Review
Chapter 6



Chemistry 221
Professor Michael Russell

Last update: 4/29/24

MAR

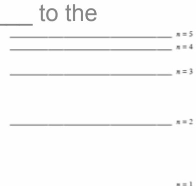
How many emission lines are possible considering only the five quantum levels of hydrogen shown below?

A. 3
B. 4
C. 7
D. 8
E. 10

MAR

Photons of the highest frequency will be emitted in a transition from the level with n = ____ to the level with the n = _____.

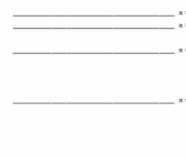
A. from n = 1 to n = 2
B. from n = 2 to n = 1
C. from n = 3 to n = 1
D. from n = 4 to n = 1
E. from n = 5 to n = 1



MAR

The emission line having the longest wavelength corresponds to a transition from the level with n = ___ to the level with n = ___.

A. from n = 1 to n = 2
B. from n = 2 to n = 1
C. from n = 4 to n = 1
D. from n = 5 to n = 1
E. from n = 5 to n = 4



MAR

Excited H atoms emit *visible light* when electrons fall from higher levels to n = 2 (this is called the Balmer series of lines). If green light comes from the transition from n = 4 to n = 2, is the light from the n = 3 to n = 2 transition expected to be red or blue?

A. Red
B. Blue
C. ultramagnetic

MAR

Calculate the wavelength in nanometers associated with an energy change of 182.3 kJ/mol.

A. 3.027×10^{-19}
B. 6.563×10^{-7}
C. 302.7
D. 656.3
E. *billions!*

MAR

What is the observed wavelength for an electron (mass = 9.109×10^{-28} g) traveling at a speed of 1.20×10^{8} m/s? *(hint: use **kg** for mass!)*

A. 6.06×10^{-3} m
B. 1.17×10^{-5} fm
C. 3.00×10^{8} m
D. 6.06×10^{-3} nm
E. none of the above

MAR

Which of the following is NOT a valid set of quantum numbers?

A. n = 4, ℓ = 1, and m_ℓ = -1
B. n = 6, ℓ = 5, and m_ℓ = 0
C. n = 2, ℓ = 2, and m_ℓ = +1
D. n = 3, ℓ = 2, and m_ℓ = -2
E. n = 1, ℓ = 0, and m_ℓ = 0

MAR

For a certain orbital, n = 3, ℓ = 1, and m_ℓ = -1. What type of orbital is this?

A. 3d
B. 3s
C. 3p
D. 4d
E. 1f

MAR

If an electron subshell has 7 orbitals, what is the ℓ value for this subshell?

A. two
B. three
C. four
D. five
E. seven

MAR

What type of orbital has 2 nodal planes?

A. *s*
B. *p*
C. *d*
D. *f*
E. *g*

MAR

Which of the following orbitals has 2 spherical nodes?

A. 1*s*
B. 2*p*
C. 3*d*
D. 3*p*
E. 4*p*

MAR

The electron configuration for neutral chlorine is

A. $1s^2\ 2s^2 2p^6\ 3s^5$
B. $1s^2\ 2s^2 2p^6\ 3s^2 3p^5$
C. $1s^2\ 2s^2 2p^5$
D. $1s^2\ 2s^2 2p^6\ 3s^2 3p^6$
E. [Xe]

MAR

What neutral element has the electron configuration $1s^2\ 2s^2 2p^6\ 3s^2 3p^6\ 3d^{10}\ 4s^2$?

A. Zn
B. Ca
C. Ge
D. Ni
E. H

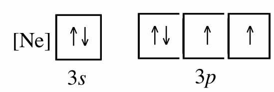
MAR

The electron configuration for neutral tin is:

A. [Ne] $4s^2\ 3d^{10}\ 4p^2$
B. [Ar] $4s^2\ 3d^{10}\ 4p^2$
C. [Kr] $5s^2\ 4d^{10}\ 5p^2$
D. [Xe] $5s^2\ 4d^{10}\ 5p^2$
E. [Uuo] or [Og]
Z = 118! :)

MAR

What neutral element has the following electron configuration?



A. Cl
B. S
C. Be
D. Ti
E. Jq

MAR

What neutral element has the electron configuration [Xe] $4f^{14}\ 5d^{10}\ 6s^2\ 6p^2$?

A. Hf
B. Lu
C. Pb
D. Sn
E. Jq

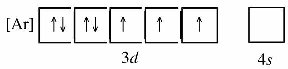
MAR

What is the electronic configuration of P^{3-}?

A. [Ne] $3s^2\ 3p^6$
B. [Ne] $3s^2\ 3p^3$
C. [Ne] $3s^2$
D. [Ne] $3p^6$
E. [Ne]

MAR

What ion corresponds to the following electron configuration?



A. Fe^{3+}
B. Rh^{3+}
C. Co^{2+}
D. Ni^{2+}
E. Li^{1+}

MAR

Which of the following ions is diamagnetic?

A. Ti^{2+}
B. V^{2+}
C. Mg^{2+}
D. Cr^{2+}
E. none are diamagnetic

MAR

Which of the following is the correct electronic configuration for the nickel(II) ion?

A. [Ar] $3d^8$
B. [Kr] $4s^2$ $3d^6$
C. [Ar] $4s^2$ $3d^6$
D. [Kr] $3d^8$
E. [He]

MAR

Which of the following is the correct electronic configuration for the tin(II) ion?

A. [Kr] $5s^2$ $5p^2$ $4d^{10}$
B. [Kr] $5s^2$ $4d^{10}$
C. [Kr] $5s^2$ $5p^2$ $4d^8$
D. [Kr] $5p^2$ $4d^{10}$
E. [He] $2s^2$ $2p^2$

MAR

Compare the elements Na, B, Al, and C with regard to the following properties:
Which has the largest atomic radius?

A. Na
B. B
C. Al
D. C
E. Jq

MAR

Which of the following is expected to have the largest radius?

A. P^{3-}
B. Cl^{-}
C. S^{2-}
D. Ar
E. need a table to determine

MAR

Compare the elements Na, B, Al, and C with regard to the following properties:
Which has the largest (most negative) electron affinity?

A. Na
B. B
C. Al
D. C
E. Jq

MAR

Which of the following groups of elements is arranged correctly in order of increasing first ionization energy?

A. Mg < C < N < F
B. N < Mg < C < F
C. Mg < N < C < F
D. F < C < Mg < N
E. I need Google to answer this question

MAR

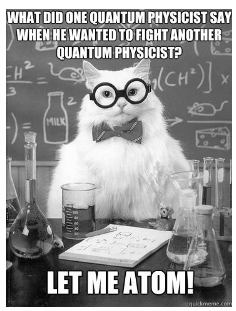
End of Review - *Good luck with your final exams!*

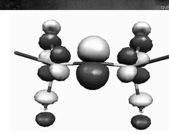
Need more practice?

- *Practice Problem Sets (online)*
- *Concept Guides (Companion and online)*
- *Chapter Guides (online)*
- *End of Chapter Problems in Textbook (every other question has answer at end)*

Good luck with your studying!

MAR





"Quick & Dirty" Metric Guide

Prefix	Base	Symbol	Acronym	Example with mass
kilo	10^3	k	*kiss*	kilogram, kg
hecto	10^2	h	*him/her*	hectogram, hg
deka	10^1	da	*dearly*	dekagram, dag
base	10^0	base	*but*	gram, g
deci	10^{-1}	d	*don't*	decigram, dg
centi	10^{-2}	c	*call*	centigram, cg
milli	10^{-3}	m	*me*	milligram, mg

Example: There are 10^3 grams in one kilogram,
or there are 10^{-3} kilograms in one gram

Example: There are 10^{-2} grams in one centigram,
or there are 10^2 centigrams in one gram

Example: There are 10^2 liters in one hectoliter,
or there are $(10^{-2}\ hL/L)(10^{-2}\ L/cL) = 10^{-4}\ hL\ /\ cL$

S.I. Base Units

Quantity	SI Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electric current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol

S.I. Derived Units

Quantity	Unit	Symbol	SI Units
Energy	joule	J	$kg\ m^{2}\ s^{-2}$
Force	newton	N	$kg\ m\ s^{-2}$
Power	watt	W	$kg\ m^{2}\ s^{-3}$
Electric charge	coulomb	C	A s
Electric potential	volt	V	$kg\ m^{2}\ s^{-3}\ A^{-1}$
Magnetic flux density	tesla	T	$kg\ s^{-2}\ A^{-1}$
Frequency	hertz	Hz	s^{-1} *(cycle per second)*

Metric Prefixes

deci	d	0.1	deca	da	10
centi	c	0.01	hecto	h	100
milli	m	0.001	kilo	k	1000
micro	μ	10^{-6}	mega	M	10^{6}
nano	n	10^{-9}	giga	G	10^{9}
pico	p	10^{-12}	tera	T	10^{12}
femto	f	10^{-15}	peta	P	10^{15}
atto	a	10^{-18}	exa	E	10^{18}

Also note these common *non-SI* units:

Angstrom	Å	10^{-10} m
Micron	μ	10^{-6} m
Calorie	cal	4.184 J
Gauss	G	10^{-4} T
Debye	D	$3.3356 * 10^{-30}$ C m

The "Quick & Dirty" Sig Fig* Chart

* Sig Fig = Significant Figures

<u>Rounding Up or Down</u>: *Round up **only** if first extraneous number is 5 or larger*

18.3<u>4</u> = 18.3
18.3<u>5</u> = 18.4
18.3<u>6</u> = 18.4

<u>Note</u>: *Round up or down only at the **end** of the calculation*

<u>Multiplying & Dividing</u>: *Use lowest number of sig figs in equation*

$$\frac{(4\ sig\ figs)}{(3\ sig\ figs)} \quad \frac{0.01208}{0.0236} = \mathbf{0.512}\ (3\ sig\ figs)$$

2.6 * 2.6 = **6.8**

<u>Adding & Subtracting</u>:

*# of decimal places in **answer** = # of decimal places in value with **fewest places***

140.76 *(hundredths)* - 22.1 *(tenths)* = **118.7** *(tenths)*

140 *(tens)* + 22.1 *(tenths)* = **160** *(tens, or $1.6 * 10^2$)*

CH 221 Chapter One Study Guide

- Define physical properties and chemical properties and be able to give examples of each.
- Recognize the different states of matter (solids, liquids and gases) and give characteristics of each.
- Understand the basic concepts of the Kinetic Molecular Theory of Matter.
- Appreciate the difference between matter represented at the macroscopic level and at the particulate ("microscopic") level.
- Convert temperatures between Celsius, Kelvin and Fahrenheit scales.
- Realize that the Kelvin scale does not use the degree symbol (˚).
- Know how to find density, mass or volume if two of the three quantities are known.
- Know the differences between atoms and elements as well as molecules and compounds.
- Be able to recognize chemical compounds and identify the number and identity of atoms in the compound.
- Be able to recognize (memorize?) at least the first 20 elements in the periodic table.
- Understand the difference between qualitative and quantitative measurements.
- Comprehend SI units and the prefixes that modify the sizes of metric units.
- Know how to use dimensional analysis in calculations, conversions, etc.
- Know the differences between accuracy and precision.
- Be able to report a value correctly using the appropriate number of significant figures.
- Use the concept of percent in chemistry.
- Be able to solve and understand the assigned problems in problem set #1.

The "q&d" Guide to Electric Charge

"q&d" = "Quick 'n' Dirty"

- Charge may be either of two types, POSITIVE or NEGATIVE
- **Protons** are positive and **Electrons** are negative
- **Neutrons** are neutral (no charge present)
- **Unlike charges attract** (i.e. protons and electrons) while **like charges repel** (i.e. protons and protons, electrons and electrons)
- Charge may be transferred from one object to another by contact or induction
- The force of attraction (F) is inversely proportional to the square of the distance (d) by **Coulomb's Law**:

$$F = k\frac{\left(n^{+}e\right)\left(n^{-}e\right)}{d^{2}}$$

n^+ = number of positive charges
n^- = number of negative charges
e = charge on an electron = $1.602*10^{-19}$ C
k = proportionality constant

Particle	Charge	Mass (g)	Mass (amu)
Proton	$+1.6*10^{-19}$ C	$1.7*10^{-24}$ g	1.0073
Neutron	zero	$1.7*10^{-24}$ g	1.0087
Electron	$-1.6*10^{-19}$ C	$9.1*10^{-28}$ g	$5.5*10^{-4}$

Remember: Atoms are usually electrically neutral, indicating ***equal numbers of protons and electrons!***

CH 221 Chapter Two Part 1 Study Guide

- Explain the historical development of the atomic theory and identify some of the key scientists who made important contributions to this field (Democritus, Dalton, Curie, Rutherford, Thompson, Millikan, Mendeleev, etc.)

- Know the differences between and identities of alpha, beta and gamma radioactive particles.

- Describe electrons, protons and neutrons and the general structure of the atom.

- Understand the atomic mass unit (amu) and elementary charge (e).

- Be able to calculate the atomic mass of an element from isotopic abundances.

- Define isotope and be able to give the mass number and number of neutrons for a specific isotope.

- Explain the difference between atomic number and atomic mass for an element. Be able to find this information from a periodic table.

- Memorize the value of Avogadro's Number to at least four significant figures ($6.022 * 10^{23}$).

- Explain the concept of the mole. Be able to find the mass per mole from the periodic table.

- Know how mass per mole relates to mass per atom on the periodic table and know how to use this in calculations.

- Understand how to convert from moles of an element to mass of an element and from the mass of an element to moles of an element.

- Be able to identify the following groups from the periodic table: metals, nonmetals, metalloids, alkali, alkaline earths, pnictogens, chalcogens, halogens, noble gases, transition metals, lanthanides and actinides.

- Use the periodic table to predict properties of elements.

- Be able to solve and understand the assigned problems in problem set #2.

Molecular Formula: Actual number of different atoms in molecule

Empirical Formula: Smallest whole number ratio of atoms in molecule

Structural Formula: Molecular formula written to designate molecule's orientation in space

Compound	Molecular	Empirical	Structural
Water	H_2O	H_2O	HOH
Hydrogen Peroxide	H_2O_2	HO	HOOH
Ethylene	C_2H_4	CH_2	H_2CCH_2
Ethane	C_2H_6	CH_3	H_3CCH_3
Ethanol	C_2H_6O	C_2H_6O	H_3CCH_2OH
Dimethyl ether	C_2H_6O	C_2H_6O	H_3COCH_3

“Quick & Dirty” Ionic Charge Guide

Cations (Positive Charges)

For Groups 1A, 2A and 3A: Charge on metal = **Positive**
Magnitude of charge = **Group number**

Examples:

Lithium (Group 1A)	***positive* +1 charge**	$\mathbf{Li^{+}}$
Strontium (Group 2A)	***positive* +2 charge**	$\mathbf{Sr^{+2}}$
Gallium (Group 3A)	***positive* +3 charge**	$\mathbf{Ga^{+3}}$

-=-

Anions (Negative Charges)

For Groups 5A, 6A and 7A: Charge on nonmetal = **Negative**
Magnitude of charge = **8 - Group number**

Examples:

Bromine (Group 7A)	***negative* -1 charge (8 - 7 = 1)**	$\mathbf{Br^{-}}$
Oxygen (Group 6A)	***negative* -2 charge (8 - 6 = 2)**	$\mathbf{O^{-2}}$
Nitrogen (Group 5A)	***negative* -3 charge (8 - 5 = 3)**	$\mathbf{N^{-3}}$

-=-

Remember: Noble Gases do not react! Inert!

Three Types of Compounds

1) Main Group metal + nonmetal

No variable charge or Greek prefixes

Al_2O_3	aluminum oxide
Na_2SO_4	sodium sulfate
NaCl	sodium chloride

2) Transition metal + nonmetal

Watch variable charge!

Fe_2O_3	iron(III) oxide
FeO	iron(II) oxide
Fe_2O_5	iron(V) oxide
$CuNO_3$	copper(I) nitrate

3) Nonmetal + nonmetal (covalent)

Use Greek di-, tri-, etc.
Least electronegative (i.e. more metallic) named first
Least electronegative does not need "mono"

P_2O_3	diphosphorus trioxide
P_2O_5	diphosphorus pentaoxide
S_2Cl_{10}	disulfur decachloride
OF_2^{-1}	oxygen difluoride
NO	nitrogen monoxide
NO_2	nitrogen dioxide
N_2O	dinitrogen monoxide

Greek Prefixes for Naming Multiple Atoms

1	*mono*	*6*	*hexa*
2	*di*	*7*	*hepta*
3	*tri*	*8*	*octa*
4	*tetra*	*9*	*nona*
5	*penta*	*10*	*deca*

Common Polyatomic Ions and the Corresponding Acids

There is a pattern associated with many of the polyatomic ions in chemistry that can aid you when learning names and the relationships with the corresponding acids. Some combinations of a central atom with oxygen are found more often in nature, and they are designated the "common" form of the polyatomic... yet due to oxygen's "social nature", several other combinations of the central atom with oxygen can exist. A pattern exists which relates the number of oxygen atoms relative to the "common" form... and this pattern can be extended to a host of oxygen-containing acids.

First, remember this phrase:

"**Nick** the **Camel Brat** ate **Icky Clam** for **Supper** in **Phoenix**"

This phrase helps you remember the **central atom**, the **number of oxygen atoms in the "common" form** of the polyatomic, and the **charge** on the polyatomic ion. ***All of the common form polyatomic ions get an "ate" suffix.***

- The **number of consonants** = the **number of oxygen atoms** in the common form of the polyatomic ion
- The **number of vowels** = the **negative charge** on the polyatomic ion

Nick = **nitr**ate, NO_3^{-1}
Camel = **carb**onate, CO_3^{-2}
Brat = **brom**ate, BrO_3^{-1}
Icky = **iod**ate, IO_3^{-1} *(note that* ***y*** *is a consonant and not a vowel in this context!)*
Clam = **chlor**ate, ClO_3^{-1}
Supper = **sulf**ate, SO_4^{-2}
Phoenix = **phosph**ate, PO_4^{-3}

- Polyatomic ions in the **common** form have an "**ate**" suffix (i.e. chlor**ate**, ClO_3^{-1})
- Polyatomic ions with **one more oxygen** than the common form get a "**per**" prefix and an "**ate**" suffix (i.e. **per**chlor**ate**, ClO_4^{-1})
- Polyatomic ions with **one less oxygen** than the common form get an "**ite**" ending (i.e. chlor**ite**, ClO_2^{-1})
- Polyatomic ions with **two less oxygen atoms** than the common form get a "**hypo**" prefix and the "**ite**" suffix (i.e. **hypo**chlor**ite**, ClO^{-1})

The following table shows the various polyatomic ions and all of their known variations:

	nitrogen	*carbon*	*bromine*	*iodine*	*chlorine*	*sulfur*	*phosphorus*
-2 oxygen	-	-	**hypo**brom**ite**, BrO^{-1}	**hypo**iod**ite**, IO^{-1}	**hypo**chlor**ite**, ClO^{-1}	-	-
-1 oxygen	nitr**ite**, NO_2^{-1}	-	brom**ite**, BrO_2^{-1}	iod**ite**, IO_2^{-1}	chlor**ite**, ClO_2^{-1}	sulf**ite**, SO_3^{-2}	phosph**ite**, PO_3^{-3}
common	nitr**ate**, NO_3^{-1}	carbon**ate**, CO_3^{-2}	brom**ate**, BrO_3^{-1}	iod**ate**, IO_3^{-1}	chlor**ate**, ClO_3^{-1}	sulf**ate**, SO_4^{-2}	phosph**ate**, PO_4^{-3}
+1 oxygen	-	-	**per**brom**ate**, BrO_4^{-1}	**per**iod**ate**, IO_4^{-1}	**per**chlor**ate**, ClO_4^{-1}	-	-

Entries with a "-" are not known to exist and can be ignored.

Polyatomic ions readily make acids. An acid is a compound with a hydrogen atom that reacts readily with other substances. In chemistry, we list the acidic hydrogen first to designate its reactivity.

As before, a naming pattern exists for acids containing an oxygenated polyatomic ion:

- Acidic polyatomic ions in the **common** form have an "**ic acid**" suffix (i.e. chlor**ic acid**, $HClO_3$)
- Acidic polyatomic ions with **one more oxygen** than the common form get a "**per**" prefix and an "**ic acid**" suffix (i.e. **per**chlor**ic acid**, $HClO_4$)
- Acidic polyatomic ions with **one less oxygen** than the common form get an "**ous acid**" ending (i.e. chlor**ous acid**, $HClO_2$)
- Acidic polyatomic ions with **two less oxygen atoms** than the common form get a "**hypo**" prefix and the "**ous acid**" suffix (i.e. **hypo**chlor**ous acid**, HClO)
- Acidic polyatomic ions with **no oxygen atoms** get a "**hydro**" prefix and the "**ic acid**" suffix (i.e. **hydro**chlor**ic acid**, HCl)

The following table shows the acidic form of the polyatomic ions with all of their known variations:

	nitrogen	*carbon*	*bromine*	*iodine*	*chlorine*	*sulfur*	*phosphorus*
no oxygen	-	-	**hydro**brom**ic acid**, HBr	**hydro**iod**ic acid**, HI	**hydro**chlor**ic acid**, HCl	**hydro**sulfur**ic acid**, H_2S	-
-2 oxygen	-	-	**hypo**brom**ous acid**, HBrO	**hypo**iod**ous acid**, HIO	**hypo**chlor**ous acid**, HClO	-	-
-1 oxygen	nitr**ous acid**, HNO_2	-	brom**ous acid**, $HBrO_2$	iod**ous acid**, HIO_2	chlor**ous acid**, $HClO_2$	sulfur**ous acid**, H_2SO_3	phosphor**ous acid**, H_3PO_3
common	nitr**ic acid**, HNO_3	carbon**ic acid**, H_2CO_3	brom**ic acid**, $HBrO_3$	iod**ic acid**, HIO_3	chlor**ic acid**, $HClO_3$	sulfur**ic acid**, H_2SO_4	phosphor**ic acid**, H_3PO_4
+1 oxygen	-	-	**per**brom**ic acid**, $HBrO_4$	**per**iod**ic acid**, HIO_4	**per**chlor**ic acid**, $HClO_4$	-	-

Finally, please note that this list is not 100% inclusive... but similar patterns can be applied to polyatomic ions not on this list. For example,

- H_2SeO_4 = selenic acid *and* H_2SeO_3 = selenous acid
- AsO_4^{-3} = arsenate ion *and* AsO_3^{-3} = arsenite ion

And if you cannot get enough polyatomic ions… here's another useful phrase:

"**Simon** and **Bonnie Aspired** to **Search** the **Creepy Count** for the **Icky Clam**"

Simon = SiO_3^{2-} = silicate
Bonnie = BO_3^{3-} = borate
Aspired = AsO_4^{2-} = arsenate
Search = SeO_4^{2-} = selenate
Creepy = CrO_4^{2-} = chromate
Count = CO_3^{2-} = carbonate
Icky = IO_3^{1-} = silicate
Clam = ClO_3^{1-} = chlorate

"Quick & Dirty" Rules for Ionic Compounds

I. *Generally*, metals form *ionic* compounds

II. Nonmetals *generally* form *ionic* compounds when allowed to react with metals

III. *Metalloids* are difficult to predict

IV. The farther apart the elements are in the periodic table, the *better the chances* the compound will be ionic

Ex: $Na^+ + Cl^- \rightarrow NaCl$

Ex: $Cd^{2+} + Te^{2-} \rightarrow CdTe$ (?)

CH 221 Chapter Two Part 2 Study Guide

- Define molecular formula, empirical formula and structural (or "condensed") formula and know how to use them. Define allotropes and give several examples.

- List the elements that exist as diatomic molecules and be able to predict which elements are monatomic (Noble gases).

- Understand the definitions of cation (positive charge) and anion (negative charge). Metals are usually cations (lose electrons) while nonmetals will often be anions (gain electrons).

- Understand the "quick and dirty" ionic charge guide for predicting the ionic charges on atoms in groups 1A-3A and 5A-8A. Recognize that transition metal elements often exist in a variety of positively charged "oxidation" states.

- Be able to give the names, formulas and ionic charges for the polyatomic ions listed in the textbook and in the Nomenclature lab.

- Be able to write the formulas for a number of ionic compounds using groups 1A-3A and 5A-7A.

- Explain the general properties of ionic compounds. Understand the importance of Coulomb's Law and how it relates to electrostatic forces. We will be revisiting this concept in CH 222.

- Be able to determine the name of an ionic (metal plus nonmetal) or covalent (nonmetal plus nonmetal) compound using the rules outlined in this chapter.

- Understand the concepts of formula mass and molar mass (i.e. molecular weight) and how they relate to the mole and Avogadro's number. Be able to calculate the molar mass for *any* given compound. Master the skills necessary to convert moles to grams and grams to moles.

- Understand and be able to use percent composition in relation to empirical formulas.

- Understand the difference between empirical and molecular formulas and what is needed to calculate the molecular formula from an empirical formula (i.e. a molar mass determination such as from mass spectrometry).

- Be able to use experimental data to calculate the number of water molecules in a hydrated compound.

- Be able to solve and understand the assigned problems in problem set #3.

Stoichiometry Calculations

$$\begin{tikzcd}[column sep=huge, row sep=huge]
\textbf{grams of A} \arrow[d, "*\left(\frac{mol\ A}{grams\ A}\right)"'] \arrow[r, dashed, "\text{Direct calculation not possible}", "\times"'] & \textbf{grams of B} \\
\textbf{moles of A} \arrow[r, "\text{* Stoichiometric Factor}", "*\left(\frac{\beta\ mol\ B}{\alpha\ mol\ A}\right)"'] & \textbf{moles of B} \arrow[u, "*\left(\frac{grams\ B}{mol\ B}\right)"']
\end{tikzcd}$$

Valid for:

$$\boldsymbol{\alpha}\ \mathbf{A} + \boldsymbol{\beta}\ \mathbf{B} \rightarrow \gamma\ C + \delta\ D \quad \textit{or}$$
$$\boldsymbol{\alpha}\ \mathbf{A} + \gamma\ C \rightarrow \boldsymbol{\beta}\ \mathbf{B} + \delta\ D \quad \textit{or}$$
$$\gamma\ C + \boldsymbol{\beta}\ \mathbf{B} \rightarrow \boldsymbol{\alpha}\ \mathbf{A} + \delta\ D \quad \textit{or}$$
$$\gamma\ C + \delta\ D \rightarrow \boldsymbol{\alpha}\ \mathbf{A} + \boldsymbol{\beta}\ \mathbf{B}$$

CH 221 Chapter Four Part 1 Study Guide

- Be able to balance simple chemical equations and understand the information conveyed by the equation (number of moles of reactants, etc.)

- Know how to interpret a chemical equation – states of matter, quantity of reacting materials, etc.

- Understand how to convert from the mass of element A to the mass of element B using chemical equations. Remember to travel through the "molar bridge" when converting masses.

- Understand the concept of stoichiometric factor and be able to convert between different quantities (mass, moles, products, reactants, etc.)

- Define limiting reagent and know how to determine which reactant is limiting.

- Explain the differences between actual yield, theoretical yield and percent yield. Know how to calculate or determine the values for each type of yield.

- Be able to use stoichiometry principles to analyze a mixture or to find the empirical formula of an unknown compound.

- Understand the process whereby element percentages (i.e. %C) arrive.

- Be able to solve and understand the assigned problems in problem set #4.

Dissolve, Dissociate and Electrolyte Guide

$AgCl_{(s)} \xrightarrow{H_2O}$ ***No reaction,*** *does not dissolve or dissociate*

silver(I) chloride

$NaCl_{(s)} \xrightarrow{H_2O} NaCl_{(aq)} \xrightarrow{\text{time}} Na^+_{(aq)} + Cl^-_{(aq)}$

sodium chloride — *dissolves*

100% dissociation
strong electrolyte
conducts electricity

$C_6H_{12}O_{6(s)} \xrightarrow{H_2O} C_6H_{12}O_{6(aq)} \xrightarrow{\text{time}} C_6H_{12}O_{6(aq)}$

sugar — *dissolves*

No dissociation
nonelectrolyte
does not conduct electricity

$$HC_2H_3O_{2(l)} \xrightarrow{H_2O} HC_2H_3O_{2(aq)} \xrightarrow{\text{time}} \begin{cases} HC_2H_3O_{2(aq)} & (99\%) \\ H^+_{(aq)} & (\sim 1\%) \\ C_2H_3O_2^-{}_{(aq)} & (\sim 1\%) \end{cases}$$

acetic acid — *dissolves*

Partial dissociation
weak electrolyte
partial conductor of electricity

Balancing Net Ionic Equations

1. Balance the **molecular equation**. Find stoichiometric coefficients; do not change the subscripts or states of matter.

   Example: Balance the double displacement reaction between sodium hydroxide and aluminum chloride.

   $$NaOH_{(aq)} + AlCl_{3(aq)} \text{ -> } NaCl_{(aq)} + Al(OH)_{3(s)}$$

   $NaOH$, $AlCl_3$ and $NaCl$ are strong electrolytes; $Al(OH)_3$ is insoluble in water, so:

   $$\mathbf{3\ NaOH_{(aq)} + AlCl_{3(aq)} \text{ -> } 3\ NaCl_{(aq)} + Al(OH)_{3(s)}}$$

2. Write the **total ionic equation** by rewriting the molecular equation with the strong electrolytes separated into ions. Do not "ionize" solids, liquids or gases; only aqueous species should be separated.

   Example:

   $$\mathbf{3\ Na^+_{(aq)} + 3\ OH^-_{(aq)} + Al^{3+}_{(aq)} + 3\ Cl^-_{(aq)} \text{ -> } 3\ Na^+_{(aq)} + 3\ Cl^-_{(aq)} + Al(OH)_{3(s)}}$$

3. Write the **net ionic equation** by rewriting the total ionic equation and canceling the **spectator ions** (the species that appear on both the product and reactant sides of the total ionic equation.) Remember that atoms are *not* the same as ions (i.e. $Mg_{(s)}$ is not the same as $Mg^{2+}_{(aq)}$.)

   Example: Na^+ and Cl^- appear on both sides of the equation, so they are spectator ions

   $$\cancel{3\ Na^+_{(aq)}} + 3\ OH^-_{(aq)} + Al^{3+}_{(aq)} + \cancel{3\ Cl^-_{(aq)}} \text{ -> } \cancel{3\ Na^+_{(aq)}} + \cancel{3\ Cl^-_{(aq)}} + Al(OH)_{3(s)}$$

   $$\mathbf{Al^{3+}_{(aq)} + 3\ OH^-_{(aq)} \text{ -> } Al(OH)_{3(s)}}$$

4. **Check** that the **total ionic charge** on the reactant side balances the total ionic charge on the product side. The equation will now be balanced for both mass and charge.

   Example: $\mathbf{Al^{3+}_{(aq)} + 3\ OH^-_{(aq)} \text{ -> } Al(OH)_{3(s)}}$ To check the total ionic charge:
   Reactant side: **+3** (from Al^{3+}) **+3(-1)** (from OH^-) = **0**
   Product side: **0** (no charge on molecular solids)

   Since the charge on the reactant side equals the charge on the product side, the total ionic charge for this reaction is balanced.

Stoichiometry Calculations II

$$
\begin{tikzcd}[column sep=huge, row sep=huge]
\textbf{grams of A} \arrow[d, "*\left(\frac{mol\ A}{grams\ A}\right)"'] \arrow[r, dashed, "\text{Direct calculation not possible}", "\times"'] & \textbf{grams of B} \\
\textbf{moles of A} \arrow[r, "\text{* Stoichiometric Factor}", "*\left(\frac{\beta\ mol\ B}{\alpha\ mol\ A}\right)"'] & \textbf{moles of B} \arrow[u, "*\left(\frac{grams\ B}{mol\ B}\right)"'] \arrow[d] \\
\textbf{(Molarity of A)*(Liters of A)} \arrow[u] & \textbf{(Molarity of B)*(Liters of B)}
\end{tikzcd}
$$

$\boxed{\textit{Molarity} = \textit{mol / Liter}}$

$\boxed{\textit{1000 mL = 1 L}}$

Combustion Reactions:

$$C_mH_n + x\,O_{2(g)} \rightarrow y\,CO_{2(g)} + z\,H_2O_{(g)}$$

Acid-Base Reactions:

$$\{\text{acid}\} + \{\text{base}\} \rightarrow \{\text{salt}\} + H_2O_{(l)}$$

Precipitation Reactions:

$$M_iX_{i(aq)} + M_jX_{j(aq)} \rightarrow M_iX_{j(s)} + M_jX_{i(aq)}$$

Gas-Forming Reactions:

$$\text{A} + \text{B} \rightarrow \{\text{gas } \textit{or } H_2CO_3\} + \text{other product(s)}$$

Oxidation-Reduction Reactions:

$$Cu_{(s)} + 2\,Ag^{+}{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + 2\,Ag_{(s)}$$

Determining Oxidation Numbers

1) Each atom in a pure element has an oxidation number = 0.

2) For ions consisting of a single atom, the oxidation number is equal to the charge on the ion.

3) **Fluorine** is ***always*** -1.

4) **Chlorine**, **bromine** and **iodine** are ***always*** -1 ***except*** when combined with oxygen or fluorine.

5) The oxidation number of **hydrogen** is +1 and of **oxygen** is -2. ***Exceptions:*** hydrides (H^{-1}), peroxides (O^{-1}), OF compounds.

6) In neutral compounds, the sum of the oxidation numbers must be zero. In polyatomic ions, the sum of the oxidation numbers must be equal to the ion charge.

Examples:

Cu(s)	Cu: 0	**NaCl**	Cl: -1 Na: +1
NaF	F: -1 Na: +1	$\mathbf{ClO^-}$	Cl: +1 O: -2
$\mathbf{Al^{3+}}$	Al: +3	$\mathbf{Fe_2O_3}$	Fe: +3 O: -2
$\mathbf{H_2O_2}$	H: +1 O: -1	$\mathbf{AlH_4^-}$	Al: +3 H: -1

$\mathbf{KMnO_4}$ K: +1 Mn: +7 O: -2

CH 221 Chapter Four Part 2 Study Guide

- Understand what an electrolyte is. Know the differences between strong electrolytes, weak electrolytes, and non-electrolytes. Be able to give examples of each category.

- Be able to predict the solubility of ionic compounds in water. Be able to recognize what types of ions are created upon dissolving ionic compounds or acids and bases in water. Know how to predict the products of precipitation reactions by looking at the cations and anions.

- Define acids and bases and know their characteristic behavior towards each other. Be able to recognize acid-base equations. Memorize the names of the common acids and bases.

- Understand the differences between strong and weak acids *and* strong and weak bases. Memorize the neutralization reaction and know when it applies. Understand that the net ionic equation for the reaction of a strong acid and strong base will *always* be $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$ Know how to calculate the pH of a solution with a strong acid. Know how to find $[H_3O^+]$ from the pH value.

- Know the general formula for combustion reactions, including anticipated products and reactants. Be able to recognize precipitation reactions and gas-forming reactions. Know the importance of H_2CO_3. Understand net ionic equations and be able to derive them from normal chemical equations.

- Be able to give the oxidation number of any element or compound. Oxidation numbers are *very important* for many chemical systems.

- Know the definitions of reduced, oxidized, reducing agent and oxidizing agent. Be able to recognize an oxidation-reduction reaction.

- Define molarity, solute, solvent and solution. Know how to calculate molarity, volume and moles if only two of the three quantities are known.

- Understand the importance of dilution in the chemistry laboratory. Know how to utilize the formula $\mathbf{M_1V_1 = M_2V_2}$ (also known as $\mathbf{c_1V_1 = c_2V_2}$). Be able to derive this equation from $moles_1$ and $moles_2$ if required.

- Be able to solve stoichiometric problems using solution concentrations and volumes. Explain how a titration is performed. Understand the significance of standardization. Know the definitions for indicators and the equivalence point. Be able to calculate concentrations or amounts of reactants using titration data.

- Be able to solve and understand the assigned problems in problem set #4 and #5.

Key Concepts for Enthalpy Reactions

- When a reaction is reversed, the *magnitude* of ΔH remains the same, but the *sign* of ΔH changes.

$$H_{2(g)} + {}^{1}/_{2}\, O_{2(g)} \rightarrow H_2O_{(g)} \quad \Delta H^{\circ}_{f} = -285 \text{ kJ/mol}$$
$$H_2O_{(g)} \rightarrow H_{2(g)} + {}^{1}/_{2}\, O_{2(g)} \quad \Delta H^{\circ}_{f} = +285 \text{ kJ/mol}$$

- When a balanced equation for a reaction is multiplied by an integer, the value of ΔH for that reaction must be multiplied by the same integer.

$$2\,\{H_{2(g)} + {}^{1}/_{2}\, O_{2(g)} \rightarrow H_2O_{(g)}\} \;=\; 2\,H_{2(g)} + O_{2(g)} \rightarrow 2\,H_2O_{(g)} \quad \Delta H^{\circ}_{f} = 2(-285 \text{ kJ/mol}) = \mathbf{-570\ kJ/mol}$$

- The change in enthalpy for a given reaction can be calculated from the enthalpies of formation of the reactants and products:

$$\boldsymbol{\Delta H^{\circ}_{\text{reaction}}} = \Sigma n_p \Delta H^{\circ}_{f}(\text{products}) - \Sigma n_r \Delta H^{\circ}_{f}(\text{reactants})$$

- Elements in their *standard states* are not included in the $\Delta H_{\text{reaction}}$ calculations (*i.e.* ΔH°_{f} for an element in its standard state is zero.)

- Reactions may be combined through addition and subtraction to provide **net** values of $\Delta H^{\circ}_{\text{reaction}}$ (**Hess's Law**).

CH 221 Chapter Five Study Guide

- Understand the terms reactant- and product-favored. We shall study these again during CH 223.
- Understand the difference between kinetic and potential energy. Know the general equations from physics for each energy type.
- Be able to utilize the joule in energy and heat calculations. Be able to convert between joules and calories, and be aware of the differences between calories and Calories.
- Be able to use specific heat in calculations. Know how to utilize the magnitude of specific heat to predict temperature changes, etc.
- Understand the sign conventions of **q** regarding heat transfer.
- Be able to use heat of fusion and heat of vaporization values to find the quantity of thermal energy involved in changes of state. Be able to apply the system and surroundings concepts to chemical reactions.
- Understand the definition of exothermic and endothermic and be able to predict these if given the sign of ΔT, ΔH or q_{sys}.
- Know the first law of thermodynamics (the law of energy conservation).
- Understand enthalpy. Enthalpy must be measured relative to something else (a "change in"), and enthalpy is not a specific value for a given reactant. Know that $\Delta H = q_p$. Is $\Delta H = q_v$ true? ***(No!** What is q_v equal to?)*
- Explain what a state function is and give examples of state functions and non-state functions.
- Be able to apply Hess' Law to find values of enthalpy.
- Know the definition of standard conditions (i.e., °) in thermodynamics.
- Be able to write balanced chemical equations that define the standard molar enthalpy of formation, ΔH°_f, for a compound.
- Know the difference between the standard molar enthalpy of formation, ΔH°_f, and the enthalpy change for a reaction, ΔH°_{rxn}.
- Understand the theory of calorimetry as discussed in lab and lecture.
- Be able to solve and understand the assigned problems in problem set #5.

CH 221 Guide to Quantum Numbers

Quantum Number	***Quantum Name***	***Values***
n	**shell**	**$1, 2, 3, 4, \ldots \infty$**
l	**subshell**	**$0, 1, 2, \ldots (n - 1)$**
m_l	**orbital**	**$-l \ \ldots 0 \ldots +l$**
m_s	**electron spin**	**$+^1/_2$ *or* $-^1/_2$**

Each electron in an atom can have its own unique "address" or *set of quantum numbers*.

Example: Consider a **Beryllium** atom with four electrons. Beryllium is in the second period, so possible n values are 1 and 2. Electrons are filled using the lowest value of n (or n + l), so the electrons will be placed into the n=1 shell before they enter the n=2 shell.

<u>When n = 1</u>, the only allowed value of l is 0; likewise, the only allowed value of $m_l = 0$. We will place the first two electrons in a 1s orbital. Each electron can have either a "spin up" ($m_s = +^1/_2$) or "spin down" ($m_s = -^1/_2$) configuration.

The first electron's set of quantum numbers (or address) will be: **$n = 1, l = 0, m_l = 0, m_s = +^1/_2$**

The second electron's set of quantum numbers (or address) will be: **$n = 1, l = 0, m_l = 0, m_s = -^1/_2$**

<u>When n = 2</u>, allowed values of l are 0 and 1. Lowest (n + l) values are filled first; hence, a (n + l) value of (2 +0) = 2 will be filled before a (n + l) value of (2 + 1) = 3. When $l = 0$, the only allowed value of $m_l = 0$. We will place the next two electrons in a 2s orbital. Each electron can have either a "spin up" ($m_s = +^1/_2$) or "spin down" ($m_s = -^1/_2$) configuration.

The third electron's set of quantum numbers (or address) will be: **$n = 2, l = 0, m_l = 0, m_s = +^1/_2$**

The fourth electron's set of quantum numbers (or address) will be: **$n = 2, l = 0, m_l = 0, m_s = -^1/_2$**

CH 221 Chapter Six Part 1 Study Guide

- Be able to define wavelength, frequency, wave amplitude and node.
- Understand the relationship between frequency, wavelength and the speed of light; know how to use this relationship in calculations.
- Know the difference between standing waves and moving waves.
- *Memorize* the value for the speed of light, **c** = $2.998*10^8$ m/s.
- Know the *relative positions* of these sections of the electromagnetic spectrum: visible, ultraviolet, infrared, radio, gamma, X-ray and microwaves.
- Understand the relationships amongst the energy of a photon, the frequency of the photon and Planck's constant. Be able to convert the frequency to wavelength if required; also be able to convert between one photon and a mole of photons.
- *Memorize* the value for Planck's constant, **h** = $6.626*10^{-34}$ J·s.
- Be able to describe in general terms the Bohr model for the hydrogen atom. Be able to explain how it accounts for the emission line spectra of excited atoms.
- Be able to calculate the energy levels of the hydrogen atom using the Bohr equation. You will *not* have to memorize neither this equation nor the Rydberg constant.
- Understand the de Broglie equation and know how it is used and for what systems.
- Recognize the significance of wave or quantum mechanics in describing the modern view of atomic structure.
- Understand that an orbital for an electron in an atom corresponds to an allowed energy of that electron.
- Know that the position of the electron is not known with certainty due to the Heisenberg uncertainty principle; only the probability of the electron being within a given region of space can be calculated.
- Be able to describe the allowed energy states of an electron in an atom using the quantum numbers **n**, ***l*** and $\mathbf{m}_l$. Be able to describe the shapes of the orbitals.
- Be able to solve and understand the assigned problems in problem set #6.

Predicting Atomic Electron Configurations

1) Electrons occupy the lowest energy orbitals available - *the **n+l Rule***
Begin assigning electrons at **1s** and continue in the following order:

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p *etc.*

Examples: **Li**: $1s^2 2s^1$ **Na**: $1s^2 2s^2 2p^6 3s^1$ **Ca**: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

2) **s** orbitals have one subshell; **p** orbitals have three subshells; **d** orbitals have five subshells; **f** orbitals have seven subshells. Or:

Number of subshells = 2l+1

where $\mathbf{m_l = \{l, (l-1), \ldots, 0, \ldots, (-l+1), -l\}}$ … *and* $\mathbf{l = \{0, 1, \ldots n - 1\}}$

3) No two electrons in an atom can have the same set of four quantum numbers - ***Pauli Exclusion Principle***. Each subshell can hold only two electrons, and the two electrons must have opposite values of spin (i.e. $\mathbf{m_s}$).

4) The most stable arrangement of electrons is that with the maximum number of unpaired electrons - ***Hund's Rule***. Single electrons must occupy every subshell in an orbital before they "pair up" or are "spin paired".

Example: **Ti**: $[Ar]3d^2 4s^2$ Titanium has two unpaired electrons

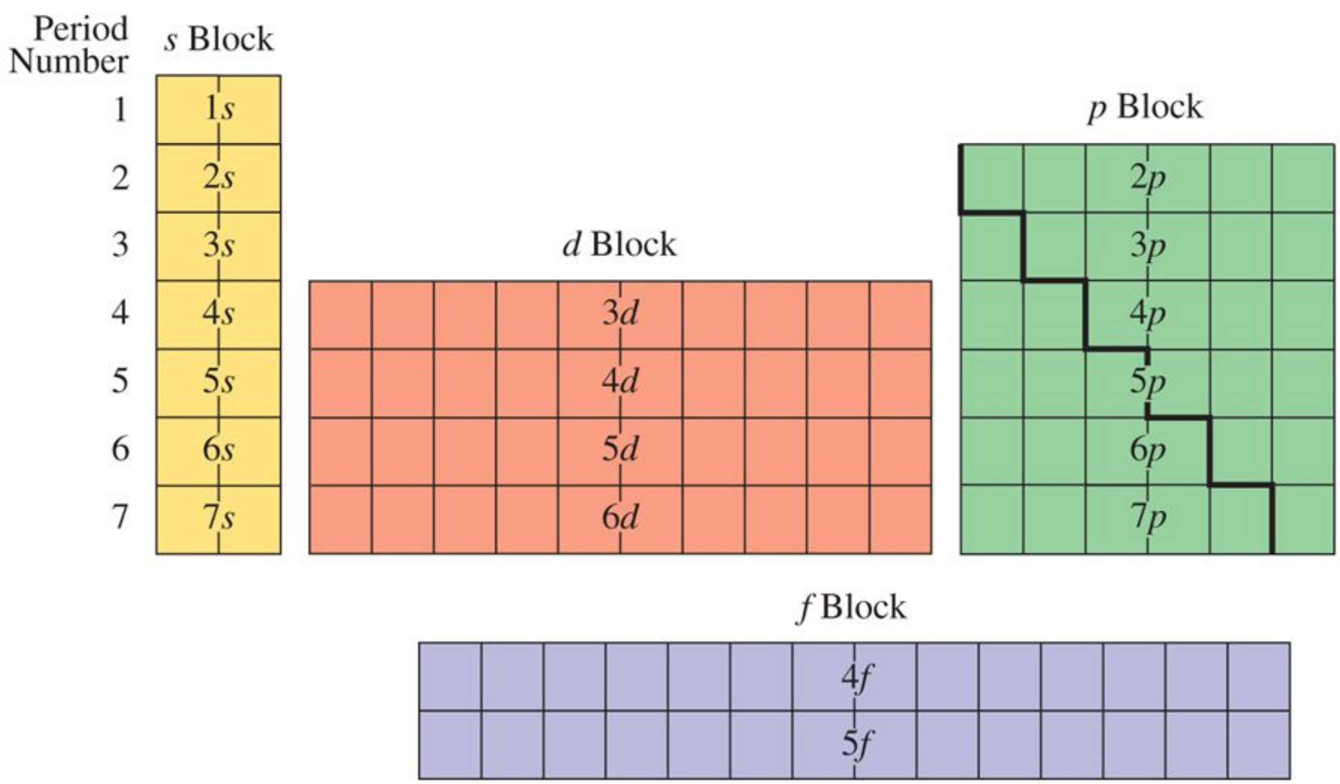
5) ***Paramagnetic*** compounds contain unpaired electrons.
Diamagnetic compounds contain electrons that are exclusively "spin paired." No unpaired electrons exist in diamagnetic compounds.

Examples: **Zn**: $[Ar]3d^{10} 4s^2$ *(diamagnetic)* **Li**: $[He]2s^1$ *(paramagnetic)*

6) ***Atomic ion configurations*** can be assigned using the rules given above and while remembering that the electrons easiest to remove will generally come from the highest energy orbital available.

Examples: **Cu**: $[Ar]3d^{10} 4s^1$ $\mathbf{Cu^{2+}}$: $[Ar]3d^9$

Periodic Table Blocks



CH 221 "q&d" Guide to Periodic Trends

"q&d" = "Quick 'n' Dirty"

- Each column represents a family or group of elements whose properties are similar

- Periodic Trends are ***generalizations***; subtle deviations in trend exist for almost all periodic properties

- **Metallic Properties** increase going down a column and moving left across a period

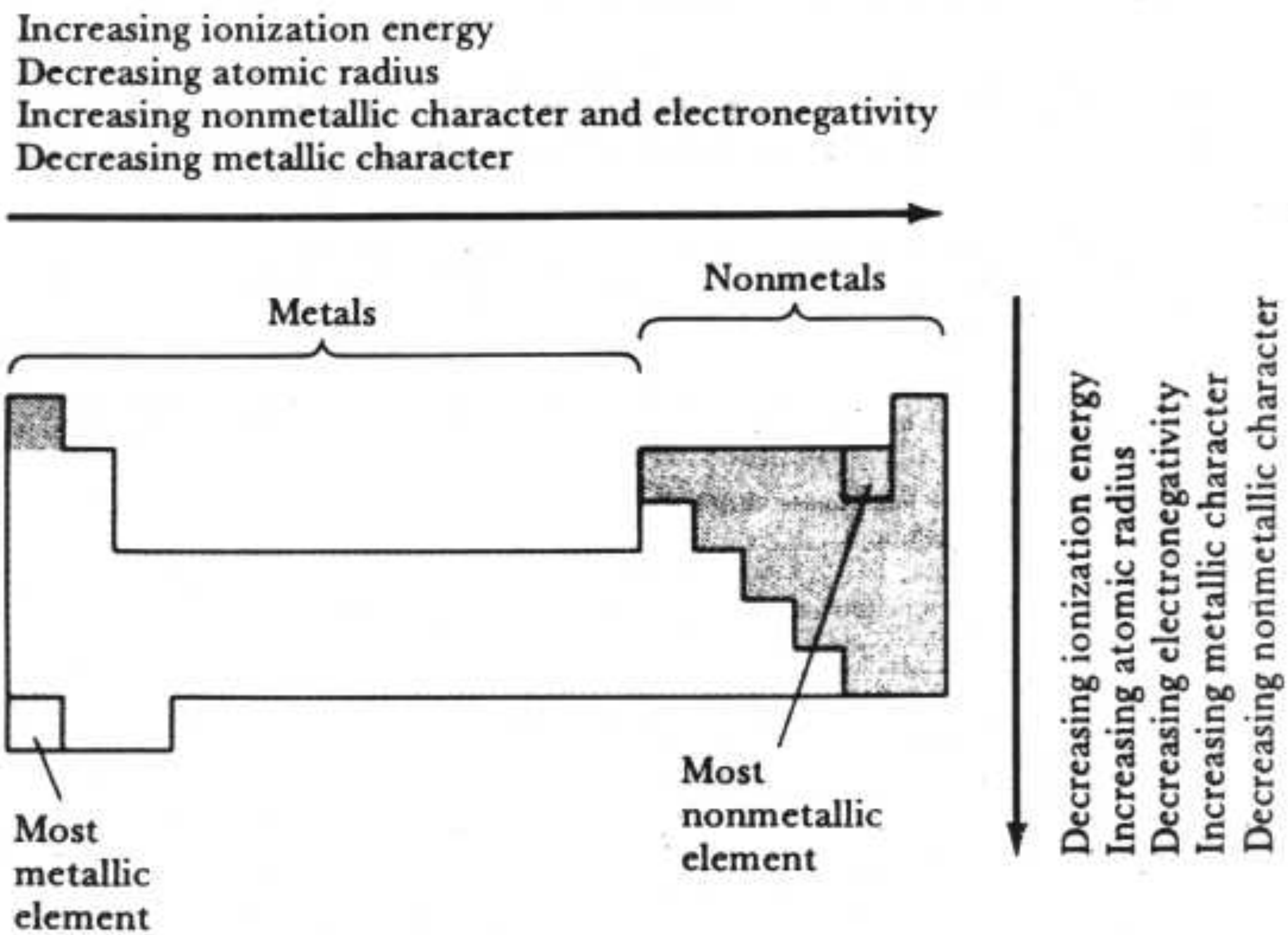
- **Atomic Size** increases going down a column and moving left across a period

- **Cationic Size** increases going down a column and moving left across a period

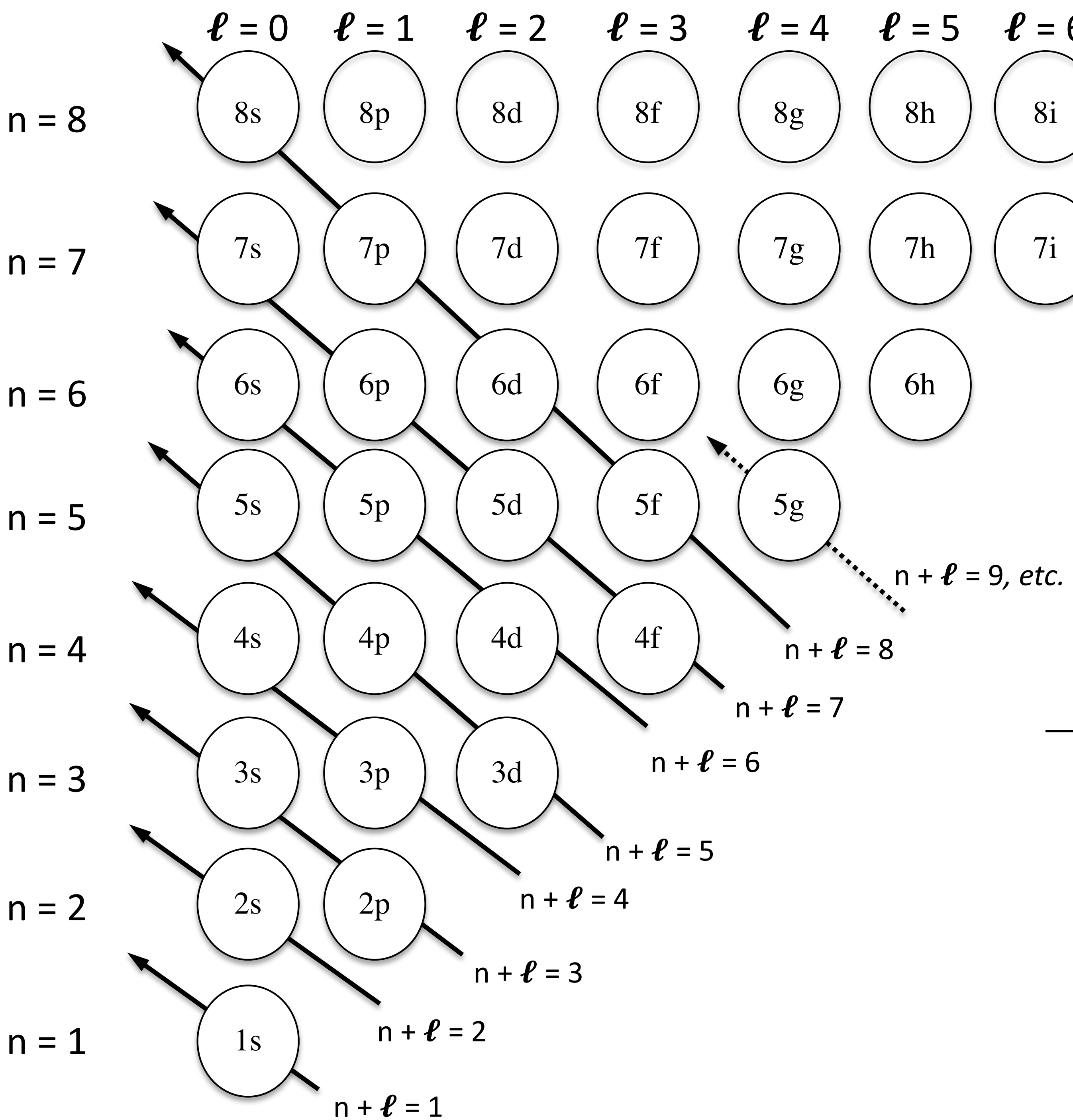
- **Anionic Size** increases going down a column and moving left across a period

- **First Ionization Energy** increases going up a column and moving right across a period ($M \rightarrow M^+ + e^-$)

- **Electron Affinity** increases going up a column and moving right across a period *if* the noble gases are omitted ($X + e^- \rightarrow X^-$)


Increasing ionization energy
Decreasing atomic radius
Increasing nonmetallic character and electronegativity
Decreasing metallic character
Metals
Nonmetals
Most metallic element
Most nonmetallic element
Decreasing ionization energy
Increasing atomic radius
Decreasing electronegativity
Increasing metallic character
Decreasing nonmetallic character

Aufbau Diagram



ℓ	letter	
0	s	*stupid*
1	p	*people*
2	d	*drive*
3	f	*freakin'*
4	g	*gas*
5	h	*hogs*

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p…

CH 221 Chapter Six Part 2 Study Guide

- Be able to classify substances as diamagnetic or paramagnetic.
- Realize that most paramagnetic compounds are not normally attracted to magnetic fields, with the exception of ferromagnetic and antiferromagnetic materials.
- Know that the spin quantum number, m_s, has values of + $^1/_2$ and - $^1/_2$. Know what these values refer to in the presence of a magnetic field.
- Recognize that each electron in an atom has a different set of the four quantum numbers - the Pauli Exclusion Principle.
- Recognize that the Pauli Exclusion Principle leads to the conclusion that no atomic orbital can be assigned more than two electrons *and* that the two electrons must have opposite spins (i.e. opposite values of m_s.)
- Using the periodic table as a guide, be able to depict electron configuration of the elements and monatomic ions by the orbital box notation or the spectroscopic notation. Understand the significance and relevance of the noble gas notation.
- Understand that electrons are generally assigned to the subshells of an atom in order of increasing subshell energy.
- Recognize that subshell energies in the hydrogen atom depend on both the n and *l* quantum numbers.
- When assigning electrons to atomic orbitals, be able to apply the Pauli Exclusion Principle and Hund's rule.
- Predict how properties of atoms - size, ionization energy and electron affinity - change on moving down a group or across a period of the periodic table.
- Be able to solve and understand the assigned problems in problem set #6.

CH 221 Introductory Mathematics Concept Guide

Right from the beginning of your study of chemistry you will face the task of solving problems. The goal of this section is to provide you with a review of the problem-solving skills that you will need.

1. Algebra Basics

a. Variables

Algebra is used to solve problems where some of the variables are known, but others are not. The unknowns are represented by letters. Problems are represented as equations and solved by finding the value of the unknown that makes the equation a true statement.

When solving equations, we can perform any operation (addition, subtraction, multiplication, or division) as long as we appropriately do the same thing to both sides of the equation.

When a number and a variable are written together, multiplication is indicated. 4x means four times x. Like terms can be added or subtracted.

$3x + 4x = 7x$.

Unlike terms cannot be combined.

$3x + 4x + 2y = 7x + 2y$

There may be situations for which you will be required to add algebraic equations together. To do this, line up the equal signs and add each side of the equations separately.

b. Notation Conventions

There are certain notation conventions that you may come across in example or solved problems in your chemistry work. Understanding these rules will not only help you follow problem solutions, but will help you keep track of your math as you solve problems on your own.

To avoid confusion with the variable x, a multiplication dot (·) is frequently used instead of the times symbol (x). When equations contain parentheses, do the operation within the parentheses first. When parentheses are written next to each other, or a number is written directly outside parentheses, multiplication is indicated.

Subtracting a positive number is equivalent to adding a negative number. Subtracting a negative number is equivalent to adding a positive number.

When multiplying numbers, multiplying numbers of opposite sign gives a negative product, and multiplying numbers of the same sign results in a positive product.

2. Dimensional Analysis

Dimensional analysis is a systematic way of solving numerical problems by the conversion of units. Frequently, your data will be recorded in one unit, but it will be necessary to do the calculation using a different unit of measurement. This means you should multiply the number you wish to convert by a conversion factor to produce a result in the desired unit. This way the units in the denominator cancel the units of the original data, leaving the desired units.

For example, if you pour a quarter liter of water from a full 500 mL container, how much remains in the container? To solve this problem, we need to determine how much water was poured out in milliliters (mL). We can use the conversion factor (1000 mL/1L).

0.25 L (1000 mL/1L) = 250 mL
500 mL - 250 mL = 250 mL left in the container

Question
A 200. cm^3 volume of liquid weighs 226 g. What volume of this liquid will weigh 5.0 g?

Solution
(200. cm^3/226 g)(5.0 g) = 4.4 cm^3

Question
What is the mass in kilograms of a 7.00 cm^3 piece of lead? The density of lead is 11.3 g/cm^3.

Solution
(7.00 cm^3 Pb)(11.3 g/cm^3 Pb)(1 kg/1000 g) = 0.0791 kg Pb

3. Quadratic Equation

The most likely place you will encounter quadratic equations is in the chapter on chemical equilibria, during the second semester of General Chemistry. Solving quadratic equations may seem overwhelming at first, but if you take one step at a time you will quickly see that it requires no more algebra then you have already learned.

A quadratic equation with one variable, x, is in the form:

$$ax^2 + bx + c = 0 \qquad a \neq 0$$

The coefficients a, b, and c may be either positive or negative numbers. If no coefficient is written, assume it to be one. The two roots of the equation can be found using the quadratic formula:

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

When you have solved a quadratic expression, you should always check your values by substitution into the original equation.

How do you know which of the two answers is the correct one? You have to decide in each case which root has physical significance. In chemistry, it is often the case that the negative value is not significant.

4. Logarithms

The logarithm of a positive number N to a given base b (written $\log_b N$) is the exponent of the power to which b must be raised to produce N. The most common log is base 10, usually written log N. In calculus, the most useful system of logarithms is the natural system in which the base is an irrational number symbolized by the letter e.

Despite the different bases of the two logarithms, they are used in the same manner. Our discussion will focus on common logarithms. A common logarithm is the power to which you must raise ten to obtain the number, $\log 10^x = x$. For example, the log of 100 is 2, since you must raise 10 to the second power to obtain 100.

To obtain the common log of a number other than a simple power of 10, you will need a calculator. For example, log 5.15 = 0.7118, which means that $10^{0.7118} = 5.15$.

Operations involving logarithms follow the same rules as those for exponents.

General laws of logarithms

- The log of the product of two or more positive numbers is equal to the sum of the logs of several numbers. For example, log (a x b)= log a + log b.
- The log of the quotient of two positive numbers is equal to the log of the dividend minus the log of the divisor. For example, log a/b = log a - log b.
- The log of a power of a positive number is equal to the log of the number, multiplied by the exponent of the power. For example, log a^n = n log a.
- The log of a root of a positive number is equal to the log of the number, divided by the index of the root. For example, log $\sqrt[n]{a}$ = (log a)/n
- The log of 1 = 0, log of numbers >1 are positive
- You can't have logs of negative numbers or of zero.

5. Significant Figures

The precision of a measurement indicates how well several determinations of the same quantity agree. In the laboratory, chemists attempt to set up experiments so that the greatest possible accuracy can be achieved. For each individual experiment, several measurements are usually made and their precision determined. Usually, better precision is taken as an indication of better experimental work. A calculated result can be no more precise than the least precise piece of information that went into the calculation. This is why the rules of significant figures are used.

Rule 1

To determine the number of significant figures in a number, read the number from left to right and count all the digits, starting with the first digit that is not zero. If the last digit of a number does not contain a decimal point, then the number of significant figures is equal to the number of non-zero digits in the number. Zeros to the left of 1 only locate the decimal point. This is clearer when it is written in scientific notation.

Rule 2

When adding or subtracting, the number of decimal places in the answer should be equal to the number of decimal places in the number with the fewest places.

Rule 3

When multiplying or dividing, the number of significant figures in the answer should be the same as the number with the fewest significant figures.

Rule 4

When a number is rounded off (the number of significant figures is reduced), the last digit is increased by 1 only if the following digit is 5 or greater. When calculating, you should do the calculation using all of the digits allowed by the calculator and round off only at the end of the problem. Rounding off in the middle of the problem can cause errors.

Question
How many significant figures are there in the numbers
(a) 57 (b) 62.9 (c) 20.000 (d) 0.003 (e) 0.0403 (f) 0.04030?

Solution
(a) Beginning with the 5 and counting gives two significant figures.
(b) Beginning count with the 6 gives three significant figures.
(c) All zeros here are significant. There are five significant figures.
(d) None of the zeros here are significant, giving one significant figure.
(e) The zeros to the left are insignificant, but the zero in the middle is a significant digit, giving three significant figures.
(f) The zero at the end and the zero in the middle are significant digits, giving four significant figures.

Problem
Perform the following calculations and express the answers to the proper number of significant figures:
(a) 14 - 0.052 (b) 32.1/3.21 (c) 13.79/0.0002 (d) (0.0801)(10)

Solution
(a) 14 (b) 10.0 (c) 70,000 (d) 0.801
(Note: the exact factor of 10 does not limit the number of significant figures in the answer.)

6. Percents and Fractions

Fractions
A fraction is a part of a whole. In its simplest form, the value of a fraction is less than one. An example commonly used to explain fractions is a pie: cut the pie into eight slices, then eat two. Two-eighths of the pie is gone. There are six pieces left, so six-eighths (6/8) of the pie is left.

a. Reducing Fractions

Imagine you have three pies cut into sections: one into fourths, one into eighths, and one into sixteenths. Take one piece from the first pie, two from the second, and four from the third. How much has been eaten from each pie? From pie one: one-fourth (1/4), from pie two: two eighths (2/8), and from pie three: four-sixteenths (4/16). The same amount, however, has been taken from each pie: one quarter (1/4). As you can see, it is frequently more convenient to talk about fractions in the most reduced, or simplest terms.

To reduce a fraction:
1. Find a number that divides evenly into the numerator and the denominator.
2. Check to see if another number goes in evenly. Repeat until the fraction is reduced as far as possible.

Example: 48/64
(48/8)/(64/8) = 6/8
(6/2)/(8/2) = 3/4

b. Converting Between Improper Fractions and Whole/ Mixed Numbers
An improper fraction is one in which the numerator is larger than the denominator. For example, if you were told you had six-fourths (6/4) of a pie left, you would know that you had one whole pie (4/4) plus one-half of a pie (2/4). It can be more useful to express the fraction as a mixed number, a number containing a whole number and a fraction. Thus rather than 6/4 of a pie, you have 1 1/2 pies.

To change an improper fraction into a mixed number:

1. Divide the denominator into the numerator.
2. Write the remainder as a fraction over the original denominator.
3. Reduce the remaining fraction.

c. Changing a Mixed Number into an Improper Fraction

When carrying out mathematical operations, it is usually necessary to work with improper fractions rather than mixed numbers. To change a mixed number into an improper fraction:

1. Multiply the whole number by the denominator.
2. Add the numerator to the product.
3. Place the sum in the numerator, over the original denominator.

d. Multiplying and Dividing Fractions

To multiply fractions, simply multiply the numerators together and multiply the denominators together. When multiplying fractions, you can cross reduce: reduce as you normally would a fraction, but use the numerator of one fraction and the denominator of the other. This will save you from having to reduce the fraction product. When multiplying a fraction by a whole number, write the whole number as a fraction over one, and multiply as usual.

Example: (15/7) x (3/5)

1. divide both 15 and 5 by 5
2. (3/7) x (3/1) = 9/7

To divide by a fraction, invert the fraction to the right of the division sign and multiply. Example: 1/2 / 1/8

1. (1/2)(8/1) = 8/2 = 4

e. Adding and Subtracting Fractions

Many problems will require mathematic manipulations of fractions. To continue with the pie example, if you have 3/8 of one pie and 1/4 of another pie, how much pie do you have?

To add or subtract fractions, they must have the same denominator. Then you simply add the numerators. Thus, we have 3/8 of one pie and 2/8 of another, leaving us with 5/8 of a pie.

If the fractions you want to add do not have a common denominator, you must change one or both of the fractions. You can do so by multiplying one or more of the fractions by an expression equivalent to one.

f. Decimals

In one of our earlier examples, we determined that a quarter of the pie had been eaten. We expressed one quarter as a fraction, 1/4. It can also be expressed as a decimal, 0.25, which is read as twenty-five hundredths. Any fraction can be expressed as a decimal by dividing the numerator by the denominator.

g. Rounding Off

When a fraction's denominator does not divide evenly into the numerator, the decimal equivalent can be long and too cumbersome to work with. For example, 1/3 = 0.3333333... The decimal is infinite, so we will want to use an abbreviated version for our records and calculations. An acceptable equivalent, rounded to two decimal places, is 0.33. When rounding off, increase the last digit retained by one if the following digit is greater than or equal to 5. Leave the last digit unchanged if the following digit is less than 5.

Percents

The percent symbol (%) means per hundred. 15% is equivalent to 15/100 or 0.15. Thus, when a quarter of our pie was eaten, 0.25, or 25% was gone. Any percentage may be expressed in decimal form by dividing by 100 and dropping the percent symbol. For example, 52.3% = 52.3/100 = 0.523. To calculate percentage, you must change the percent to decimal form (divide the percent by 100) and multiply.

CH 221 Chapter One Concept Guide

Much scientific work depends on physical measurements, and much of the progress of science depends on communication. Throughout time, a system of standardized units of measurement has evolved. The system utilizes the metric system and a standard set of units, called *Système International d'Unités* (International System of Units), abbreviated SI units.

1. Common Units

SI System
Most measurements in chemistry are made in terms of powers of ten of these standard units.

<u>Units of Measurement</u>

Measurement	Name of Unit	Abbreviation
Mass	kilogram *	kg
Length	meter *	m
Time	second *	s
Temperature	Kelvin *	K
Amount of substance	mole *	mol
Charge	coulomb	C
Electric current	ampere *	A
Force	Newton	N
Work and energy	joule	J
Frequency	hertz	Hz
Pressure	Pascal	Pa
Volume	cubic meter *	m^3
Volume	liter	L

* indicates SI Base Unit

Selected Prefixes Used in the Metric System

mega-	M	10^6
kilo-	k	10^3
deci-	d	10^{-1}
centi-	c	10^{-2}
milli-	m	10^{-3}
micro-	μ	10^{-6}
nano-	n	10^{-9}
pico-	p	10^{-12}

2. Mass

Mass is a physical property that represents the quantity of matter in an object. Weight, on the other hand, is the force exerted on the object by the pull of gravity on the mass of that object and is expressed in units of force.

- 1 lb = 0.4536 kg
- 1 g = 0.0353 oz
- 1 metric ton = 10^3 kg

Problem
Make the following conversion: 16.0 pounds to grams.

Solution
We need the following conversion factor to convert pounds to grams:
0.4536 kg = 1 lb.
(16.0 lbs)(0.4536 kg/1 lb)(1000 g/1 kg) = 7260 g

3. Energy

In the SI system, the unit for energy of all types is the joule (J). One J = 1 kg * m^2/s^2, where m is meters and s is seconds. Both joules and kilojoules (kJ), as well as the older metric units of calorie and kilocalorie, are used, where 1 cal = 4.184 J, exactly.

4. Temperature

Heat is a form of energy that arises from the motions of atoms and molecules in a substance. Temperature, on the other hand, is a measurement that determines whether heat can transfer from one object to another, and it indicates the direction of that transfer. Three temperature scales are in common use today: Fahrenheit (F), Celsius (C), and Kelvin (K). When these scales measure the same temperature, they give different numbers. For example, 212 °F, 100 °C, and 373.15 K represent identical temperatures. The conversions between different scales are:

$$K = {}^\circ C + 273.15 = ({}^\circ F + 495.67)/1.8$$
$${}^\circ C = ({}^\circ F - 32)/1.8 = K - 273.15$$
$${}^\circ F = ({}^\circ C * 1.8) + 32 = (K * 1.8) - 459.67$$

Problem
Convert 70.0 °F to (a) Celsius and (b) Kelvin.

Solution
(a) We will need to substitute given information, T(°F) = 70.0 °F, into the following equation:
$${}^\circ C = (70.0{}^\circ F - 32)/1.8 = 21.1\ {}^\circ C$$
(b) Now, use the following equation to convert from Celsius to Kelvin.
$${}^\circ C = K - 273.15$$
$21.1{}^\circ C = K - 273.15 = 294.\ 25$ K which gives **294.3 K** with correct significant figures.

5. Density

When mass and volume are combined in a ratio, they yield density: mass per unit volume of substance. The most common SI unit for this property is g/cm^3, which is equivalent to g/mL. The density of solids and liquids generally has these dimensions, whereas it is common to use g/L for gases.

When we know a substance's density, we can convert between its mass and volume. For this reason, density is a common conversion factor.

Problem
Vinegar has a density of 1.0056 g/cm^3. What is the mass of 1.5000 L of vinegar?

Solution
$$(1.5000\text{ L vinegar})(1000\text{ mL}/1\text{ L})(1\text{ cm}^3/1\text{ mL})(1.0056\text{ g}/1\text{ cm}^3)$$
$$= 1.5084 \times 10^3\text{ g}$$

CH 221 Chapter Two Part 1 Concept Guide

1. Origins of Atomic Theory

The simple picture of atoms as tiny spheres in constant motion goes back a long way in history - to the Greek philosopher Levucippus and his student, Democritus (460 - 370 BC). Democritus reasoned that if a bit of matter was divided into smaller and smaller pieces, one would ultimately arrive at a tiny particle that could not be further divided. He described this particle with the word "atom," which means uncuttable.

Democritus used his concept of atoms to explain physical properties and actions, such as density, hardness, and evaporation. The theory was untestable and unsupported, however, and remained so for more than 2000 years.

By the time a useful theory of atoms was developed early in the 19th century, much progress had been made in science and chemistry. The existence of gases was recognized and widely studied. Some basic physical laws, such as the conservation of matter, had been elucidated. Mathematics was recognizes as a key to understanding the universe. The value of experimentation and measurement was established. Science had evolved into a form we can recognize today.

The most significant progress was made in a concentrated period near the end of the 19th and at the beginning of the 20th centuries. By the dawn of the 19th century, certain basic rules or scientific laws had been formulated and were widely accepted: the law of conservation of matter, the law of constant composition, and the law of multiple proportions. To be accepted, any theory of matter must satisfy these laws.

In 1803, John Dalton revived the idea of atoms as real objects having size and mass, and he developed the first useful atomic theory. He linked the existence of elements, which cannot be chemically decomposed, to the idea of atoms as invisible units.

Dalton's postulates were generally accepted at the time because they were useful and violated none of the established scientific laws.

Another step in understanding atomic structure was the discovery of radioactivity, which helped determine that atoms can in fact break down, and by implication, have a structure we can analyze. Marie Curie (1867 - 1934) shared the Nobel Prize in Physics in 1903 with her husband, Pierre Curie (1859 - 1906) and Henry Becquerel (1852 - 1908) for discovering the phenomenon of "radioactivity," a process in which atoms of naturally radiating substances emit their unusual rays as they spontaneously disintegrate.

2. Electrons

The idea of electrons existed long before the particles were found to be a part of atoms. The English scientist Michael Faraday (1791-1867) performed many experiments with electricity, which led to the concept of a fundamental particle of electricity (then believed to be distinct from atoms). The name "electron" was suggested in 1891, and in 1897, J. J. Thomson of Cambridge University in England proved the particle nature of electrons in an experiment using cathode rays.

Robert Andrews Millikan (1868-1953) followed Thomson, devising an elegant, now classic experiment to measure the charge of the electron using oil droplets with negative charges, now called the "Millikan Oil Drop Experiment." Millikan's measurements of charges on oil droplets led to the discovery of the charge on the electron. Since the charge-to-mass ratio of the electron was known, its mass could be calculated. The currently accepted value for the mass is 9.109389×10^{-28} g, and that for the charge is $1.60217733 \times 10^{-19}$ C, where C, the Coulomb, is the SI unit of electric charge.

3. Protons

It was a student of Thomson's, Ernest Rutherford, who, in 1914, discovered the proton. Rutherford focused on the particles of canal rays, those that flow in the opposite direction to cathode rays in gas-discharge tubes. He found that the lightest and simplest canal-ray particles are formed when a gas-discharge tube contains hydrogen. Eventually he showed that particles identical to hydrogen atoms with one electron missing, H^+, are present in all matter. He named these particles protons, a fundamental subatomic particle with a positive charge equal in magnitude to the negative charge of the electron.

4. Neutrons

Because they have no charge, neutrons are much harder to detect than protons and electrons, and the search for the neutron lasted many years. It began in 1920 with Rutherford's idea that the nucleus might contain an uncharged particle with a mass close to that of the hydrogen atom. James Chadwick, a member of Rutherford's laboratory in Cambridge, saw the answer in alpha-particle experiments reported by other workers, in which a "highly penetrating" radiation from beryllium knocked protons out of paraffin with great force. Chadwick believed this radiation was a beam of uncharged particles, each with the mass expected for a neutron. After performing his own experiments, Chadwick proved conclusively the existence of the neutron in 1932.

5. The Modern View of an Atom

The existence of protons and electrons led to what became known as the Thomson model of the atom. Thomson proposed that an atom consists of a positively charged, uniform sphere of relatively large volume and low density, in which negatively charged electrons are embedded. A single experiment, however, destroyed this model.

Rutherford's now famous experiment involved bombarding a thin gold foil with positively charged alpha particles.

If Thomson's model were correct, the positively charged alpha particles would have plunged through the atoms of the foil like bullets through jelly. Instead, some of the particles were deflected backward. Obviously, they were hitting something very dense and very small. Rutherford and his group had discovered the nucleus of the atom.

6. Basics of Matter

All objects have physical properties, such as color, boiling point, melting point, magnetism, and viscosity. Physical properties are observed and measured without changing the composition of a substance.

a. Atoms, Molecules, Elements, and Compounds

All matter is constructed of building blocks called atoms. These particles are the simplest form of matter and the smallest that retain the chemical properties of that element.

Some pure substances, such as carbon, oxygen, and hydrogen, are composed of only one type of atom and are classified as elements. Atoms can link together to form bigger building blocks, or molecules. When two or more different kind of atoms combine, the combination is called a chemical compound. More specifically, the definition of a compound is limited to "pure" substances, the atoms of which are in a fixed ratio. For example, water is a compound that has two hydrogen atoms for each oxygen atom.

A molecule is the smallest entity that retains the chemical properties of the compound. Each molecule has a definite number of atoms, represented by the compound's formula. Letters in a chemical formula are symbols for the element; numbers used as subscripts indicate the number of atoms of that element in a single molecule.

Problem
Name the element represented by each of the following symbols:
(a) As (b) Be (c) B (d) V (e) Tl

Solution
(a) Arsenic (b) Beryllium (c) Boron (d) Vanadium (e) Thallium

Problem
Write the symbol for each of the following elements:
(a) xenon (b) magnesium (c) cobalt (d) copper (e) lead (f) silver (g) gold.

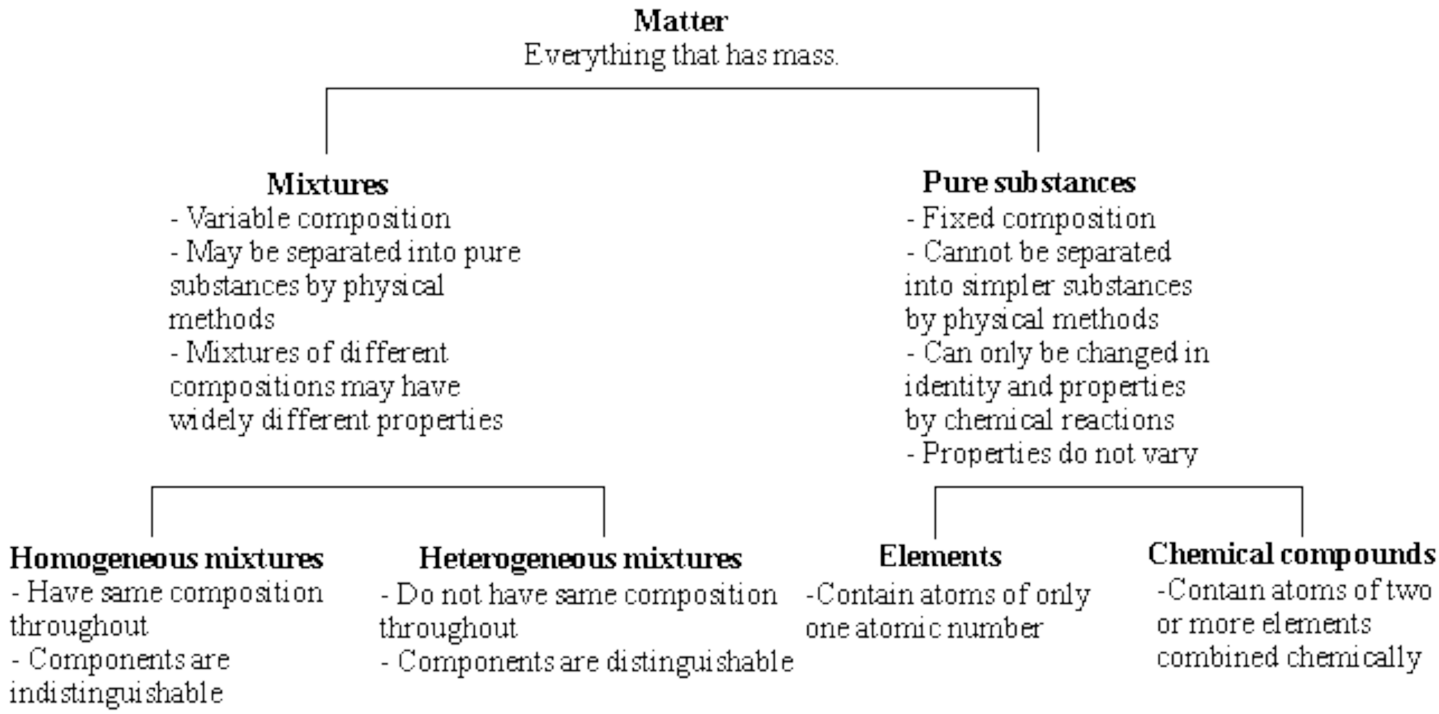
Solution
(a) Xe (b) Mg (c) Co (d) Cu (e) Pb (f) Ag (g) Au

b. States of Matter

An easily observed property of matter is its physical state, or phase. Almost all substances are solids, liquids, or gases, and virtually all matter is found in the solid state at low temperatures. At higher temperatures, solids generally melt to form liquids. Further heating may cause liquids to evaporate to form gases.

c. Mixtures and Pure Substances

All matter can be classified as a pure substance or a mixture by examining its properties and composition. Typically, there is further classification of mixtures into a homogeneous mixture or heterogeneous mixture, and of pure substances into elements and chemical compounds. For example, air is a homogeneous mixture, whereas dirt is a heterogeneous mixture because the components remain physically separate and can be seen as separate components. The general classification of matter is summarized in the diagram below.



Reference: Bailar, JC; Moeller T; Kleinberg J; Guss CO; Castellion ME; Metz C. Chemistry, 3rd ed. San Diego: Harcourt Brace Jovanovich, 1989.

Problem: Using the "Classification of Matter" terminology given in this lesson, describe each of the following and identify the number of phases in each sample:

(a) a drop of mercury (b) an ice cube (c) a melting ice cube (d) a puddle of water

Solution

(a) matter, pure substance, element; 1 phase
(b) matter, pure substance, compound; 1 phase
(c) matter, heterogeneous mixture; 2 phases
(d) matter, pure substance, compound; 1 phase

7. Atomic Mass

Question

Atoms frequently gain or lose electrons during chemical processes. Does this substantially affect their mass?

Approach

Select an element, add or subtract electrons, and calculate its new mass. Compare this new mass to the mass of the element.

Solution

As an example, let's look at an iron atom with a nucleus containing 26 protons and 30 neutrons having a total mass of 9.28836×10^{-23} g. If iron loses 2 electrons (mass of electron = 9.109389×10^{-28} g) to form Fe^{2+}, its mass is 9.28818×10^{-23} g.

$$9.28836 \times 10^{-23} - (2)(9.109389 \times 10^{-28}\text{ g}) = 9.28818 \times 10^{-23}\text{ g}$$

Based on this calculation, it is clear that the loss of electrons does not significantly change the mass of the element. The mass of Fe and Fe^{2+} is approximately the same, with a difference of only 1.8×10^{-27} g, or 0.002%.

8. Radioactivity and Atomic Composition

Radioactivity is the spontaneous emission of electromagnetic radiation and/or unstable nuclei of particles. Nuclides that spontaneously break down, or *decay*, are called radioisotopes, radionuclides, or radioactive nuclides.

Three different types of radiation from specific elements are common, although others are possible: alpha particles, beta particles, and gamma rays. Alpha decay is the emission of an alpha particle by a radionuclide. When this occurs, the mass decreases by 4 units and the atomic number decreases by 2. For example, uranium-238 decays to thorium-234.

$$^{238}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{234}_{90}\mathrm{Th}$$

An alpha particle is a helium nucleus of higher energy.

Beta decay is the emission of high-energy electrons that have been created by the decay of neutrons within the nucleus. These high-energy electrons are more penetrating of alpha particles.

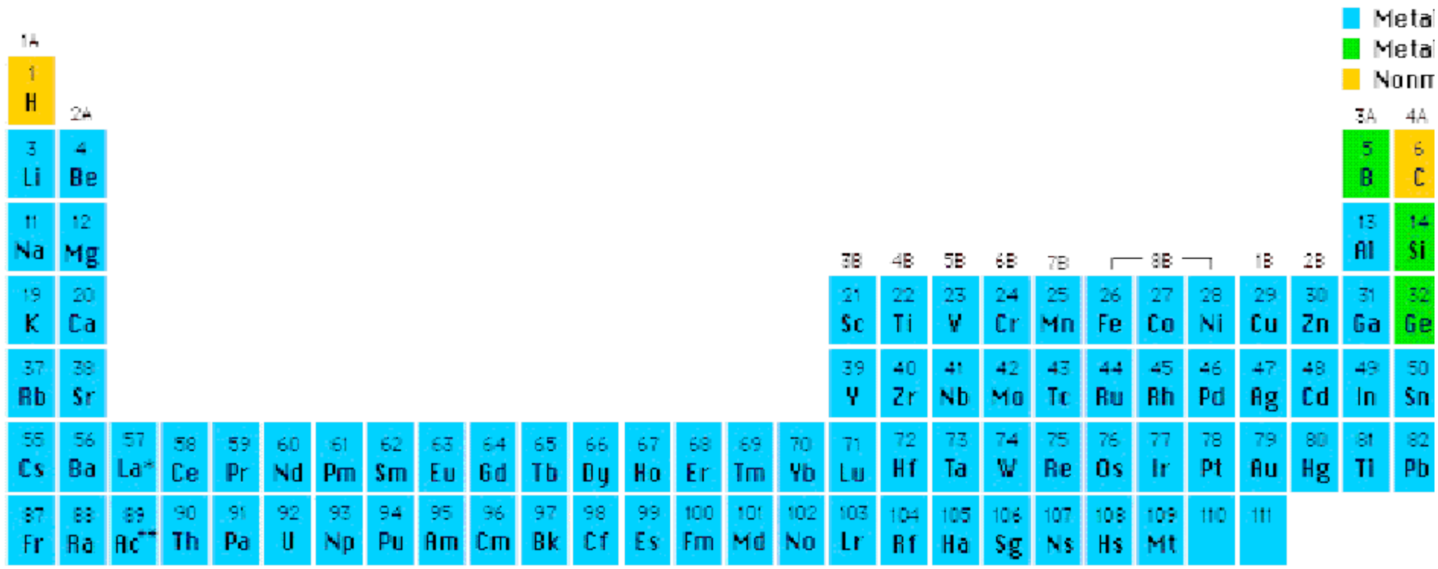
When beta decay occurs, the atomic number increases by 1, yet there is no change in atomic mass. Essentially, a neutron in the nucleus is converted into a proton and an electron is ejected. For example, carbon-14, an isotope commonly used in a technique to determine age, is radioactive and decays to nitrogen-14 by beta emission.

$$^{14}_{6}\mathrm{C} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{14}_{7}\mathrm{N}$$

9. Examining the Periodic Table

Question
What would the periodic table look like if the Lanthanide and Actinide series were not in a separate section?

Solution



10. Element Characteristics

Problem
List five attributes of chlorine.

Approach
Consider where chlorine is in the periodic table.

Solution
Chlorine is:

1. in Group 7A	2. a halogen
3. a non-metal	4. in Period 3
5. a Main Group element	6. a gas
7. very reactive	8. an element that reacts violently with alkali metals to form salts

11. Element Characteristics

Problem
List five attributes of sodium.

Approach
Consider where sodium is on the periodic table.

Solution
Sodium is:

1. an alkali metal	2. in Group 1A
3. a Main Group element	4. in Period 3
5. a solid	6. a conductor of electricity
7. malleable and ductile	8. reactive with water to produce hydrogen gas and an alkaline solution (NaOH)
9. not found free in nature	10. always found combined with other elements as compounds

12. Mendeleev's Periodic Table

Question
Examine Mendeleev's periodic table (circa 1872).

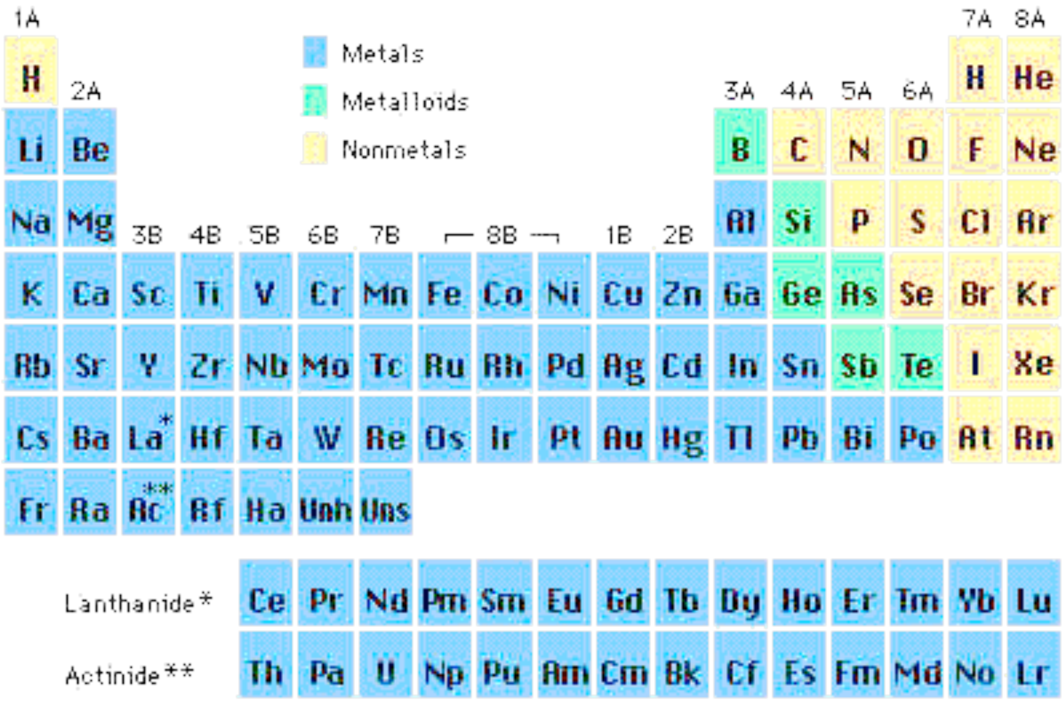
TABELLE II

Reihen	Gruppe I. — R^2O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	— = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63.
5	(Cu = 63)	Zn = 65	— = 68	— = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	— = 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108.
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	J = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er = 178	?La = 180	Ta = 182	W = 184	—	Os = 195, Ir = 197, Pt = 198, Au = 199.
11	(Au = 199)	Hg = 200	Tl = 204	Pb = 207	Bi = 208	—	—	
12	—	—	—	Th = 231	—	U = 240	—	— — — —

To what elements on the modern periodic table do the missing elements having masses of 68 and 70 grams per mole correspond? How do these predicted masses on Mendeleev's periodic table compare to those on the modern periodic table?

Solution

On Mendeleev's periodic table, the element having a mass of 68 refers to Gallium and that having a mass of 70 refers to Germanium. This is plausible because these elements fall between zinc and arsenic, as they do on Mendeleev's periodic table. In addition, the mass of Gallium is 69.72 grams per mole, which is only a difference of +1.72 grams per mole compared to Mendeleev's estimated 68 grams per mole. With regard to Germanium, the mass of this element is 72.59 grams per mole, a difference of only 2.59 grams per mole from Mendeleev's approximation. This is remarkable considering that Mendeleev did not observe these elements, but predicted their existence and properties by inference of other known elements and the structure of his table.


1A
7A
8A
Metals
Metalloids
Nonmetals
2A
3A
4A
5A
6A
3B
4B
5B
6B
7B
8B
1B
2B
H
H He
Li Be
B C N O F Ne
Na Mg
Al Si P S Cl Ar
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe
Cs Ba La* Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn
Fr Ra Ac** Rf Ha Unh Uns
Lanthanide* Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu
Actinide** Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr

CH 221 Chapter Two Part 2 Concept Guide

1. Ion Charges

Question
What charge are the following ions expected to have?
(a) ionic barium (b) ionic oxygen (c) ionic potassium

Solution
(a) Barium is expected to form cations. Elements in periodic Group 2A form ions of +2 charge, therefore barium is expected to form Ba^{2+}.
(b) Oxygen is expected to form anions. It is in periodic Group 6A, and forms O^{2-}.
(c) Potassium is expected to form cations. Elements in periodic Group 1A form ions of +1 charge, therefore potassium will form K^{+}.

2. Ion Charge and Empirical Formula

Problem
Aluminum acts as a metal and oxygen acts as a nonmetal when they react. Predict the empirical formula for aluminum oxide.

Approach
Aluminum loses 3 valence electrons to form Al^{3+}, whereas oxygen gains 2 valence electrons to form O^{2-}.

Solution:
When ionic compounds form from elements, the total charge on the cations must balance out the total negative charge on the anions. We will need two Al ions, which give a +6 charge, for every 3 oxygen ions, which give a -6 charge. The empirical formula is: Al_2O_3.

3. Electrostatic Forces

Question
Which compound's ions are held together by stronger forces: LiBr or MgS?

Solution
The most significant difference between the two compounds is the charges on the individual ions. Li and Br have +1 and -1 charges, respectively. Mg and S have +2 and -2 charges, respectively. The higher charges on Mg and S (+2 and -2, in relation to +1 and -1 in LiBr) lead to stronger electrostatic forces.

4. Ionic Compounds

Problem
Give the number and identify the constituent ions in the following ionic compounds:
(a) NaF (b) $CaCl_2$ (c) $Cu(NO_3)_2$ (d) $NaCH_3CO_2$.

Solution

(a) 1 Na^+ and 1 F^- ion (b) 1 Ca^{2+} ion and 2 Cl^- ions

(c) 1 Cu^{2+} ion and 2 NO_3^- ions (d) 1 Na^+ and 1 $CH_3CO_2^-$ ion

5. Nomenclature

Problem

Give the formula for each of the following ionic compounds:

(a) ammonium nitrate (b) cobalt(II) sulfate (c) nickel(II) cyanide.

Solution

(a) NH_4NO_3 (b) $CoSO_4$ (c) $Ni(CN)_2$

6. Nomenclature

Problem

Name the following ionic compounds:

(a) Li_2CO_3 (b) $KHSO_3$ (c) $CuCl$ and $CuCl_2$.

Solution

(a) Lithium carbonate (b) Potassium hydrogen sulfite

(c) Copper(I) chloride and copper(II) chloride

7. Nomenclature

Question

What are the names of each of these molecules?

(a) CO_2 (b) S_2F_{10} (c) BF_3

Solution

The symbol of the cation is always given first, followed by the anion symbol. The correct names for the above molecules are:

(a) carbon dioxide (b) disulfur decafluoride (c) boron trifluoride

8. Nomenclature

Problem

Give the name for each of the following compounds:

(a) PI_3 (b) SCl_2 (c) XeO_3

Solution

(a) Phosphorus triiodide (b) Sulfur dichloride (c) Xenon trioxide

9. Naming Hydrogen-containing Compounds

Question
What is the name of HBr?

Solution
Hydrogen monobromide or hydrobromic acid

10. Nomenclature

Problem
Give the formula for each of the following compounds:
(a) zinc(II) carbonate (b) sodium phosphate (c) aluminum chloride.

Solution
(a) $ZnCO_3$ (b) Na_3PO_4 (c) $AlCl_3$

11. Molar Mass

Question
What is the molar mass of NaOH?

Approach
To find the molar mass of a compound, we must list the elements in the compound's formula and determine the number of atoms of each element in the formula. Then, for each element, we need to look up the molar mass. Once we have found the mass contributed by each element for one mole of the compound, the molar mass is calculated by adding these individual masses.

Solution
1 mol Na in NaOH = (1)22.98977 g/mol = 22.98977 g
1 mol O in NaOH = (1)15.9994 g/mol = 15.9994 g
1 mol H in NaOH = (1)1.0079 g/mol = 1.0079 g

Molar mass of 1 mol NaOH = 22.98977 g + 15.9994 g + 1.0079 g = 39.9971 g/mol NaOH

12. Molar Mass

Question
What is the molar mass of $Cu(NO_3)_2$?

Approach
To find the molar mass of a compound, we must list the elements in the compound's formula and determine the number of atoms of each element in the formula. Then, for each element, we need to look up the molar mass. Once we have found the mass contributed by each element for one mole of the compound, the molar mass is calculated by adding these individual masses.

Solution

1 mol Cu in $Cu(NO_3)_2$ = (1)63.546 g/mol = 63.546 g
2 mol N in $Cu(NO_3)_2$ = (2)14.0067 g/mol = 28.0134 g
6 mol O in $Cu(NO_3)_2$ = (6)15.9994 g/mol = 95.9964 g

Molar mass of 1 mol $Cu(NO_3)_2$ = 63.546 g + 28.0134 g + 95.9964 g = 187.556 g/mol of $Cu(NO_3)_2$

13. Converting Mass to Moles

Question
What quantity, in moles, does 107 g HBr represent?

Solution
The molar mass of HBr is 80.912 g/mol. The number of moles of HBr is:
107 g HBr * 1 mol HBr/80.912 g HBr = 1.32 mol HBr

14. Molar Mass and Moles

Question
Which represents a greater number of moles: 4.5 g of carbon dioxide or 4.5 g of sodium chloride?

Solution
The molar masses are 44.01 g for CO_2 and 58.44 g for NaCl. The numbers of moles of each compound is calculated using the molar mass:
4.5 g CO_2 * 1mol CO_2/44.01 g CO_2 = 0.10 mol CO_2
4.5 g NaCl * 1 mol NaCl/58.44 g NaCl = 0.077 mol NaCl

There are more moles of CO_2 than of NaCl.

15. Atomic Mass

Question
What is the average atomic mass of chlorine?

Mass of ^{35}Cl	34.96885 amu
Mass of ^{37}Cl	36.96712 amu
Isotopic abundance of ^{35}Cl	75.77 %
Isotopic abundance of ^{37}Cl	24.23 %

Approach
Consider the mass and abundance of the isotopes of chlorine.

Solution:
Step 1. There are 2 naturally occurring isotopes of chlorine: ^{35}Cl and ^{37}Cl. A sample of chlorine shows that the 2 isotopes are not present in equal amounts. The percent abundance is calculated by dividing the number of atoms of a given isotope by the total number of atoms of all isotopes of that element, times one hundred.

$$\text{Percent abundance} = \frac{\text{number of atoms of a given isotope}}{\text{total number of atoms of that element}} \times 100\%$$

Step 2. The average atomic mass is

$$\text{Atomic mass} = \sum_{\text{all isotopes}} \text{isotopic mass x fractional abundance}$$

For chlorine, the atomic mass is

Atomic mass = 34.96885 amu * 0.7577 + 36.96712 amu * 0.2423 = 35.45 amu

16. Converting Mass to Atoms

Question
The 1989 nutritional recommended dietary allowance (RDA) of iron for a female age 19-24 is 15 mg. How many iron atoms is this?

Approach
Mass must be converted to moles, then moles must be converted to atoms. Avogadro's number (6.022×10^{23}) will be needed, as will the molar mass of iron. Note: in molar calculations, all masses must be converted to grams.

Solution:
Follow the series of multiplication steps below to convert mass to atoms.

$$(15 \text{ mg Fe})\left(\frac{1 \text{ g}}{1000 \text{ mg}}\right)\left(\frac{1 \text{ mol Fe}}{55.847 \text{ g Fe}}\right)\left(\frac{6.022\text{x}10^{23}}{1 \text{ mol Fe}}\right) = 1.6 \times 10^{20} \text{ Fe atoms}$$

In 15 mg of iron there are 1.6×10^{20} Fe atoms.

17. Converting Mass to Molecules and Atoms

Question
How many ozone molecules and how many oxygen atoms are contained in 48.00 g ozone, O_3?

Approach
First, write the equation for the synthesis of ozone. Then, to calculate the number of ozone molecules in 48.00 g ozone, convert mass to molecules using the molecular weight of ozone and Avogadro's number (6.022×10^{23}). Next, to calculate the number of oxygen atoms, start with the mass of ozone, convert this to grams, use a mole-to-mole ratio to convert ozone to O_2, and finally, use Avogadro's number to obtain the number of oxygen atoms.

Solution:
Step 1. Write the equation for the synthesis of ozone, and calculate its mass.

$3\ O_2(g) \rightarrow 2\ O_3(g)$ $\quad\quad$ O_3 = 3(15.999 g/mol) = 47.997 g/mol

Step 2. To calculate the number of ozone molecules from its mass, multiply grams of ozone, molecular weight of ozone, and Avogadro's number.

$$48.00\text{g } O_3 \cdot \frac{1 \text{ mol } O_3}{49.997\text{g } O_3} \cdot \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol } O_3} = 6.022 \times 10^{23} \text{ molecules } O_3$$

The number of molecules of ozone in 48.00 grams is 6.022×10^{23} molecules.

Step 3. To calculate the number of atoms of oxygen, multiply three atoms of oxygen per molecule of ozone and the number of molecules of ozone.

$$6.022 \times 10^{23} \text{ molecules } O_3 \cdot \frac{3 \text{ atoms O}}{1 \text{ molecule } O_3} = 1.807 \times 10^{24} \text{ atoms O}$$

The number of oxygen atoms in 48.00 g of ozone is 1.807×10^{24} atoms.

18. Empirical Formula

Question
A 1.27 g sample of an oxide contains 0.55 g phosphorus and 0.72 g oxygen. What this oxide's empirical formula?

Approach
First, it is necessary to determine from the experimental data the number of moles of atoms of each element present. The simplest ratio is then found by dividing the numbers of moles of each element by the number of moles of the element present in the smallest amount.

Solution
Step 1. Calculate the number of moles of phosphorus and oxygen.
(0.55 g P)(1 mol P/30.97 g P) = 0.018 mol P
(0.72 g O)(1 mol O/16.00 g O) = 0.045 mol O

Step 2. Divide the numbers of moles by the number of moles of the element present in the smallest amount: P.
0.018 mol P/0.018 mol P = 1.0
0.045 mol O/0.018 mol P = 2.5

Thus, the empirical formula is $P_{1.0}O_{2.5}$.

Step 3. Double all numbers in the formula to convert the fraction to a whole number.
$2(P_{1.0}O_{2.5}) = P_2O_5$

The empirical formula is P_2O_5.

19. Percent Composition

Question
A 2.91 g sample of potassium metal when burned in oxygen formed a compound weighing 6.11 g and containing only potassium and oxygen. What is the percent composition of each element in this compound?

Approach
The percent composition is the percent by mass of each element in the compound, which is given by the mass of that element divided by the total mass of the compound, times 100.

Solution
The percent potassium in this compound is:
%K = (2.91 g / 6.11 g compound)(100%) = 47.6 %

The percent oxygen in this compound is:
6.11 g compound - 2.91 g K = 3.20 g O
%O = (3.20 g / 6.11 g compound)(100%) = 52.4 %

The percent composition of the compound is 47.6% potassium and 52.4% oxygen.

CH 221 Chapter Four Part I Concept Guide

1. Balancing Chemical Equations

Description
When chlorine gas, Cl_2, is added to solid phosphorus, P_4, a reaction occurs to produce liquid phosphorus trichloride, PCl_3, and heat.

$P_4(s) + 6\ Cl_2(g) \rightarrow 4\ PCl_3(l)$

Question
If you want to make 150.0 grams of PCl_3, how many moles of chlorine gas must you begin with?

Approach
The balanced chemical equation relates the number of moles of each species involved in the reaction. If we can determine the number of moles of PCl_3 we want to make, we can use the stoichiometric coefficients in the balanced equation to determine the quantity in moles of each reactant that is necessary.

Solution
Step 1. To determine the quantity of PCl_3 in units of moles, we must make use of the compound's molar mass, which is 137.33 g/mol. We convert units using the molar mass as a unit conversion factor, remembering to place the unit we are converting to in the numerator of the ratio.

Convert PCl_3 from grams to moles.

$(150.0\ g\ PCl_3)\ (1\ mol\ PCl_3/137.33\ g\ PCl_3) = 1.092\ mol\ PCl_3$

Step 2. We can determine the quantity in moles of Cl_2 needed to make 1.092 mol of PCl_3 using the fact that 6 mol of Cl_2 are used to form 4 mol of PCl_3. These are used in the form of a ratio to convert from one to the other.

Determine the number of moles of Cl_2 needed.

$(1.092\ mol\ PCl_3)(6\ molCl_2/4\ mol\ PCl_3) = 1.638\ mol\ Cl_2$

We must begin with 1.638 moles of Cl_2 gas to make 150.0 grams of PCl_3.

2. Balancing Chemical Equations

Description
Butane reacts with oxygen gas to produce carbon dioxide and water vapor.

$C_4H_{10}(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Question
What is the balanced form of this reaction equation?

Approach
Balance the numbers of elements on each side of the equation one compound at a time.

Solution

Step 1. Start with carbon. There are four carbon atoms on the left side of the arrow, so we need to put a 4 in front of the CO_2 on the right.

$C_4H_{10}(g) + O_2(g) \rightarrow 4\ CO_2(g) + H_2O(g)$

Step 2. Increase the number of hydrogen atoms on the right to match the number found on the left.

$C_4H_{10}(g) + O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)$

Step 3. Oxygen is the last element to be balanced. There are a total of thirteen oxygen atoms on the right side of the equation. To balance this, we must put 13/2 in front of the O_2 on the left.

$C_4H_{10}(g) + 13/2\ O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)$

Step 4. Although the equation is now balanced, we are not finished. Because we are looking at this reaction at the molecular scale, we cannot talk about half molecules. Therefore, the whole reaction equation must be doubled.

$2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$

3. Balancing Chemical Equations

Description

Xenon Tetrafluoride gas and water react to give xenon, oxygen, and hydrogen fluoride gases.

Question

What is the balanced form of this reaction?

Approach

Write out the reaction. Then balance the numbers of atoms on each side of the equation one element at a time.

Solution

Step 1. Write out the reaction. Indicate the physical state of each reactant and product.

$XeF_4(g) + H_2O(l) \rightarrow Xe(g) + O_2(g) + HF(g)$

Step 2. It is best to start with an element that appears in only one species on each side of the equation. Start by writing a coefficient of 4 for HF, thus obtaining 4 fluorine atoms on each side.

$XeF_4(g) + H_2O(l) \rightarrow Xe(g) + O_2(g) + 4\ HF(g)$

Step 3. Now consider the xenon atoms. There is one xenon atom on each side, therefore, the xenon atoms are balanced.

$XeF_4(g) + H_2O(l) \rightarrow Xe(g) + O_2(g) + HF(g)$

Step 4. There are 2 hydrogen atoms on the left side of the reaction, and 4 hydrogen atoms on the right side. To obtain 4 hydrogen atoms on the left, write a coefficient of 2 for H_2O.

$XeF_4(g) + 2\ H_2O(l) \rightarrow Xe(g) + O_2(g) + 4\ HF(g)$

Step 5. Finally, consider oxygen, the last element to be balanced. There are now 2 oxygen atoms on the left, and 2 oxygen atoms on the right, thus the oxygen atoms are balanced as is. This is the final balanced equation for the reaction of xenon tetrafluoride gas and water to give xenon, oxygen, and hydrogen fluoride gases.

$XeF_4(g) + 2\ H_2O(l) \rightarrow Xe(g) + O_2(g) + 4\ HF(g)$

4. Mole Ratios

Problem
Calcium carbide is produced by the reaction of calcium oxide with carbon at high temperatures:

$CaO(s) + 3\ C(s) \rightarrow CaC_2(s) + CO(g)$

What are the mole ratios that give
(a) the amount of calcium carbide produced by each mole of calcium oxide that reacts,
(b) the amount of carbon required by each mole of calcium oxide that reacts, and
(c) the amount of calcium carbide produced by each mole of carbon that reacts.

Approach
Use the balanced chemical equation to determine each mole ratio.

Solution
(a) 1 mol CaC_2 / 1 mol CaO
(b) 3 mol C / 1 mol CaO
(c) 1 mol CaC_2 / 3 mol C

5. Mole Ratios

Problem
How many moles of silicon dioxide would be required to produce 4.5 mol of P_4?

$2\ Ca_3(PO_4)_2(s) + 6\ SiO_2(s) + 10\ C(s) \rightarrow 6\ CaSiO_3(l) + 10\ CO(g) + P_4(g)$

Solution
(4.5 mol P_4)(6 mol SiO_2 / 1 mol P_4) = 27 mol SiO_2

6. Stoichiometry

Question
What mass of hydrogen fluoride can be produced by the reaction of 15.0 g of calcium fluoride with excess sulfuric acid?

$CaF_2(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + 2\ HF(aq)$

Approach
Convert known masses to moles. Use mole ratios to find the unknown in number of moles. Finally, convert from moles to the desired quantity, which in this case is mass.

Solution
Step 1. Convert the mass of calcium fluoride into moles using molecular mass.

(15.0 g CaF_2)(1 mol CaF_2 / 78.1 g CaF_2) = 0.192 mol CaF_2

Step 2. Multiply the moles of calcium fluoride times the mole ratio of calcium fluoride to hydrogen fluoride.

(0.192 mol CaF_2)(2 mol HF / 1 mol CaF_2) = 0.384 mol HF

Step 3. Now, multiply the moles of hydrogen fluoride times its molecular weight to convert from moles HF to grams HF.

(0.384 mol HF)(20.0 g HF / 1 mol HF) = 7.68 g HF

7.68 g hydrogen fluoride can be produced using 15.0 g calcium fluoride and excess sulfuric acid.

7. Limiting Reactants

Question

Suppose 378 g of CO are mixed with 60.0 g of H_2 to form CH_3OH. Which is the limiting reactant?

Approach

Write the balanced equation. Then, convert known quantities to number of moles. Use mole ratios to find the number of moles of CO required for all the H_2 to react, and the number of moles of H_2 required for all the CO to react.

Solution

Step 1. Write the balanced chemical equation.

$CO(g) + 2\ H_2(g) \rightarrow CH_3OH(l)$

Step 2. Convert mass of CO to moles of CO, and mass of H_2 to moles of H_2.

(378 g CO)(1 mol CO / 28.01 g CO) = 13.5 mol CO

(60.0 g H_2)(1 mol H_2 / 2.02 g H_2) = 29.7 mol H_2

Step 3. Use two mole ratios to determine (a) the number of moles of CO necessary for all the H_2 to react, and (b) the number of moles of H_2 for all the CO to react.

(13.5 mol CO)(2 mol H_2 / 1 mol CO) = 27.0 mol H_2

(29.7 mol H_2)(1 mol CO / 2 mol H_2) = 14.9 mol CO

The amount of CO available, 13.5 mol, is less than the required amount, 14.9 mol. The amount of H_2 available, however, is sufficient. Therefore, CO is the limiting reactant.

8. Percent Yield

Problem

When 50.0 g of calcium carbide reacted with an excess of water, 14.2 g of ethyne (acetylene) were produced. What is the percent yield of ethyne for the reaction?

$CaC_2(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(aq) + C_2H_2(g)$

Approach

First, calculate the theoretical yield, the stoichiometric amount that could be produced. Then, calculate the percent yield using the following equation:

% yield = actual yield / theoretical yield * (100%)

Solution

Step 1. Calculate the theoretical yield of ethyne.

(50.0 g CaC_2)(1 mol CaC_2 / 64.10 g CaC_2)(1 mol C_2H_2 / 1 mol CaC_2)(26.04 g C_2H_2 / 1 mol C_2H_2) = 20.3 g C_2H_2

Step 2. Divide the actual yield by the theoretical yield and multiply by 100 to calculate the percent yield.

% yield = (14.2 g ethyne / 20.3 g ethyne)(100) = 70.0%

9. Empirical Formula

Problem

0.225 g of magnesium burns in nitrogen to form 0.312 g of magnesium nitride. What is the empirical formula of magnesium nitride?

Approach

We would expect Mg^{2+} and N^{3-} ions to react to form Mg_3N_2, yet we need to determine if this is in fact correct given the experimental data. We need to calculate the number of moles of magnesium and nitrogen.

Solution

Step 1. Calculate the moles of magnesium.

(0.225 g Mg)(1 mol Mg / 24.32 g Mg) = 0.00925 mol Mg

Step 2. Calculate the mass of nitrogen.

0.312 g magnesium nitride - 0.225 Mg = 0.0870 g N

Step 3. The number of moles of nitrogen can be calculated from its mass.

(0.0870 g N)(1 mol N / 14.01 g N) = 0.00621 mol N

Step 4. Divide all moles by the smallest number of moles.

0.00925 mol Mg / 0.00621 mol N = 1.49 mol Mg / N

0.00621 mol N / 0.00621 mol N = 1.00 mol N

The mole ratio is, therefore, 1.49 mol Mg : 1.00 mol N.

The empirical formula for magnesium nitride is $Mg_{1.5}N_{1.}$ To obtain all whole numbers, multiply by a factor of 2: Mg_3N_2. This result suggests that the empirical formula of magnesium nitride is what was expected: Mg_3N_2.

10. Empirical Formula

Problem

Potassium dichromate contains three elements: potassium, chromium, and oxygen. A chemical analysis of a sample of potassium chromate resulted in 13.3 g K, 17.7 g Cr, and 19.0 g O. Calculate the empirical formula for potassium dichromate.

Approach

Calculate the number of moles of potassium, chromate, and oxygen. Then, divide all moles by the smallest number of moles to calculate the mole ratio.

Solution

Step 1. Calculate the number of moles of potassium, chromium, and oxygen.

(13.3 g K)(1 mol K / 39.10 g K) = 0.340 mol K

(17.7 g Cr)(1 mol Cr / 52.00 g Cr) = 0.340 mol Cr

(19.0 g O)(1 mol O / 16.00 g O) = 1.19 mol O

Step 2. Divide all moles by the smallest number of moles.

0.340 mol K / 0.340 mol K = 1.00 mol K
0.340 mol Cr / 0.340 mol K = 1.00 mol Cr/K
1.19 mol O / 0.340 mol K = 3.50 mol O/K

The mole ratio is, therefore, 1.00 mol K:1.00 mol Cr: 3.50 mol O

Step 3. Multiply by a factor of 2 to convert all moles to whole numbers. This is the empirical formula for potassium dichromate: $K_2Cr_2O_7$.

CH 221 Chapter Four Part II Concept Guide

1. Solubility

Why are some compounds soluble and others insoluble? In solid potassium permanganate, $KMnO_4$, the potassium ions, which have a charge of +1, are attracted to permanganate ions, which have a charge of -1. There's an electrostatic force between them that locks them into the lattice structure. When $KMnO_4$ dissolves, ionic bonds are broken as water molecules surround the ions. The partial positive charges on water molecules are attracted to the negatively charged permanganate ion.

When an ionic compound dissolves, forces are broken between ions of the ionic compounds and between the water molecules.

Water molecules that surround, or hydrate, an ion are less able to interact with other water molecules. So when an ionic compound dissolves, two different electrostatic forces are broken: the forces between positive and negative ions of the ionic compound, and the forces between water molecules that break when a water molecule hydrates an ion.

Two kinds of forces also form when an ionic compound dissolves: forces between cations and water molecules, and forces between anions and water molecules.

The relative strengths of the forces being made and broken are one factor that determines solubility. If the forces between the ions in an ionic compound are fairly weak compared to the forces between the ions and the water molecules, then the compound should be fairly soluble because stronger forces occur when the compound dissolves. However, if the forces within the lattice are stronger than the forces between the ions and the water molecules, then a compound will tend to be insoluble.

The relative strength of forces is only one of several different aspects of solubility. Another relates to the degree of order or disorder involved in the dissolution process. Entry of ions into solution causes an ordering of water molecules and ions. This effect deters solubility. Solubility is difficult to predict because the force that causes an ion to be attracted to a water molecule also causes it to be attracted to ions of the opposite charge. An ion that is strongly attracted to its neighboring ions will also be strongly attracted to water molecules.

2. Electrolytes

Compounds that form ions in aqueous solution are called electrolytes. The name reflects the fact that aqueous solutions containing ions can conduct electricity through the movement of those ions. Ions can be formed through dissolution of ionic compounds or by other reactions, including reactions of neutral molecular compounds with water molecules.

A strong electrolyte is a compound that ionizes completely when it dissolves. That is, nearly all of the compound forms cations and anions. Aqueous solutions of strong electrolytes are good conductors of electricity.

- **Strong electrolyte:** a compound that breaks up completely to form ions in aqueous solution.

Hydrogen chloride, HCl, is an example of a strong electrolyte. When hydrogen chloride is in aqueous solution, the molecules break up into hydrogen cations and chloride anions. Each molecule of HCl is neutral. When it dissolves, chlorine gains an electron to become a chloride ion with a negative charge. The hydrogen atom has lost its bonding electron, so it is now a hydrogen ion, which is really just a proton. Because this dissociation process happens to almost every hydrogen chloride molecule in aqueous solution, hydrogen chloride is called a strong electrolyte. Weak electrolytes conduct electricity only moderately well because only a small percentage of the compound reacts to form ions in aqueous solution.

Acetic acid, the main dissolved component in vinegar, is a weak electrolyte. In water, acetic acid partially dissociates into a proton and a polyatomic acetate anion. But the molecule doesn't remain dissociated for long. The acetate anion eventually links up with another proton. At the same time, other acetic acid molecules are dissociating into ions and then re-forming. About 0.42 percent of the acetic acid molecules in vinegar are dissociated.

A nonelectrolyte forms no ions and therefore cannot conduct an electric charge.

- **Nonelectrolyte:** a compound that dissolves but does not ionize in aqueous solution.

For example, the compound ethanol, which is the alcohol in alcoholic beverages, is a nonelectrolyte. Ethanol dissolves easily in water because, like water, it is a polar molecule.

The -OH group has a partial negative charge on the oxygen atom and a partial positive charge on the hydrogen atom. When ethanol dissolves in water, water molecules surround the polar ethanol molecules just as they surround ions, but aqueous ethanol is not ionic, and so cannot conduct electricity.

- **Examples of Electrolytes**

Strong Electrolytes	Weak Electrolytes	Nonelectrolytes
NaCl	HF	H_2O
$MgBr_2$	HNO_2	CH_3CH_2OH (ethanol)
HCl	NH_3	CH_3COCH_3 (acetone)
NaOH	CH_3CO_2H (acetic acid)	$C_2H_4(OH)_2$

3. Net Ionic Equations

Problem
When aqueous solutions of aluminum nitrate and sodium carbonate are mixed, a precipitate of aluminum carbonate forms. Write the net ionic equation for the reaction between aluminum nitrate and sodium carbonate.

Approach
The general approach to writing net ionic equations is to write the complete, balanced equation, then write the equation in terms of the individual ions that are in solution. Finally, eliminate any spectator ions, taking care to cross out an equal number of ions on each side of the equation.

Solution
Step 1. Write out the formulas of the reactants and determine the products.

$$Al(NO_3)_3(aq) + Na_2CO_3(aq) \rightarrow Al_2(CO_3)_3(s) + NaNO_3(aq)$$

Step 2. Balance the equation.

$$2\ Al(NO_3)_3(aq) + 3\ Na_2CO_3(aq) \rightarrow Al_2(CO_3)_3(s) + 6\ NaNO_3(aq)$$

Step 3. Write out the ions in solution for the aqueous compounds.

$$2\ Al^{3+}(aq) + 6\ NO_3^-(aq) + 6\ Na^+(aq) + 3\ CO_3^{2-}(aq) \rightarrow$$
$$Al_2(CO_3)_3(s) + 6\ NO_3^-(aq) + 6\ Na^+(aq)$$

Step 4. Cross out equal numbers of spectator ions-ions that appear on both sides of the equation that do not participate in the reaction:

$$2\ Al^{3+}(aq) + \cancel{6\ NO_3^-(aq)} + \cancel{6\ Na^+(aq)} + 3\ CO_3^{2-}(aq) \rightarrow$$
$$Al_2(CO_3)_3(s) + \cancel{6\ NO_3^-(aq)} + \cancel{6\ Na^+(aq)}$$

Step 5. The balanced net ionic equation is:

$$2\ Al^{3+}(aq) + 3\ CO_3^{2-}(aq) \rightarrow Al_2(CO_3)_3(s)$$

4. Reaction of Iron (III) Nitrate and Sodium Hydroxide

Question

Iron (III) nitrate reacts with sodium hydroxide to make iron (III) hydroxide and sodium nitrate. Does this reaction form a precipitate?

$$Fe(NO_3)_3(aq) + 3\ NaOH(aq) \rightarrow Fe(OH)_3(?) + 3\ NaNO_3(?)$$

Solution

Notice that we are considering the compounds that we'd have if the pairs of cations and anions were exchanged. We know that the original compounds are soluble, so an insoluble product would have to be either iron ions paired with hydroxide ions or sodium ions paired with nitrate ions.

If either of the products is insoluble, the reaction will produce a precipitate. Sodium salts are soluble, and nitrate salts are soluble. Therefore, sodium nitrate is soluble, and will not precipitate. All hydroxide ionic compounds are insoluble, with the exception of the alkali metal cations. Iron is not an alkali metal, so we expect iron (III) hydroxide to be insoluble. Thus, we expect a precipitate to form.

$$Fe(NO_3)_3(aq) + 3\ NaOH(aq) \rightarrow Fe(OH)_3(s) + 3\ NaNO_3(aq)$$

When the reaction is conducted, a brown precipitate of iron (III) hydroxide forms. Our prediction was correct. The net ionic equation is

$$Fe^{3+}(aq) + 3\ OH^-(aq) \rightarrow Fe(OH)_3(s)$$

5. Synthesis of Iron (II) Carbonate.

Problem

Propose a reaction that yields iron (II) carbonate as a product. This compound is insoluble.

Solution

In our desired product, iron (II) carbonate, the iron (II) ion serves as the cation and the carbonate ion is the anion. The reactants, therefore, must include an iron (II) cation and a carbonate anion, which when reacted, produce this insoluble compound. The other product must be soluble in order for pure iron (II) carbonate to be obtained.

Step 1. Start by writing the product.

$$FeCO_3(s)$$

Step 2. Write the net ionic equation formation of this product.

$Fe^{2+}(aq) + CO_3^{2-}(aq) \rightarrow FeCO_3(s)$

Step 3. Choose a soluble compound that will release Fe^{2+} ions when dissolved, and another compound that will release carbonate ion. Choices here are iron (II) nitrate and sodium carbonate, because all nitrate and sodium salts are soluble. (Note: there are several other compounds that would work here. Sodium and nitrate salts are just one possibility.)

$Fe(NO_3)_2(aq) + Na_2CO_3(aq)$

Step 4. Write out the balanced equation and indicate the physical state of each reactant and product. This is a final reaction for the synthesis of iron (II) carbonate using iron (II) nitrate and sodium carbonate.

$Fe(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow FeCO_3(s) + 2\ NaNO_3(aq)$

6. Acid-Base Reactivity: Determining Net Ionic Equations

Question

What is the net ionic equation for the acid-base reaction between HNO_3 and NaF?

Solution

Step 1. Write out the complete reaction.

$HNO_3(aq) + NaF(aq) \rightarrow HF(aq) + NaNO_3(aq)$

Step 2. Write out the reaction in terms of ions in solution.

$H^+(aq) + NO_3^-(aq) + Na^+(aq) + F^-(aq) \rightarrow HF(aq) + Na^+(aq) + NO_3^-(aq)$

HF is a weak acid, so it exists in solution mainly in the undissociated form. We therefore write it as HF instead of dissociated ions.

Step 3 . Cross out spectator ions.

$H^+(aq) + \cancel{NO_3}^-\ (aq) + \cancel{Na}^+(aq) + F^-(aq) \rightarrow HF(aq) + \cancel{Na}^+(aq) + \cancel{NO_3}^-(aq)$

Step 4. Write the net ionic equation.

$H^+(aq) + F^-(aq) \rightarrow HF(aq)$

7. Polyprotic Acids

Question

What are the three ionization steps for the polyprotic acid, H_3PO_4, reacting with H_2O?

Approach

The ionizations of polyprotic acids occur stepwise, one at a time. Phosphoric acid has three acidic protons, thus there are three ionization steps.

Solution

Step 1. Write out the reaction for phosphoric acid and water.

$H_3PO_4(aq) + H_2O(aq) \rightarrow H_3O^+(aq) + H_2PO_4^-(aq)$

Step 2. Now, write the reaction for second ionization step, i.e., the removal of a total of 2 protons from phosphoric acid. Use $H_2PO_4^-$ from the first step and react with water.

$H_2PO_4^-(aq) + H_2O(aq) \rightarrow H_3O^+(aq) + HPO_4^{2-}(aq)$

Step 3. Finally, the third ionization step is the reaction of HPO_4^{2-} with water. This reaction yields the tribasic form of phosphate. There are no additional ionization steps because the product of this step, PO_4^{3-}, does not have any protons.

$HPO_4^{2-}(aq) + H_2O(aq) \rightarrow H_3O^+(aq) + PO_4^{3-}(aq)$

8. Acids Without Hydrogen Atoms

Question

What is the reaction for the synthesis of HNO_3?

Approach

Oxides of the non-metals are usually acids. Nonmetal oxides react with water molecules to increase the acidity of the solution. The reaction, therefore, should be that of NO_2 and H_2O. $H_2(g)$ is also generated through the oxidation of NO_2.

Solution

Step 1. Write the synthesis reaction of HNO_3 using NO_2 and H_2O as reactants.

$NO_2(aq) + H_2O(aq) \rightarrow HNO_3(aq) + {}^1/_2\, H_2(aq)$

9. Acid-base Synthesis Reaction

Question

What is an appropriate example of an acid-base reaction for the synthesis of K_2SO_4?

Solution

Step 1. Choose a strong acid to give SO_4^{2-} ions.

$H_2SO_4(aq)$

Step 2. Choose a strong base to give K^+ ions.

$KOH(aq)$

Step 3. Write the balanced reaction for the synthesis for the dissolved salt, K_2SO_4.

$H_2SO_4(aq) + 2\ KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(aq)$

Step 4. The net ionic equation for the reaction is:

$H^+(aq) + OH^-(aq) \rightarrow H_2O(aq)$

10. Acids without Hydrogen Atoms

An aqueous solution of dissolved carbon dioxide gas is acidic: it has an H^+ concentration greater than that of pure water. How can CO_2 act as an acid if it has no hydrogen to donate?

Carbon dioxide reacts with the hydroxide ions in solution. When CO_2 and OH^- come together, the oxygen of the OH^- bonds to carbon, forming hydrogen carbonate ion, also known as bicarbonate ion. This leaves an excess of hydronium ions in solution, which agrees with the Arrhenius definition of an acid.

When dissolved in water, oxides of the nonmetals yield acids, and are sometimes called acid anhydrides. Nonmetal oxides can also react with intact water molecules, rather than with free hydroxide ions, to increase the acidity of the solution. For example, when coal is burned, sulfur compounds can be released into the atmosphere. There they react to form SO_3. which then reacts with water to make sulfuric acid:

$$SO_3(aq) + H_2O(l) \rightarrow H_2SO_4\ (aq)$$

The reaction takes place in droplets of water in the atmosphere, and the product, sulfuric acid, falls to Earth dissolved in raindrops. Burning coal without trapping the sulfur generated during combustion can increase the acidity of raindrops.

Although most oxides of non-metals yield acidic aqueous solutions through reactions such as those of CO_2 and SO_3, these compounds are not usually listed in tables of acids**.** It's important to recognize that they act as acids in aqueous solution.

11. Metal Oxides are Bases

We know that oxides of nonmetal elements act as acids. Oxides of the metallic elements act as bases and are sometimes called basic anhydrides. A metal oxide can react in water to make a basic solution.

For example, solid calcium oxide reacts with water to form calcium hydroxide:

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$$

Some of the product, $Ca(OH)_2$, dissociates, adding hydroxide ions to solution:

$$Ca(OH)_2(s) \rightarrow Ca^{2+}\ (aq) + 2\ OH^-\ (aq)$$

12. Redox Reaction of NaCl

Question
When solid sodium and gaseous chlorine are combined in a flask, a vigorous reaction occurs to produce sodium chloride. Which species is oxidized and which is reduced?

Approach
To answer this question, we must examine the oxidation numbers of the elements before and after the reaction.

Solution
The charges on both sodium metal and chlorine gas are zero. Once the reaction occurs, and sodium chloride is produced, the charge of sodium is plus one, and the charge of chlorine is minus one.

	$2\ Na(s)$ +	$Cl_2(g) \rightarrow$	$2\ NaCl(s)$
charge per atom =	0	0	+1 -1

Na loses $1e^-$ per atom. Na is oxidized.
Cl gains $1e^-$ per Cl atom. Cl is reduced.

The sodium atom becomes more positively charged, because it loses one electron to each chlorine atom. Sodium is oxidized. Each chlorine atom gains one electron from the sodium atom and becomes more negatively charged. Chlorine is reduced.

Chlorine is the oxidizing agent because it is the substance that gains or accepts electrons from sodium. The sodium metal is the reducing agent because it is losing or donating electrons to chlorine.

13. Oxidation Numbers

Question
Which species is oxidized and which species is reduced in the combustion reaction of methane and oxygen?

Approach
Examine the oxidation numbers of the elements before and after the reaction.

Solution
Step 1. Write the balanced reaction of methane and oxygen, and indicate the physical state of each species.

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$

Step 2. Assign oxidation numbers to each element.

$CH_4(g)$ +	$2\ O_2(g)$	$\rightarrow$	$CO_2(g)$ +	$2\ H_2O(g)$
-4 +1	0		+4 -2	+1 -2

Step 3. Using oxidation numbers, determine which element gains electrons (reduction) and which element loses electrons (oxidation).

Oxidation state of carbon: $-4 \rightarrow +4$
Oxidation state of oxygen: $0 \rightarrow -2$

In the reaction of methane and oxygen, carbon is oxidized and oxygen is reduced. Hydrogen is neither oxidized nor reduced, as its oxidation number remains +1 after the reaction has occurred.

14. Oxidation Numbers

Question
Which species is oxidized and which species is reduced in the reaction of potassium nitrite, potassium iodide, and sulfuric acid?

Approach
Examine the oxidation numbers of the elements before and after the reaction.

Solution
Step 1. Write the balanced reaction of potassium nitrite, potassium iodide, and sulfuric acid, and indicate the physical state of each species.

$2\ KNO_2(aq) + 2\ KI(aq) + 2\ H_2SO_4(aq) \rightarrow$
$2\ NO(aq) + 2\ K_2SO_4(aq) + 2\ H_2O(aq) + I_2(aq)$

Step 2. Assign oxidation numbers to each element.

2 KNO_2(aq) +	2 KI(aq) +	2 H_2SO_4(aq)	→
+1 +3 -2	+1 -1	+1 +6 -2	
2 NO(aq) +	2 K_2SO_4(aq) +	2 H_2O(aq) +	I_2(aq)
+2 -2	+1 +6 -2	+1 -2	0

Step 3. Using oxidation numbers, determine which element gains electrons (reduction) and which element loses electrons (oxidation).

Oxidation state of iodine: -1 → 0
Oxidation state of nitrogen: +3 → +2

In the reaction of potassium nitrite, potassium iodide, and sulfuric acid, iodine is oxidized and nitrogen is reduced. Hydrogen, potassium, sulfur, and oxygen are neither oxidized nor reduced, as their oxidation numbers remain the same after the reaction has occurred. Moreover, SO_4^{2-} stays intact, thus there is no need to consider it in terms of being oxidized or reduced.

15. Oxidation Numbers

Question
Which species is oxidized and which species is reduced in the reaction of chlorine and hydrogen sulfide?

Approach
Examine the oxidation numbers of the elements before and after the reaction.

Solution
Step 1. Write the balanced reaction of chlorine and hydrogen sulfide, and indicate the physical state of each species.

8 Cl_2(g) + 8 H_2S(aq) → S_8(s) + 16 HCl(aq)

Step 2. Assign oxidation numbers to each element.

8 Cl_2(g) +	8 H_2S(aq)	→	S_8(s) +	16 HCl(aq)
0	+1-2		0	+1 -1

Step 3. Using oxidation numbers, determine which element gains electrons (reduction) and which element loses electrons (oxidation).

Oxidation state of chlorine: 0 → -1
Oxidation state of sulfur: -2 → 0

In the reaction of chlorine and hydrogen sulfide, sulfur is oxidized and chlorine is reduced. Hydrogen is neither oxidized nor reduced, as its oxidation number remains the same after the reaction has occurred.

16. Molarity

Problem
A student prepared a solution by dissolving 1.455 g of potassium nitrate in enough water to make 20.00 mL of solution. Calculate the molarity of the solution.

Approach
Molarity is calculated by dividing the number of moles of solute by the volume of solution (in liters).

Solution
Step 1. Calculate the number of moles of solute, KNO_3.
Number of moles of KNO_3 = (1.455 g KNO_3)(1 mol / 101.11 g KNO_3) = 0.01439 moles KNO_3

Step 2. Calculate the molarity of the solution.
Molarity = 0.01439 moles KNO_3 / 0.02000 L solution = 0.7195 M KNO_3

17. Molarity

Question
How many moles of NaCl are present in 15.00 mL of a 1.60 M NaCl?

Approach
To calculate the number of moles, multiply the molarity of the solution with the volume of solution (in liters).

Solution
Number of moles NaCl = 1.60 mol/L NaCl * 0.01500 L solution = 0.0240 mol NaCl

18. Molarity

Question
What volume of a 1.35×10^{-3} M $C_6H_{12}O_6$ solution should you transfer to obtain a solution that contains 2.44×10^{-6} moles of glucose?

Approach
Use the molarity of the glucose solution to convert from units of moles to units of liters.

Solution
Volume of solution to transfer = 2.44×10^{-6} moles glucose / 1.35×10^{-3} M $C_6H_{12}O_6$
= 0.00181 L
= 1.81 mL of 1.35×10^{-3} M $C_6H_{12}O_6$

19. Dilution

Problem
Calculate the volume of 0.0111 M HCl that we should use to prepare 100 mL of a 4.23×10^{-4} M HCl solution.

Approach
Use the following equation to calculate the final volume of the solution. The other variables, $M_{initial}$, $V_{initial}$, and M_{final} are all known values.

$M_{initial} \times V_{initial} = M_{final} \times V_{final}$

Solution
$M_{initial} \times V_{initial} = M_{final} \times V_{final}$

$(0.0111 \text{ M HCl})(V_{initial}) = (4.23 \times 10^{-4} \text{ M HCl})(100 \text{ mL})$

$V_{initial} = 3.81 \text{ mL}$

3.81 mL of a 0.0111 M HCl(aq) solution are needed to prepare 100 mL of a 4.23×10^{-4} M HCl(aq) solution.

20. Titrations

Question
If 20.00 mL of a solution of oxalic acid, $H_2C_2O_4$, was titrated with 0.500 M NaOH(aq) and the endpoint was reached when 30.0 mL of the solution of base had been added, what is the molarity of the oxalic acid solution?

Approach
First, we need to write the balanced chemical equation. Calculate the number of moles of NaOH that were added. Then calculate the number of moles of $H_2C_2O_4$ consumed using the mole ratios indicated in the balanced equation.

Solution
Step 1. Write the balanced chemical equation.

$H_2C_2O_4(aq) + 2\ NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\ H_2O(l)$

Step 2. Calculate the number of moles of NaOH added.

0.03000 L NaOH x 0.500 mo/L NaOH = 0.0150 mol NaOH

Step 3. Calculate the number of moles of $H_2C_2O_4$.
The stoichiometric relation we need is 2 mol NaOH: 1 mol $H_2C_2O_4$.
Number of moles of $H_2C_2O_4$
= 0.0150 mol NaOH x (1 mol $H_2C_2O_4$ / 2 mol NaOH)
= 0.00750 mol $H_2C_2O_4$

Step 4. The molarity of the oxalic acid solution can be calculated by dividing the moles of $H_2C_2O_4$ by the volume of base added in liters.

0.00750 mol $H_2C_2O_4$ / 0.0200 L solution = 0.375 M $H_2C_2O_4$

21. Titrations

Question
A student prepared a sample of aqueous HCl that contained 0.45 g of HCl in 250 mL of solution. This solution was used to titrate 20.0 mL of a solution of $Ca(OH)_2$ and the equivalence point was reached when 14.4 mL of acid solution had been added. What was the molarity of the $Ca(OH)_2$ solution?

Approach
First, we need to write the balanced chemical equation. Then the molarity of the HCl solution must be calculated. From this value, the number of moles of HCl can be determined, and the number of moles of $Ca(OH)_2$ can be calculated using a mole ratio. Then, the molarity of the $Ca(OH)_2$ solution can be calculated.

Solution
Step 1. Write the balanced chemical equation.
$2\ HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) + 2\ H_2O(l)$

Step 2. Calculate the molarity of HCl.
(0.45 g HCl)(1 mol HCl / 36.46 g HCl) / 0.250 L solution = 0.049 mol/L HCl

Step 3. Calculate the number of moles of HCl.
0.049 mol/L HCl x 0.0144 L HCl = 7.1×10^{-4} mol HCl

Step 4. Calculate the number of moles of $Ca(OH)_2$.
The stoichiometric relation we need is 2 mol HCl : 1 mol $Ca(OH)_2$.
Number of moles of $Ca(OH)_2$
= 7.1×10^{-4} mol HCl * (1 mol $Ca(OH)_2$ / 2 mol HCl)
= 3.6×10^{-4} mol $Ca(OH)_2$.

Step 5. The molarity of the $Ca(OH)_2$ solution can be calculated by dividing the moles of $Ca(OH)_2$ by the volume of $Ca(OH)_2$ solution, in liters.
3.6×10^{-4} mol $Ca(OH)_2$ / 0.020 L $Ca(OH)_2$ = 0.018 mol/L $Ca(OH)_2$

22. Stoichiometry of Reacting Gases

Problem
Calculate the volume of sulfur dioxide produced at 25 ˚C and 1.00 atm by the combustion of 15.6 g of sulfur.
$S_8(s) + 8\ O_2(g) \rightarrow 8\ SO_2(g)$ *Note that* 1 mol of gas at 25 ˚C occupies 24.47 L

Approach
The first step is to convert to moles of S_8. Second, use a mole ratio to convert moles S_8 to SO_2. Then, use molar volume and the number of moles of SO_2 to calculate the volume of SO_2.

Solution
Step 1. Convert grams to moles of S_8.
(15.6 g S_8)(1 mol S_8 / 256.48 grams S_8) = 0.0608 mol S_8

Step 2. Use a mole ratio to convert from moles of sulfur to moles of sulfur dioxide.
(0.0608 mol S_8)(8 mol SO_2 / 1 mol S_8) = 0.486 mol SO_2

Step 3. Finally, use molar volume to convert moles of SO_2 to liters of SO_2.
0.486 mol SO_2 * 24.47 L/mol = 11.9 L SO_2

23. Limiting Reactants and Molarity

Question
What mass of hydrogen will be produced from the reaction of 8.0 g of zinc with 20 mL of 5.0 M hydrochloric acid?

Approach
Write the balanced equation. Then, convert known quantities to number of moles. Use mole ratios to find the unknown in number of moles. Finally, convert from moles to the desired quantity, which is in this case, mass.

Solution
Step 1. Write the balanced chemical equation.
$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Step 2. Convert mass of Zn to moles of Zn. Also, convert the molarity of HCl to moles of HCl.
(8.0 g Zn)(1 mol Zn / 65.38 g Zn) = 0.122 mol Zn
(20 mL)(1 L / 1000 mL)(5.0 mol HCl / 1 L) = 0.10 mol HCl

Step 3. Use two mole ratios to determine (a) the number of moles of HCl necessary for all the Zn to react, and (b) the number of moles of Zn for all the HCl to react.
(0.122 mol Zn)(2 mol HCl / 1 mol Zn) = 0.244 mol HCl
(0.10 mol HCl)(1 mol Zn / 2 mol HCl) = 0.050 mol Zn

The amount of HCl available, 0.10 mol, is less than the required amount, 0.244 mol HCl. The amount of Zn available, however, is sufficient. Therefore, HCl is the limiting reactant, and the amount of H_2 formed is determined by the amount of HCl:
(0.10 mol HCl)(1 mol H_2 / 2 mol HCl) = 0.050 mol H_2

Step 4. To calculate the mass of H_2 produced in this reaction, multiply the moles of H_2 times its molecular weight.
(0.050 mol H_2)(2.02 g H_2 / 1 mol H_2) = 0.10 g H_2

24. Synthesis Using Acid-Base Reactions

Compounds can be synthesized using acid-base reactions. Consider the compound calcium chloride, $CaCl_2$. Calcium chloride is soluble in water. Thus, given an aqueous $CaCl_2$ solution, we could isolate the $CaCl_2$ by evaporating the water.

When a strong acid and a strong base react, they form a dissolved salt and water. To form the salt calcium chloride, we need to react a strong acid that will donate a chloride ion to water with a strong base that will donate a calcium ion to water.

One calcium chloride producing reaction uses hydrochloric acid, HCl, for the acid and calcium hydroxide, $Ca(OH)_2$, for the base:
$Ca(OH)_2\ (aq) + 2HCl(aq) \rightarrow CaCl_2\ (aq) + 2H_2O(l)$

The water can be evaporated to isolate the pure calcium chloride compound.

Another reaction we could use to generate calcium chloride begins with HCl and calcium carbonate, $CaCO_3$, as reactants. Metal carbonates react with strong acids to generate a salt, water, and carbon dioxide:
$CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

The CO_2 gas escapes from solution, leaving water and calcium chloride. Again, evaporating the water could isolate the pure calcium chloride compound.

CH 221 Chapter Five Concept Guide

1. Specific Heat

Question
How much heat is needed to raise the temperature of 28 g of aluminum from 27 °C to 163 °C? No phase changes occur. The specific heat of aluminum is 0.902 J/K*g.

Approach
The amount of heat needed is found from the known mass of Al, the change in temperature, and the specific heat of Al. The following equation relates heat to these values: q = (mass)(specific heat)(temperature change).

Solution
ΔT = 163 °C - 27 °C = 136 °C, or 136 K.
q = (mass)(specific heat)(temperature change)
q = (28 g)(0.902 J/K*g)(136 K)
q = 3400 J

Note that the temperature increased and q is positive. Heat was added to the aluminum, so the direction of heat flow was into the system.

2. Specific Heat

Question
What quantity of heat is needed to raise the temperature of 130 g of water by 5.0 °C? The specific heat of water is 4.184 J/K*g.

Approach
The following equation relates heat to mass and temperature change:
q = (mass)(specific heat)(temperature change).

Solution
ΔT = 5.0 °C, or 5.0 K
q = (mass)(specific heat)(temperature change)
q = (130 g)(4.184 J/K*g)(5.0 K)
q = 2700 J

Note that the temperature increased and q is positive.

3. Molar Heat Capacity

Problem
125 J of heat was absorbed by 1.00 mol of iron, resulting in a temperature change of 4.02 °C. Calculate the molar heat capacity of iron.

Approach

Molar heat capacity can be calculated using the following equation:
q = (number of moles)(molar heat capacity)(temperature change).

Solution
q = (number of moles)(molar heat capacity)(temperature change).
125 J = (1.00 mol iron)(molar heat capacity of iron)(4.02 °C)
molar heat capacity of iron = 125 J / (1.00 mol iron)(4.02 °C)
molar heat capacity of iron = 31.1 J/mol*K

4. Molar Heat Capacity

Heat capacity values are tabulated in two ways: as molar heat capacity, and as specific heat capacity, which has already been discussed in this lesson. Molar heat capacity differs from specific heat capacity in that the former is the amount of heat required to raise the temperature of one mole of a substance by one Kelvin (J/K*g). You may recall that the specific heat is the amount of heat required to raise the temperature of one gram of a substance by one Kelvin (J/K*g).

The molar heat capacity may be calculated using the following equation:

q = (number of moles)(molar heat capacity)(temperature change)

where q is heat in joules. The molar heat capacity has units of J/K*g. For example, the specific heat of Cu is 0.385 J/K*g. Its molar heat capacity is, therefore, 24.5 J/K*mol.

5. Hess's Law: Heats of Reaction

Problem
Using the following thermochemical data, calculate $\Delta H°$ for:
$CaC_2(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(aq) + C_2H_2(g)$

Thermochemical Equation	$\Delta H°$ (kJ)
1. $Ca(s) + 2\ C(graphite) \rightarrow CaC_2(s)$	-59.8
2. $Ca(s) + {}^1/_2\ O_2(g) \rightarrow CaO(s)$	-635.09
3. $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$	-653.1
4. $C_2H_2(g) + 5/2\ O_2(g) \rightarrow 2\ CO_2(g) + H_2O(l)$	-1300.
5. $C(graphite) + O_2(g) \rightarrow CO_2(g)$	-393.509

Approach
We need to add a series of reactions for which $\Delta H°$ is known, to yield the overall reaction of interest. We can manipulate the known reactions as we do this. If necessary, reverse some of the known thermochemical

equations so that the major reactants and products of the desired equation are on the appropriate sides. Then, if necessary multiply the known equations by appropriate coefficients so the major reactants and products have the same coefficients as in the desired equation. Add the known equations and the values of ΔH° to obtain ΔH° for:

$CaC_2(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(aq) + C_2H_2(g)$

Solution

The first, third, and fourth equations include the major reactants and products, yet equations 1 and 4 must be reversed so that the reactants and products are on the appropriate sides according to the desired equation. Only multipliers are needed, as the coefficients are the same in the known equations as in the desired equation.

$CaC_2(s) \rightarrow \cancel{Ca(s)} + \cancel{2\ C(graphite)}$	ΔH° = 59.8 kJ
$\cancel{CaO(s)} + H_2O(l) \rightarrow Ca(OH)_2(aq)$	ΔH° = -653.1
$\cancel{2\ CO_2(g)} + H_2O(l) \rightarrow C_2H_2(g) + \cancel{5/2\ O_2(g)}$	ΔH° = 1300.
$\cancel{Ca(s)} + \cancel{^1/_2\ O_2(g)} \rightarrow \cancel{CaO(s)}$	ΔH° = -635.09
$\cancel{2\ C(graphite)} + \cancel{2\ O_2(g)} \rightarrow \cancel{2\ CO_2(g)}$	ΔH° = -787.018
$CaC_2(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(aq) + C_2H_2(g)$	ΔH° = -715 kJ

This reaction is exothermic and has a ΔH° = -715 kJ.

6. Hess's Law

Problem

The following data are the thermochemical equations for iron oxides. Find ΔH° for:

$4\ FeO(s) \rightarrow Fe(s) + Fe_3O_4(s)$

Thermochemical Equation	ΔH° (kJ)
1. $Fe(s) + {}^1/_2\ O_2(g) \rightarrow FeO(s)$	-272.0
2. $3\ Fe(s) + 2\ O_2(g) \rightarrow Fe_3O_4(s)$	-1118.4

Approach

We need to add a series of reactions for which ΔH° is known, to yield the overall reaction of interest. We can manipulate the known reactions as we do this. If necessary, reverse some of the known thermochemical equations so that the major reactants and products of the desired equation are on the appropriate sides. Then, if necessary multiply the known equations by appropriate coefficients so the major reactants and products have the same coefficients as in the desired equation. Add the known equations and the values of ΔH° to obtain ΔH° for:

$4\ FeO(s) \rightarrow Fe(s) + Fe_3O_4(s)$

Solution
Equation 1 must be reversed and multiplied by 4 so that the reactants and products are on the appropriate side and the coefficients are the same as in the desired equation.

$4\ Fe(s) + 2\ O_2(g) \rightarrow 4\ FeO(s)$ $\Delta H° = 1088.\ kJ$

$3\ Fe(s) + 2\ O_2(g) \rightarrow Fe_3O_4(s)$ $\Delta H° = -1118.4\ kJ$

$4\ FeO(s) \rightarrow Fe(s) + Fe_3O_4(s)$ $\Delta H° = -30.4\ kJ$

This reaction is exothermic and has $\Delta H° = -30.4$ kJ.

7. Heats of Reaction

Problem
Given that the standard heat of formation of PbO(s) at 298 K is -218.99 kJ/mol, write the thermochemical equation for the formation of PbO(s).

Approach
The equation must use the elements in their standard states at 298 K as reactants, and one mole of PbO(s) in its standard state as the product. In their standard states, lead is a solid and oxygen is gas.

Solution
The unbalanced reaction is:
$Pb(s) + O_2(g) \rightarrow PbO(s)$

Now, balance the equation by putting a coefficient of $^1/_2$ before O_2. Remember: we must form a single mole of PbO_2.
$Pb(s) + {}^1/_2\ O_2(g) \rightarrow PbO(s)$

The heat of reaction of PbO(s) is given as -218.99 kJ/mol, so the complete reaction is:
$Pb(s) + {}^1/_2\ O_2(g) \rightarrow PbO(s)$ $\Delta H° = -218.99\ kJ$

8. Calorimetry

Problem
A 40.00 mg sample of naphthalene, $C_{10}H_8$, was combusted in a calorimeter containing 750. g H_2O. The temperature of the calorimeter increased from 25.000 °C to 25.404 °C. The calorimeter has a heat capacity (calorimeter constant) of 465 J/°C. Calculate $\Delta H°_{combustion}$ for naphthalene in kJ/mol.

Approach
First, calculate the energy to heat the water, q_{water}. Then, calculate the energy to heat the calorimeter vessel, $q_{calorimter}$. The total energy transferred is the sum of these two numbers, q_{system}. Next, convert the mass of the naphthalene sample into moles. Finally, solve for the heat transfer using q_{system} and the moles of naphthalene.

Solution

Step 1. Solve for q_{water}.

$q(J)_{water}$ = 750. g * 4.184 J/g K * 0.404 K = 1.27×10^3 J

Step 2. Solve for $q_{calorimeter}$.

$q(J)_{calorimeter}$ = 465 J/K * 0.404 K = 1.88×10^2 J

Step 3. Solve for q_{system}, which is the negative of the total energy transferred.

Total energy transferred = $-q(J)_{system}$

= $-(1.27 \times 10^3$ J + 1.88×10^2 J) = -1.46 kJ

Step 4. Convert mass of naphthalene to moles naphthalene using the molar mass.

0.0400 g naphthalene * 1 mol naphtlanene/128.17 g naphthalene

= 3.12×10^{-4} mol naphthalene

Step 5. Finally, solve for heat transfer by dividing q_{system} in kilojoules by the moles of naphthalene.

$\Delta H°$ = -1.46 kJ/3.12×10^{-4} mol naphthalene = -4.66×10^3 kJ/mol

CH 221 Chapter Six Part I Concept Guide

1. Radiation, Wavelength, and Frequency

Question

Is the frequency of the radiation used in a microwave oven higher or lower than that from an FM radio station broadcasting at 91.7 MHz (where 1 MHz = 10^6 s^{-1})?

Solution

Microwave radiation has a frequency on the order of 10^{10} Hz, compared to FM radio, which has a frequency on the order of 10^8 Hz. FM radio is, therefore, lower in frequency than microwaves.

2. The Relationship between Wavelength and Frequency

Question

What is the wavelength of orange light of frequency 4.85 x 10^{14} Hz?

Approach

We need to convert frequency to wavelength using the following relation:

$\lambda = c / \nu$

where λ is wavelength in meters, c is the speed of light, and ν is the frequency in s^{-1}.

Solution

$\lambda = c / \nu$

$\lambda = 3.00 \times 10^8$ m/s / 4.85×10^{14} s^{-1}

$\lambda = 6.19 \times 10^{-7}$ m = 619 nm

3. Planck's Law

Problem

Compare the energy of a mole of photons of green light (5.00×10^2 nm) with the energy of a mole of photons of microwave radiation having a frequency of 2.45 GHz (1 GHz = 10^9 s^{-1}). Which has greater energy? By what factor is one greater than the other?

Approach

First, we calculate the frequency of radiation of green light. Next, we calculate the energy of the green light and the energy of the microwave radiation. A ratio of energies will result in the factor by which one is greater than the other.

Solution

Step 1: Calculate the frequency of green light.

$\nu = c/\lambda = 3.00 \times 10^8$ m s^{-1} / 5.00×10^{-7} m = 6.00×10^{14} s^{-1}

Step 2: Calculate the energies of green light and microwave radiation.

E(green light) = $h\nu$ = (6.626×10^{-34} J s/photon)(6.00×10^{14} s^{-1}) = 3.98×10^{-19} J/photon

E(microwave radiation) = $h\nu$ = (6.626×10^{-34} J s/photon)(2.45×10^9 s^{-1}) = 1.62×10^{-24} J/photon

Green light has greater energy than microwave radiation.

Step 3: Use a ratio of energy values to calculate the factor by which the energy of green light is greater than that of microwave radiation.

E(green light) / E(microwave radiation)

= 3.98×10^{-19} J/photon / 1.62×10^{-24} J/photon = 2.45×10^5

Green light is almost a quarter of a million times more energetic than microwaves.

4. Matter as Waves

Question

Does a particle exhibiting wavelike behavior have a frequency as well as a wavelength?

Solution

All matter exhibits wavelike behavior. Recall that for all waves, $\lambda\nu = c$, where λ is the wavelength, ν is the frequency, and c is the speed of light. For waves, c is replaced by v, which is the velocity of the wave: $\lambda\nu = v$.

Thus, $\lambda = v/\nu$, and $v/\nu = h/mv$. Finally, $\nu = mv^2/h$.

5. Calculating Uncertainty in the Position of an Electron

Question

What is the smallest possible uncertainty in the position of an electron having a mass of 9.109×10^{-21} kg and a velocity of $3.0 \times 10^{7} \pm 7.27 \times 10^{5}$ m/s? 1 Joule = 1 kg m^2/s^2.

Solution

The product of the uncertainty is momentum, $m\Delta v$, and the uncertainty in position, Δx, must be greater than h: $(m\Delta v)(\Delta x) > h$.

Therefore, $\Delta x > h/m\Delta v$.

$\Delta x > h/m\Delta v$

$\Delta x \sim (6.626 \times 10^{-34}\ \text{kg s m}^2/\text{s}^2) / (9.109 \times 10^{-21}\ \text{kg})(7.27 \times 10^{5}\ \text{m/s})$

$\Delta x \sim 10^{-19}$ m

6. Nodes

Question

The total number of nodes in an orbital is equal to the shell number, *n*, minus 1. These nodes are either nodal planes or nodal spheres. The number of nodal planes is equal to the value of *l* and the remainder are nodal spheres. What types of nodes exist in 3d orbitals and in 4d orbitals?

Approach

Find the total number of nodes from the shell number, the number of nodal planes for the value of *l*, and the number of nodal spheres by taking the difference.

Solution

A 3d orbital has $n = 3$, and a 4d orbital has $n = 4$, thus a 3d orbital has 2 nodes and a 4d orbital has 3 nodes. Both are d orbitals, therefore $l = 2$ and both 3d and 4d have 2 nodal planes. Finally, a 3d orbital has 2 nodes, of which 2 are nodal planes. This orbital has no nodal spheres. A 4d orbital, however, has 3 nodes, of which 2 are planes. A 4d orbital has 1 nodal sphere.

7. Quantum Numbers and Orbitals

Question

What values of the subshell quantum number correspond to the (a) d, (b) f, (c) s, (d) p, (e) g subshells?

Solution:

The first five subshells, l = 0, 1, 2, 3, 4, are identified by the letters s, p, d, f, and g.

(a) 2
(b) 3
(c) 0
(d) 1
(e) 4

CH 221 Chapter Six Part II Concept Guide

1. Writing Electron Configurations

Question
What is the complete electron configuration of the zirconium atom?

Approach
With increasing atomic number, electrons occupy the subshells available in each main energy level in order, with few exceptions:

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p.

Solution:
Forty electrons must be accommodated. The total electron configuration for Zr is:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^2\ 5s^2$.

This configuration may also be written as:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^2$.

2. Writing Electron Configurations

Question
What is the complete electron configuration for the arsenic atom?

Approach
With increasing atomic number, electrons occupy the subshells available in each main energy level in order, with few exceptions:

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p.

Solution:
Thirty-three electrons must be accommodated. The total electron configuration for As is:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^3$.

This configuration may also be written as:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$.

3. Writing Electron Configurations

Question
What is the noble gas notation for the electron configuration for the rubidium atom?

Approach
With increasing atomic number, electrons occupy the subshells available in each main energy level in specific order, with few exceptions:

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p.

The noble gas notation substitutes the symbol of the noble gas for the corresponding noble gas core in the electron configuration.

Solution:
The noble gas just prior to Rb in the periodic table is Kr. The noble gas notation for the electron configuration for Rb is: $[Kr]5s^1$.

4. Writing Electron Configurations

Question
What is the complete electron configuration for the Br^- ? With which element in the periodic table is it isoelectric?

Approach
With increasing atomic number, electrons occupy the subshells available in each main energy level in specific order:
1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p.

Solution:
The anion Br^- is formed by the addition of 1 electron to the lowest energy orbital that has a vacancy. The electron configuration for Br^- is:
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$.

Cl^- is isoelectric with krypton.

5. Electron Configurations of Transition Metal Cations

We will look at iron, a transition metal, and two of its cations, and compare their electron configurations. We will compare what we would expect from main group behavior against the real behavior of the cations. We'll examine the real behavior closely to understand why it occurs, and then come to some conclusions about why the real behavior differs from our expectations.

We will be examining Fe, Fe^{2+}, and Fe^{3+}. The electron configuration of neutral Fe is $[Ar]4s^2 3d^6$.

Question: What will be the electron configurations of Fe^{2+} and of Fe^{3+}?

Expectation (based on main group behavior): If Fe loses electrons to form cations, we might expect the electrons removed to be those that were added last. We would expect:

Fe^{2+}: $[Ar]4s^2 3d^4$ Fe^{3+}: $[Ar]4s^2 3d^3$

Real Behavior:

Fe^{2+}: $[Ar]3d^6$ Fe^{3+}: $[Ar]3d^5$

Elaboration: Most transition metal atoms have electron configurations with two electrons in an s orbital, and some electrons in the d subshell of the shell below (with a lower value of n) that of the s orbital.

When transition metal atoms lose electrons to form ions, the first electrons lost are those in the s orbital. The electron configuration of Fe^{2+} is $[Ar]3d^6$. The 4s electrons have been lost, the 3d electrons remain. If a third electron is lost, forming Fe^{3+}, it will be removed from the 3d subshell, forming an ion with an electron configuration $[Ar]3d^5$.

Explanation: In an atom of Ca, the 4s subshell fills in preference to the 3d subshell and the atom is lower in energy with an [Ar]$4s^2$ electron configuration. Having a filled s subshell is favored for transition metal atoms as well, but this does not hold true for transition metal cations. In these cases, it is energetically favorable to have the d orbital subshell fill in preference to the s subshell. This implies that orbital subshells shift both in absolute energy as well as in energy relative to one another when ions form. This is the case.

Additional problems:

1) Give the electron configurations for O, S, and Se in spectroscopic notation. In what way are they similar?

2) Give the noble gas notation electron configuration for samarium, Sm, which is in the lanthanide series. What subshell remains only partially filled? Is a samarium atom expected to be diamagnetic or paramagnetic?

3) Give three atoms or ions that have an electron configuration of: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$.

1. **Answer:**

   O: $1s^2 2s^2 2p^4$
   S: $1s^2 2s^2 2p^6 3s^2 3p^4$
   Se: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$

   They are similar in that they each have an outermost p subshell that contains four electrons. Each of the elements is located in group 4A of the periodic table. In general, elements that are found in the same periodic group have similar outermost electron configurations.

2. **Answer:**

   [Xe]$6s^2 4f^6$

   The 4f subshell is partially filled, as is expected for an f-block element. The 4f subshell has 6 electrons (the overall subshell, with seven orbitals, can hold 14 electrons). Sm has atoms among the most strongly paramagnetic of all the elements because its atoms have 6 unpaired electrons.

3. **Answer:** Notice that this electron configuration has an outermost 5p subshell that is filled. It therefore represents the noble gas element in period 5 of the periodic table. This is Xe.

   Atoms that can form anions with the same electron configuration will be found to the left of Xe in the periodic table. Therefore, I^- and Te^{2-} will each have this electron configuration. Atoms that can form cations with the same electron configuration will be the next two elements on the periodic table. Therefore, Cs^+ and Ba^{2+} will each have this electron configuration. These five species, Te^{2-}, I^-, Xe, Cs^+, and Ba^{2+} are all isoelectronic. They each have 54 electrons but, due to their differing number of protons, they have different charges.

6. Atomic Radii

Problem

Consulting the periodic table, decide whether the first atom in each of the following pairs is larger, smaller, or similar in atomic radius to the second atom:
(a) Si, Pb (b) Cs, Pb (c) Rh, Ir (d) Ti, V.

Approach
Atomic size increases moving down a group, and decreases moving left to right along a period.

Solution:
(a) Silicon is the second element in the carbon family, and lead is the fifth element. Silicon atoms should be smaller as size increases down a representative periodic group.
(b) Cesium and lead are in the same period, with lead further to the right in the period. Size decreases across the periods, and therefore cesium atoms should be larger than lead atoms.
(c) Rhodium and iridium are the second and third members of a d-transition metal group. Elements in the second and third transition series are very similar in size due to the lanthanide contraction, therefore, these atoms should be similar in size.
(d) Titanium and vanadium are adjacent elements in the same d-transition period. The decrease in size across the periods for transition elements is gradual. Atoms of these two elements should be similar in atomic radius.

7. Ionic Radii

Problem
Place the following species in order of increasing radius:
Cl Cl^- Cl^+

Approach
For species having the same number of protons, the more electrons, the larger the species.

Solution:
The positively charged chlorine atom is smaller than neutral chlorine; the former has one fewer electron than neutral chlorine. The negatively charge chlorine atom is the largest of all three; it has one additional electron than neutral chlorine.

$Cl^+ < Cl < Cl^-$

8. Ionic Radii

Problem
Arrange the following ions in order of increasing radius:
F^- Mg^{2+} Cl^- Be^{2+} S^{2-} Na^+

Approach
Ionic size increases moving down a group, and decreases moving left to right along a period. For isoelectric species, the greater the number of protons, the smaller the species.

Solution:
Mg^{2+}, Na^+, F^- all have the same number of electrons, thus they are similar in size. Mg^{2+}, however, has the largest number of protons, therefore it is the smallest of these three ions. Likewise, S^{2-} is larger than Cl^- because it has fewer protons. S^{2-} and Cl^- have an additional shell relative to Mg^{2+}, Na^+, and F^-, therefore S^{2-} and Cl^- are both larger than these three ions. Be^{2+} has one fewer shell, and is therefore smaller than Mg^{2+}, Na^+, and F^-.

$Be^{2+} < Mg^{2+} < Na^+ < F^- < Cl^- < S^{2-}$

CH 221 Practice Problem Set #1

*This is a **practice problem set** and not the actual graded problem set that you will turn in for credit.*
Answers to each problem can be found at the end of this assignment.

Covering: **Chapter One** *and* **Chapter Guide One**
Important Tables and/or Constants: 1 cm^3 = 1 mL, **273.15** *(temperature)*

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1. Give the name of each of the following elements: C, U, Tc, Si, Ne, Ir
2. Give the symbol for each of the following elements: Argon, Potassium, Radon, Lead, Bismuth, Helium
3. In each of the following pairs, decide which is an element and which is a compound:
    a. $Cu(NH_3)_6Cl_2$ and Cu
    b. zinc and zinc(II) sulfide
    c. carbon and methane
4. A chemist needs 2.00 g of a liquid compound with a density of 0.718 g/cm^3. What volume of the compound is required?
5. Make the following temperature conversions:
    a. 370. K to °C
    b. 16 °C to K
    c. 40 °C to K
6. A typical laboratory beaker has a volume of 250. mL. What is its volume in cubic centimeters? In liters? In cubic meters? In cubic decimeters?
7. Molecular distances are usually given in nanometers (1 nm = 1 x 10^{-9} m) or in picometers (1 pm = 1 x 10^{-12} m). However, the angstrom (Å) is sometimes used, where 1 Å = 1 x 10^{-10} m. If the distance between the Pt atom and the N atom in the cancer chemotherapy drug cisplatin is 1.97 Å, what is this distance in nanometers? In picometers?
8. The platinum-containing cancer drug cisplatin contains 65.0% platinum. If you have 1.53 g of the compound, what mass of platinum (in grams) is contained in this sample?
9. You can identify a metal by carefully determining its density (d). An unknown piece of metal, with a mass of 2.361 g, is 2.35 cm long, 1.34 cm wide, and 1.05 mm thick. Which of the following is this element?
    a. Nickel, d = 8.90 g/cm^3
    b. Titanium, d = 4.50 g/cm^3
    c. Zinc, d = 7.13 g/cm^3
    d. Tin, d = 7.23 g/cm^3
10. Carbon tetrachloride, CCl_4, a liquid compound, has a density of 1.58 g/cm^3. If you place a piece of a plastic soda bottle (d = 1.37 g/cm^3) and a piece of aluminum (d = 2.70 g/cm^3) in liquid CCl_4, will the plastic and aluminum float or sink?
11. Give the number of significant figures in each of the following numbers:
    a. 0.546 s
    b. 1583.3 mL
    c. 9.10 x 10^{-31} g
    d. 1.0 x 10^{21} atoms
    e. 3650. km

12. Carry out the following calculation and report the answer to the correct number of significant figures. $(22.71-2.3)\left[\frac{9.322*10^3}{103.10-92.2}\right]$
13. The aluminum in a package containing 75 ft^2 of kitchen foil weighs approximately 12 ounces. Aluminum has a density of 2.70 g/cm^3. What is the approximate thickness of the aluminum foil in millimeters? (1 oz = 28.4 g)
14. In July 1983, an Air Canada Boeing 767 ran out of fuel over central Canada on a trip from Montreal to Edmonton. (The plane glided safely to a landing at an abandoned airstrip.) The pilots knew that 22,300 kg of fuel were required for the trip, and they knew that 7682 L of fuel were already in the tank. The ground crew added 4916 L of fuel, which was only about one fifth of what was required. The crew members used a factor of 1.77 for the fuel density—the problem is that 1.77 has units of *pounds* per liter and not *kilograms* per liter! What is the fuel density in units of kg/L? What mass of fuel should have been loaded? (1 lb = 453.6 g.)
15. About two centuries ago, Benjamin Franklin showed that 1 teaspoon of oil would cover about 0.50 acre of still water. If you know that 1.0×10^4 m^2 = 2.47 acres, and that there is approximately 5.0 cm^3 in a teaspoon, what is the thickness of the layer of oil?
16. The anesthetic procaine hydrochloride is often used to deaden pain during dental surgery. The compound is packaged as a 10.% solution (by mass; d = 1.0 g/mL) in water. If your dentist injects 0.50 mL of the solution, what mass of procaine hydrochloride (in milligrams) is injected?

Answers to the Practice Problem Set:

1. Carbon, Uranium, Technetium, Silicon, Neon, Iridium
2. Ar, K, Rn, Pb, Bi, He
3. a) compound, element b) element, compound c) element, compound
4. 2.79 mL
5. a) 97 °C b) 289 K c) 310 K
6. 250 cm^3, 0.25 L, 2.5 x 10^{-4} m^3, 0.25 dm^3
7. 0.197 nm, 197 pm
8. 0.995 g Pt
9. zinc
10. The plastic will float, the metal will sink.
11. a) 3 b) 5 c) 3 d) 2 e) 4
12. 1.75 x 10^4
13. 1.8 x 10^{-2} mm
14. 0.803 kg/L; 12,200 kg (15,200 L) additional fuel needed
15. 2.5 x 10^{-7} cm
16. 50. mg procaine hydrochloride

CH 221 Practice Problem Set #2

*This is a **practice problem set** and not the actual graded problem set that you will turn in for credit.*
Answers to each problem can be found at the end of this assignment.

Covering: **Chapter Two, Chapter 3.1** *and* **Chapter Guide Two**
Important Tables and/or Constants: 1 mol = 6.022×10^{23}

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1. Give the mass number of each of the following atoms: a. magnesium with 15 neutrons, b. titanium with 26 neutrons, and c. zinc with 32 neutrons.
2. Give the complete symbol ($^{A}_{Z}X$) for each of the following atoms: a. potassium with 20 neutrons, b. krypton with 48 neutrons, and c. cobalt with 33 neutrons.
3. Thallium has two stable isotopes, ^{203}Tl and ^{205}Tl. Knowing that the atomic weight of thallium is 204.4, which isotope is the more abundant of the two?
4. Silver (Ag) has two stable isotopes, ^{107}Ag and ^{109}Ag. The isotopic mass of ^{107}Ag is 106.9051 and the isotopic mass of ^{109}Ag is 108.9047. The atomic weight of Ag, from the periodic table, is 107.868. Estimate the percentage of ^{107}Ag in a sample of the element.
   a. 0% b. 25% c. 50% d. 75%
5. Gallium has two naturally occurring isotopes, ^{69}Ga and ^{71}Ga, with masses of 68.9257 u and 70.9249 u, respectively. Calculate the percent abundances of these isotopes of gallium.
6. Calculate the mass in grams of:
   a. 2.5 mol aluminum b. 1.25×10^{-3} mol of iron
   c. 0.015 mol of calcium d. 653 mol of neon
7. Calculate the amount (moles) represented by each of the following:
   a. 127.08 g of Cu b. 0.012 g of lithium
   c. 5.0 mg of americium d. 6.75 g of Al
8. Classify the following elements as metals, metalloids, or nonmetals: N, Na, Ni, Ne, and Np.
9. Fill in the blanks in the table (one column per element):

Symbol	**^{58}Ni**	**^{33}S**		
Number of protons			**10**	
Number of neutrons			**10**	**30**
Number of electrons in the neutral atom				**25**
Name of the element				

10. Put the following elements in order from smallest to largest mass:
    a. 3.79×10^{24} atoms Fe b. 19.921 mol H_2 c. 8.576 mol C d. 7.4 mol Si
    e. 9.221 mol Na f. 4.07×10^{24} atoms Al g. 9.2 mol Cl_2
11. Dilithium is the fuel for the *Starship Enterprise*. Because its density is quite low, however, you need a large space to store a large mass. To estimate the volume required, we shall use the element lithium. If you need 256 mol for an interplanetary trip, what must the volume of the piece of lithium be? If the piece of lithium is a cube, what is the dimension of an edge of the cube? (The density for the element lithium is 0.534 g/cm^3 at 20 °C.)
12. A cylindrical piece of sodium is 12.00 cm long and has a diameter of 4.5 cm. The density of sodium is 0.971 g/cm^3. How many atoms does the piece of sodium contain? (The volume of a cylinder is $V = \pi \times r^2 \times$ length.)

13. To estimate the radius of a lead atom:
    a. You are given a cube of lead that is 1.000 cm on each side. The density of lead is 11.35 g/cm^3. How many atoms of lead are in the sample?
    b. Atoms are spherical; therefore, the lead atoms in this sample cannot fill all the available space. As an approximation, assume that 60% of the space of the cube is filled with spherical lead atoms. Calculate the volume of one lead atom from this information. From the calculated volume (V), and the formula $V = {}^4/_3\, \pi\, r^3$, estimate the radius (r) of a lead atom.
14. Reviewing the periodic table.
    a. Name an element in Group 2A.
    b. Name an element in the third period.
    c. Which element is in the second period in Group 4A?
    d. Which element is in the third period in Group 6A?
    e. Which halogen is in the fifth period?
    f. Which alkaline earth element is in the third period?
    g. Which noble gas element is in the fourth period?
    h. Name the nonmetal in Group 6A and the third period.
    i. Name a metalloid in the fourth period.

Answers to the Practice Problem Set:

1. a. 27 b. 48 c. 62
2. a. ${}^{39}_{19}K$ b. ${}^{84}_{36}Kr$ c. ${}^{60}_{27}Co$
3. Thallium-205
4. 50%
5. ${}^{69}Ga$ abundance is 60.12%, ${}^{71}Ga$ abundance is 39.88%
6. a. 68 g Al b. 0.0698 g Fe c. 0.60 g Ca d. 1.32×10^4 g Ne
7. a. 1.9998 mol Cu b. 1.7×10^{-3} mol Li c. 2.1×10^{-5} mol Am d. 0.250 mol Al
8. Metals: Na Ni Np Nonmetals: N, Ne
9. (left to right): Nickel-58, sulfur-33, neon-20, manganese-55
10. H_2 (b) < C (c) < Al (f) < Si (d) < Na (e) < Fe (a) < Cl_2 (g)
11. 3.33×10^3 cm^3 and 14.9 cm
12. 190 cm^3 and 4.9×10^{24} atoms
13. 3.299×10^{22} atoms and 1.631×10^{-8} cm
14. *Possible* answers: a. Ba b. Si c. C d. S e. I f. Mg g. Kr h. S i. As

CH 221 Practice Problem Set #3

*This is a **practice problem set** and not the actual graded problem set that you will turn in for credit. Answers to each problem can be found at the end of this assignment.*

Covering: **Chapter Two, Chapter 3.1-3.2** *and* **Chapter Guide Three**
Important Tables and/or Constants: 1 mol = 6.022×10^{23}, "Have No Fear Of Ice Clear Brew" (7 Diatomics)

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1. Give the symbol, including the correct charge, for each of the following ions:
   a. barium ion b. titanium(IV) ion c. phosphate ion d. hydrogen carbonate ion
   e. sulfide ion f. perchlorate ion g. cobalt(II) ion h. sulfate ion
2. When a potassium atom becomes a monatomic ion, how many electrons does it lose or gain? What noble gas atom is *isoelectronic* (i.e. has the same number of electrons) as a potassium ion?
3. For each of the following compounds, give the formula, charge, and the number of each ion that makes up the compound:
   a. K_2S b. $CoSO_4$ c. $KMnO_4$ d. $(NH_4)_3PO_4$ e. $Ca(ClO)_2$
4. Cobalt forms Co^{2+} and Co^{3+} ions. Write the formulas for the two cobalt oxides formed by these transition metal ions.
5. Which of the following are correct formulas for ionic compounds? For those that are not, give the correct formula.
   a. $AlCl_2$ b. KF_2 c. Ga_2O_3 d. MgS
6. Name each of the following ionic compounds:
   a. K_2S b. $CoSO_4$ c. $(NH_4)_3PO_4$ d. $Ca(ClO)_2$
7. Give the formula for each of the following ionic compounds:
   a. ammonium carbonate
   b. calcium iodide
   c. copper(II) bromide
   d. aluminum phosphate
   e. silver(I) acetate
8. Sodium ion, Na^+, forms ionic compounds with fluoride, F^-, and iodide, I^-. The radii of these ions are as follows: Na^+ = 116 pm; F^- = 119 pm; and I^- = 206 pm. In which ionic compound, NaF or NaI, are the forces of attraction between cation and anion stronger? Explain your answer.
9. Name each of the following binary, nonionic compounds:
   a. NF_3 b. HI c. BI_3 d. PF_5
10. Give the formula for each of the following compounds:
    a. sulfur dichloride
    b. dinitrogen pentaoxide
    c. silicon tetrachloride
    d. diboron trioxide
11. Calculate the molar mass of each of the following compounds:
    a. Fe_2O_3, iron(III) oxide
    b. BCl_3, boron trichloride
    c. $C_6H_8O_6$, ascorbic acid (vitamin C)

12. What mass is represented by 0.0255 mol of each of the following compounds?
    a. C_3H_7OH, propanol, rubbing alcohol
    b. $C_{11}H_{16}O_2$, an antioxidant in foods, also known as BHA (butylated hydroxyanisole)
    c. $C_9H_8O_4$, aspirin
13. Calculate the weight percent of lead in PbS, lead(II) sulfide. What mass of lead (in grams) is present in 10.0 g of PbS?
14. Succinic acid occurs in fungi and lichens. Its empirical formula is $C_2H_3O_2$ and its molar mass is 118.1 g/mol. What is its molecular formula?
15. A large family of boron-hydrogen compounds has the general formula B_xH_y. One member of this family contains 88.5% B; the remainder is hydrogen. Which of the following is its empirical formula: BH_2, BH_3, B_2H_5, B_5H_7, or B_5H_{11}?
16. A new compound containing xenon and fluorine was isolated by shining sunlight on a mixture of Xe (0.526 g) and F_2 gas. If you isolate 0.678 g of the new compound, what is its empirical formula?
17. The "alum" used in cooking is potassium aluminum sulfate hydrate, $KAl(SO_4)_2 * x\ H_2O$. To find the value of *x,* you can heat a sample of the compound to drive off all of the water and leave only $KAl(SO_4)_2$. Assume you heat 4.74 g of the hydrated compound and that the sample loses 2.16 g of water. What is the value of *x?*
18. Direct reaction of iodine (I_2) and chlorine (Cl_2) produces an iodine chloride, I_xCl_y, a bright yellow solid. If you completely used up 0.678 g of iodine and produced 1.246 g of I_xCl_y, what is the empirical formula of the compound? A later experiment showed that the molar mass of I_xCl_y was 467 g/mol. What is the molecular formula of the compound?

Answers to the Practice Problem Set:

1. a. Ba^{2+} b. Ti^{4+} c. PO_4^{3-} d. HCO_3^{-1} e. S^{2-} f. ClO_4^{-1} g. Co^{2+} h. SO_4^{2-}
2. One electron; argon.
3. *Answers:*
   a. 2 K^+ ions, 1 S^{2-} ion
   b. 1 Co^{2+} ion, 1 SO_4^{2-} ion
   c. 1 K^+ ion, 1 MnO_4^- ion
   d. 3 NH_4^+ ions, 1 PO_4^{3-} ion
   e. 1 Ca^{2+} ion, 2 ClO^- ions
4. CoO, Co_2O_3
5. a. incorrect, $AlCl_3$ b. incorrect, KF c. correct d. correct
6. a. potassium sulfide b. cobalt(II) sulfate c. ammonium phosphate d. calcium hypochlorite
7. *Answers:*
   a. $(NH_4)_2CO_3$
   b. CaI_2
   c. $CuBr_2$
   d. $AlPO_4$
   e. $AgCH_3CO_2$
8. NaF stronger, shorter cation-anion distance
9. *Answers:*
   a. nitrogen trifluoride
   b. hydrogen monoiodide
   c. boron triiodide
   d. phosphorus pentafluoride
10. a. SCl_2 b. N_2O_5 c. $SiCl_4$ d. B_2O_3
11. a. Fe_2O_3 159.69 g/mol b. BCl_3 117.17 g/mol c. $C_6H_8O_6$ 176.13 g/mol
12. a. 1.53 g b. 4.60 g c. 4.60 g
13. 86.59%, 8.66 g
14. $C_4H_6O_4$
15. B_5H_7
16. XeF_2
17. x = 12
18. I_2Cl_6

CH 221 Practice Problem Set #4

This is a ***practice problem set*** *and not the actual graded problem set that you will turn in for credit. Answers to each problem can be found at the end of this assignment.*

Covering: **Chapter Four** *and* **Chapter Guide Four**

Important Tables and/or Constants: **Solubility Table** (in the "Net Ionics" lab or here: **https://mhchem.org/sol**)- *Use the Net Ionics solubility table when answering questions about solubility in CH 221*)

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1. Balance the following equations:
   a. $Cr(s) + O_2(g) \rightarrow Cr_2O_3(s)$
   b. $Cu_2S(s) + O_2(g) \rightarrow Cu(s) + SO_2(g)$
   c. $C_6H_5CH_3(l) + O_2(g) \rightarrow H_2O(l) + CO_2(g)$
2. Balance the following equations and name each reactant and product:
   a. $Fe_2O_3(s) + Mg(s) \rightarrow MgO(s) + Fe(s)$
   b. $AlCl_3(s) + NaOH(aq) \rightarrow Al(OH)_3(s) + NaCl(aq)$
   c. $NaNO_3(s) + H_2SO_4(l) \rightarrow Na_2SO_4(s) + HNO_3(l)$
   d. $NiCO_3(s) + HNO_3(aq) \rightarrow Ni(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
3. Like many metals, aluminum reacts with a halogen to give a metal halide.
   $2\ Al(s) + 3\ Br_2(l) \rightarrow Al_2Br_6(s)$
   What mass of Br_2, in grams, is required for complete reaction with 2.56 g of Al? What mass of white, solid Al_2Br_6 is expected?
4. Aluminum chloride, $AlCl_3$, is made by treating scrap aluminum with chlorine.
   $2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3(s)$
   If you begin with 2.70 g of Al and 4.05 g of Cl_2,
   a. Which reactant is limiting?
   b. What mass of $AlCl_3$ can be produced?
   c. What mass of the excess reactant remains when the reaction is completed?
5. The deep blue compound $Cu(NH_3)_4SO_4$ is made by the reaction of copper(II) sulfate and ammonia: $CuSO_4(aq) + 4\ NH_3(aq) \rightarrow Cu(NH_3)_4SO_4(aq)$
   a. If you use 10.0 g of $CuSO_4$ and excess NH_3, what is the theoretical yield of $Cu(NH_3)_4SO_4$?
   b. If you isolate 12.6 g of $Cu(NH_3)_4SO_4$, what is the percent yield of $Cu(NH_3)_4SO_4$?
6. A sample of limestone and other soil materials is heated, and the limestone decomposes to give calcium oxide and carbon dioxide: $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
   A 1.506 g sample of limestone-containing material gives 0.558 g of CO_2, in addition to CaO, after being heated at a high temperature. What is the mass percent of $CaCO_3$ in the original sample?
7. Styrene, the building block of polystyrene, consists of only C and H. If 0.438 g of styrene is burned in oxygen and produces 1.481 g of CO_2 and 0.303 g of H_2O, what is the empirical formula of styrene?
8. Nickel forms a compound with carbon monoxide, $Ni_x(CO)_y$. To determine its formula, you carefully heat a 0.0973 g sample in air to convert the nickel to 0.0426 g of NiO and the CO to 0.100 g of CO_2. What is the empirical formula of $Ni_x(CO)_y$?

9. Menthol, from oil of mint, has a characteristic odor. The compound contains only C, H, and O. If 95.6 mg of menthol burns completely in O2, and gives 269 mg of CO_2 and 110 mg of H_2O, what is the empirical formula of menthol?
10. An unknown compound has the formula $C_xH_yO_z$. You burn 0.0956 g of the compound and isolate 0.1356 g of CO_2 and 0.0833 g of H_2O. What is the empirical formula of the compound? If the molar mass is 62.1 g/mol, what is the molecular formula?
11. Which compound or compounds in each of the following groups is (are) expected to be soluble in water?
    a. CuO, $CuCl_2$, $FeCO_3$
    b. AgI, Ag_3PO_4, $AgNO_3$
    c. K_2CO_3, KI, $KMnO_4$
12. The following compounds are water-soluble. What ions are produced by each compound in aqueous solution?
    a. KOH
    b. $LiNO_3$
    c. K_2SO_4
    d. $(NH_4)_2SO_4$
13. Decide whether each of the following is water-soluble. If soluble, tell what ions are produced.
    a. Na_2CO_3
    b. NiS
    c. $CuSO_4$
    d. $BaBr_2$
14. Predict the products of each precipitation reaction. Balance the completed equation, and then write the net ionic equation.
    a. $NiCl_2(aq) + (NH_4)_2S(aq) \rightarrow$
    b. $Mn(NO_3)_2(aq) + Na_3PO_4(aq) \rightarrow$
15. Balance the following equations, and then write the net ionic equation.
    a. $(NH_4)_2CO_3(aq) + Cu(NO_3)_2(aq) \rightarrow CuCO_3(s) + NH_4NO_3(aq)$
    b. $Pb(OH)_2(s) + HCl(aq) \rightarrow PbCl_2(s) + H_2O(l)$
    c. $BaCO_3(s) + HCl(aq) \rightarrow BaCl_2(aq) + H_2O(l) + CO_2(g)$

Answers to the Practice Problem Set:

1. *Answers:*
   a. $4\ Cr(s) + 3\ O_2(g) \rightarrow 2\ Cr_2O_3(s)$
   b. $Cu_2S(s) + O_2(g) \rightarrow 2\ Cu(s) + SO_2(g)$
   c. $C_6H_5CH_3(\ell) + 9\ O_2(g) \rightarrow 4\ H_2O(\ell) + 7\ CO_2(g)$
2. *Answers:*
   a. $Fe_2O_3(s) + 3\ Mg(s) \rightarrow 3\ MgO(s) + 2\ Fe(s)$
      iron(III) oxide, magnesium, magnesium oxide, iron
   b. $AlCl_3(s) + 3\ NaOH(aq) \rightarrow Al(OH)_3(s) + 3\ NaCl(aq)$
      aluminum chloride, sodium hydroxide, aluminum hydroxide, sodium chloride
   c. $2\ NaNO_3(s) + H_2SO_4(\ell) \rightarrow Na_2SO_4(s) + 2\ HNO_3(\ell)$
      sodium nitrate, hydrogen sulfate (sulfuric acid), sodium sulfate, hydrogen nitrate (nitric acid)
   d. $NiCO_3(s) + 2\ HNO_3(aq) \rightarrow Ni(NO_3)_2(aq) + CO_2(g) + H_2O(\ell)$
      nickel(II) carbonate, hydrogen nitrate (nitric acid), nickel(II) nitrate, carbon dioxide, water
3. 22.7 g Br_2; 25.3 g Al_2Br_6
4. a. Cl_2 b. 5.09 g c. 1.67 g Al
5. a. 14.3 g b. 88.3%
6. 84.3%
7. CH
8. $Ni(CO)_4$
9. $C_{10}H_{19}O$
10. EF = CH_3O, MF = $C_2H_6O_2$
11. a. $CuCl_2$ b. $AgNO_3$ c. all three compounds
12. a. K^+ and OH^- ions b. Li^+ and NO_3^- ions c. K^+ and SO_4^{2-} ions d. NH_4^+ and SO_4^{2-} ions
13. a. soluble, Na^+ and CO_3^{2-} ions b. insoluble c. soluble, Cu^{2+} and SO_4^{2-} ions d. soluble, Ba^{2+} and Br^- ions
14. *Answers:*
    a. $NiCl_2(aq) + (NH_4)_2S(aq) \rightarrow NiS(s) + 2\ NH_4Cl(aq)$
       $Ni^{2+}(aq) + S^{2-}(aq) \rightarrow NiS(s)$
    b. $3\ Mn(NO_3)_2(aq) + 2\ Na_3PO_4(aq) \rightarrow Mn_3(PO_4)_2(s) + 6\ NaNO_3(aq)$
       $3\ Mn^{2+}(aq) + 2\ PO_4^{3-}(aq) \rightarrow Mn_3(PO_4)_2(s)$
15. *Answers:*
    a. $(NH_4)_2CO_3(aq) + Cu(NO_3)_2 \rightarrow CuCO_3(s) + 2\ NH_4NO_3(aq)$
       $CO_3^{2-}(aq) + Cu^{2+}(aq) \rightarrow CuCO_3(s)$
    b. $Pb(OH)_2(s) + 2\ HCl(aq) \rightarrow PbCl_2(s) + 2\ H_2O(\ell)$
       $Pb(OH)_2(s) + 2\ H^+(aq) + 2\ Cl^-(aq) \rightarrow PbCl_2(s) + 2\ H_2O(\ell)$
    c. $BaCO_3(s) + 2\ HCl(aq) \rightarrow BaCl_2(aq) + H_2O(\ell) + CO_2(g)$
       $BaCO_3(s) + 2\ H^+(aq) \rightarrow Ba^{2+}(aq) + H_2O(\ell) + CO_2(g)$

CH 221 Practice Problem Set #5

*This is a **practice problem set** and not the actual graded problem set that you will turn in for credit. Answers to each problem can be found at the end of this assignment.*

Covering: **Chapter Three (3.3-3.4), Chapter Five** *and* **Chapter Guide Five**
Important Tables/Constants: **$C(H_2O)$ = 4.184 J g^{-1} K^{-1}**, **$\log_{10}$ x = ln x / ln 10** *and the* ***Thermodynamic Values*** *found in problem set #5 and here:* **http://mhchem.org/thermo**

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1. Determine the oxidation number of each element in the following ions or compounds.
   a. BrO_3^- b. $C_2O_4^{2-}$ c. F^- d. CaH_2 e. H_4SiO_4 f. HSO_4^-
2. Which two of the following reactions are oxidation–reduction reactions? Explain your answer in each case. Classify the remaining reaction.
   a. $Zn(s) + 2\ NO_3^{-1}(aq) + 4\ H^+(aq) \rightarrow Zn^{2+}(aq) + 2\ NO_2(g) + 2\ H_2O(l)$
   b. $Zn(OH)_2(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + 2\ H_2O(l)$
   c. $Ca(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$
3. In the following reactions, decide which reactant is oxidized and which is reduced. Designate the oxidizing agent and the reducing agent.
   a. $C_2H_4(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(g)$
   b. $Si(s) + 2\ Cl_2(g) \rightarrow SiCl_4(l)$
4. Some potassium dichromate ($K_2Cr_2O_7$), 2.335 g, is dissolved in enough water to make exactly 500. mL of solution. What is the molar concentration of the potassium dichromate? What are the molar concentrations of the K^+ and $Cr_2O_7^{2-}$ ions?
5. For each solution, identify the ions that exist in aqueous solution, and specify the concentration of each ion.
   a. 0.25 M $(NH_4)_2SO_4$
   b. 0.123 M Na_2CO_3
   c. 0.056 M HNO_3
6. A table wine has a pH of 3.40. What is the hydrogen ion concentration of the wine? Is it acidic or basic?
7. What volume of 0.109 M HNO_3, in milliliters, is required to react completely with 2.50 g of $Ba(OH)_2$? $2\ HNO_3(aq) + Ba(OH)_2(s) \rightarrow 2\ H_2O(l) + Ba(NO_3)_2(aq)$
8. You have 0.954 g of an unknown acid, H_2A, which reacts with NaOH according to the balanced equation: $\mathbf{H_2A(aq) + 2\ NaOH(aq) \rightarrow Na_2A(aq) + 2\ H_2O(l)}$ If 36.04 mL of 0.509 M NaOH is required to titrate the acid to the equivalence point, what is the molar mass of the acid?
9. `The specific heat capacity of copper is 0.385 J/g·K. What quantity of heat is required to heat 168 g of copper from -12.2 °C to +25.6 °C?
10. The initial temperature of a 344 g sample of iron is 18.2 °C. If the sample absorbs 2.25 kJ of heat, what is its final temperature? C_{Fe} = 0.449 J/g·K
11. One beaker contains 156 g of water at 22 °C and a second beaker contains 85.2 g of water at 95 °C. The water in the two beakers is mixed. What is the final water temperature?
12. A 237 g piece of molybdenum, initially at 100.0 °C, is dropped into 244 g of water at 10.0 °C. When the system comes to thermal equilibrium, the temperature is 15.3 °C. What is the specific heat capacity of molybdenum?

13. What quantity of heat is required to vaporize 125 g of benzene, C_6H_6, at its boiling point, 80.1 °C? The heat of vaporization of benzene is 30.8 kJ/mol.
14. Isooctane (2,2,4-trimethylpentane), one of the many hydrocarbons that make up gasoline, burns in air to give water and carbon dioxide.

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(l) \quad \Delta H°_{rxn} = -10{,}922\ kJ$$

If you burn 1.00 L of isooctane (density = 0.69 g/mL), what quantity of heat is evolved?

15. The enthalpy changes for the following reactions can be measured:

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g) \qquad \Delta H° = -802.4\ kJ$$

$$CH_3OH(g) + {}^3/_2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g) \qquad \Delta H° = -676\ kJ$$

Use these values and Hess's law to determine the enthalpy change for the reaction:

$$CH_4(g) + {}^1/_2\ O_2(g) \rightarrow CH_3OH(g) \qquad \Delta H° = ?$$

16. Enthalpy changes for the following reactions can be determined experimentally:

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g) \qquad \Delta H° = -91.8\ kJ$$

$$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g) \qquad \Delta H° = -906.2\ kJ$$

$$H_2(g) + {}^1/_2\ O_2\ (g) \rightarrow H_2O(g) \qquad \Delta H° = -241.8\ kJ$$

Use these values to determine the enthalpy change for the formation of NO(g) from the elements (an enthalpy change that cannot be measured directly because the reaction is reactant-favored) of

$${}^1/_2\ N_2(g) + {}^1/_2\ O_2(g) \rightarrow NO(g) \qquad \Delta H° = ?$$

17. Write a balanced chemical equation for the formation of $Li_2CO_3(s)$ from the elements in their standard states. Find the value of $\Delta H_f°$ for $Li_2CO_3(s)$ in the appendix of your textbook.
18. Use standard heats of formation in the appendix of your textbook to calculate standard enthalpy changes for the following:
    a. 1.0 g of white phosphorus burns, forming $P_4O_{10}(s)$
    b. 0.20 mol of NO(g) decomposes to $N_2(g)$ and $O_2(g)$
    c. 2.40 g of NaCl is formed from Na(s) and excess $Cl_2(g)$
    d. 250 g of iron is oxidized with oxygen to $Fe_2O_3(s)$
19. The Romans used calcium oxide, CaO, to produce a strong mortar to build stone structures. The CaO was mixed with water to give $Ca(OH)_2$, which reacted slowly with CO_2 in the air to give $CaCO_3$. $\mathbf{Ca(OH)_2(s) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(g)}$
    a. Calculate the standard enthalpy change for this reaction.
    b. What quantity of heat is evolved or absorbed if 1.00 kg of $Ca(OH)_2$ reacts with a stoichiometric amount of CO_2?

Answers to the Practice Problem Set:

1. *Answers:*
   a. Br is +5 and O is –2 d. Ca is +2 and H is –1
   b. C is +3 and O is –2 e. H is +1, Si is +4, and O is –2
   c. F is –1 f. H is +1, S is +6, and O is –2
2. *Answers:*
   a. oxidation-reduction reaction
      Oxidation # of Zn changes from 0 to +2, N changes from +5 to +4
   b. acid-base reaction
   c. oxidation-reduction reaction
      Oxidation number of Ca changes from 0 to +2, H from +1 to 0
3. a. C_2H_4 is oxidized / reducing agent; O_2 is reduced / oxidizing agent b. Si is oxidized / reducing agent; Cl_2 is reduced / oxidizing agent
4. $[Cr_2O_7^{2-}] = [K_2Cr_2O_7] = 0.0159$ M, $[K^+] = 0.0318$ M
5. a. 0.50 M NH_4^+; 0.25 M SO_4^{2-} b. 0.246 M Na^+; 0.123 M CO_3^{2-} c. 0.056 M H^+; 0.056 M NO_3^-
6. $[H^+] = 4.0 \times 10^{-4}$ M, acidic
7. 268 mL
8. 104 g/mol
9. 2440 J
10. 32.8 °C
11. 48 °C
12. 0.27 J/g·K
13. 49.3 kJ
14. 3.3 x 10^4 kJ heat evolved
15. -126 kJ
16. 90.3 kJ
17. $2\,Li(s) + C(s) + {}^{3}/_{2}\,O_2(g) \rightarrow Li_2CO_3(s)$ $\Delta H_f^o = -1216.04$ kJ *(OpenStax)*
18. a. -24 kJ b. -18 kJ c. -16.9 kJ d. -1800 kJ
19. a. -83.1 kJ b. -1120 kJ evolved

CH 221 Practice Problem Set #6

*This is a **practice problem set** and not the actual graded problem set that you will turn in for credit.*
Answers to each problem can be found at the end of this assignment.

Covering: **Chapter Six** *and* **Chapter Guide Six**
Important Tables and/or Constants: **c = 2.998 x 10^8 m/s, h = 6.626 x 10^{-34} J s**, the **Electromagnetic Spectrum** *and* **Subshell Filling Order diagrams** *on page 4 of Problem Set #6*. ***Memorize c and h!***

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1. Traffic signals are often now made of LEDs (light-emitting diodes). The light from an amber signal has a wavelength of 595 nm, and that from a green signal has wavelength of 500 nm. Which has the higher frequency? Which has the highest energy per photon? Calculate the frequency of amber light.
2. Place the following types of radiation in order of increasing energy per photon:
   a. yellow light from a sodium lamp
   b. x-rays from an instrument in a dentist's office
   c. microwaves in a microwave oven
   d. your favorite FM music station at 91.7 MHz
3. The most prominent line in the spectrum of aluminum is at 396.15 nm. What is the frequency of this line? What is the energy of one photon with this wavelength? Of 1.00 mol of these photons?
4. The energy emitted when an electron moves from a higher energy state to a lower energy state in any atom can be observed as electromagnetic radiation.
   a. Which involves the emission of less energy in the H atom, an electron moving from $n = 4$ to $n = 2$ or an electron moving from $n = 3$ to $n = 2$?
   b. Which involves the emission of more energy in the H atom, an electron moving from $n = 4$ to $n = 1$ or an electron moving from $n = 5$ to $n = 2$? Explain fully.
5. A rifle bullet (mass = 1.50 g) has a velocity of 7.00 x 10^2 miles per hour. What is the wavelength associated with this bullet? (0.6214 miles = 1 km)
6. a. When $n = 4$, what are the possible values of ℓ?
   b. When ℓ is 2, what are the possible values of m_ℓ?
   c. For a 4*s* orbital, what are the possible values of n, ℓ, and m_ℓ?
   d. For a 4*f* orbital, what are the possible values of n, ℓ, and m_ℓ?
7. Which of the following represent valid sets of quantum numbers? For a set that is invalid, explain briefly why it is not correct.
   a. $n = 3$, $\ell = 3$, $m_\ell = 0$, $m_s = 0$
   b. $n = 2$, $\ell = 1$, $m_\ell = 0$, $m_s = +1/2$
   c. $n = 6$, $\ell = 5$, $m_\ell = -1$, $m_s = -1/2$
   d. $n = 4$, $\ell = 3$, $m_\ell = -4$, $m_s = -1/2$

8. What is the maximum number of orbitals that can be identified by each of the following sets of quantum numbers? When “none” is the correct answer, explain your reasoning.
   a. $n = 3$, $\ell = 0$, $m_\ell = -1$
   b. $n = 5$, $\ell = 1$
   c. $n = 7$, $\ell = 5$
   d. $n = 4$, $\ell = 2$, $m_\ell = -2$, $m_s = -1/2$
9. How many nodal surfaces (planar *and* spherical) are associated with each of the following atomic orbitals?
   a. 5*f*
   b. 1*p*
   c. 4*s*
10. Answer the following questions:
   a. The quantum number *n* describes the ______ of an atomic orbital.
   b. The shape of an atomic orbital is given by the quantum number ______.
   c. A photon of green light has ______ (less or more) energy than a photon of orange light.
   d. The maximum number of orbitals that may be associated with the set of quantum numbers $n = 4$ and $\ell = 3$ is ______.
   e. The maximum number of orbitals that may be associated with the quantum number set $n = 3$, $\ell = 2$, and $m_\ell = -2$ is ______.
   f. Sketch a *d*, *s* and *p* orbital.
   g. When $n = 5$, the possible values of ℓ are ______.
   h. The number of orbitals in the $n = 4$ shell is _____.
   i. The maximum number of orbitals that may be associated with the quantum number set $n = 3$, $\ell = 3$, $m_\ell = -2$ and $m_s = -1/2$ is ______.
   j. The maximum number of orbitals that may be associated with the quantum number set $n = 6$, $\ell = 0$, $m_\ell = 0$ and $m_s = +1/2$ is ______.
11. Write the electron configurations for P and Mg using *spdf* notation and orbital box diagrams.
12. Depict the electron configuration for each of the following atoms using *spdf* and/or noble gas notations.
   a. Arsenic, As. A deficiency of As can impair growth in animals even though larger amounts are poisonous.
   b. Krypton, Kr. It ranks seventh in abundance of the gases in the earth’s atmosphere.
   c. Vanadium, V, an element found in some brown and red algae and some toadstools
13. Using orbital box diagrams and noble gas notation, depict the electron configurations of the following: (a) V, (b) V^{2+}, and (c) V^{5+}. Are any of the ions paramagnetic?
14. Arrange the following elements in order of increasing size: Al, B, C, K, and Na.
15. Name the element corresponding to each characteristic below.
   a. the element with the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^3$
   b. the alkaline earth element with the smallest atomic radius
   c. the element with the largest ionization energy in Group 5A
   d. the element whose 2+ ion has the configuration $[Kr]4d^5$
   e. the element with the most negative electron affinity in Group 7A
   f. the element whose electron configuration is $[Ar]3d^{10}4s^2$

Answers to the Practice Problem Set:

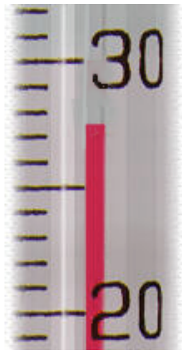
1. Green, green, 5.04×10^{14} Hz
2. FM radio (d) < microwaves (c) < yellow light (a) < X-rays (b)
3. 7.5676×10^{14} s^{-1}, 5.0144×10^{-19} J/photon, 3.02×10^{5} J/mol photons
4. a. $n = 3$ to $n = 2$ b. $n = 4$ to $n = 1$
5. 1.41×10^{-33} m
6. *Answers:*
   a. ℓ can be 0, 1, 2, 3
   b. m_ℓ can be 0, ±1, ±2
   c. $n = 4$, $\ell = 0$, $m_\ell = 0$
   d. $n = 4$, $\ell = 3$, $m_\ell = 0, \pm1, \pm2, \pm3$
7. *Answers:*
   b. and c. are valid sets of quantum numbers
   a. incorrect; when $n = 3$, the maximum value of ℓ is 2; also, m_s = *must* be $+^1/_2$ or $-^1/_2$
   d. incorrect; when $\ell = 3$, m_ℓ can only have values of 0, ±1, ±2, or ±3
8. a. none b. 3 c. 11 d. 1
9. a. 3 planar, 1 spherical b. zero planar, zero spherical c. zero planar, three spherical
10. a. size, energy b. ℓ c. more d. 7 e. 1 f. “cloverleaf”, “spherical”, “figure eight/dumbbell”
    g. 0, 1, 2, 3, 4 h. 16 i. 0 j. 1
11. P: $1s^2 2s^2 2p^6 3s^2 3p^3$ Mg: $1s^2 2s^2 2p^6 3s^2$
12. a. [Ar]$3d^{10} 4s^2 4p^3$ b. [Ar]$3d^{10} 4s^2 4p^6$ c. [Ar]$3d^3 4s^2$
13. a. [Ar]$3d^3 4s^2$ b. [Ar]$3d^3$ c. [Ar]
14. C < B < Al < Na < K
15. a. P b. Be c. N d. Tc e. Cl f. Zn

CH 221 Sig Figs / Chapter 1
Self Quiz Name: ______________________________ Lab Section: _______

This is a sample quiz for CH 221 providing examples from Chapter 1. Answers are provided at the end of this handout. *Good luck!*

---

1. How many significant digits are present in the temperature read from the thermometer illustrated to the right?



   a) 1 b) 2 c) 3 d) 4

2. The dimensions of a rectangular solid are 8.00 cm long, 4.00 cm wide, and 2.00 cm high. If the density of the solid is 10.0 g/cm^3, what is its mass?
   a) 10/64 grams
   b) 10.0 grams
   c) 64.0 grams
   d) 320. grams
   e) 640. grams

3. A metal sample weighing 30.9232 grams was added to a graduated cylinder containing 23.26 mL of water. The volume of water plus the sample was 24.85 mL. Which setup will result in the density of this metal?
   a) 30.9232 x (24.85-23.26)
   b) $\dfrac{30.9232}{24.85-23.26}$
   c) $\dfrac{24.85-23.26}{30.9232}$
   d) 30.9232 x $\dfrac{24.85}{23.26}$
   e) $\dfrac{30.9232}{24.85+23.26}$

4. The number of significant digits in 0.30500 is
   a) 1
   b) 2
   c) 3
   d) 4
   e) 5

5. A box measures 3.50 cm x 2.915 cm. The product of these numbers = 10.2025 cm^2. What is the proper way to report the area of the box?
   a) 10.20 cm^2
   b) 10.2 cm^2
   c) 10 cm^2
   d) 10. cm^2

6. The result of 2.350 x (4.0 + 6.311) is,
   a) 24
   b) 24.2
   c) 24.21
   d) 24.205

7. A student does a calculation using her calculator and the number 280.27163 is shown on the display. If there are actually three significant figures, how should she show the final answer?
   a) 280
   b) 280.3
   c) 280.27
   d) 2.80 x 10^{-2}
   e) 2.80 x 10^{2}

8. The term that refers to the reproducibility of a laboratory measurement is
   a) precision
   b) repeatability
   c) accuracy
   d) exactness

9. Which measurement below is NOT written with three significant digits?
   a) 2.00 cm
   b) 550. grams
   c) 0.003 L
   d) 12.7 mm

10. The number 6.33×10^2 equals,
    a) 6.33
    b) 0.633
    c) 633
    d) 0.0633

11. The number 6.33×10^{-2} equals,
    a) 6.33
    b) 0.633
    c) 633
    d) 0.0633

12. Calculate the following:
    150.3 - 107.240
    a) 43
    b) 43.1
    c) 43.06
    d) 43.060

13. Calculate the following:
    322.44 - 0.321 - 72.0 - 68.9555
    a) 181.1635
    b) 181.164
    c) 181.16
    d) 181.2
    e) 181

14. Calculate the following:
    18.3 * (375 - 289) / 1.16
    a) 1356.72
    b) 1356.7
    c) 1357
    d) 1360
    e) 1400

15. Which exhibits the largest length?
    a) 0.100 km
    b) 250 cm
    c) 1.7×10^6 mm
    d) 450,000 nm

16. The prefix “nano-” corresponds to what multiplication factor?
    a) 10^{-9}
    b) 10^{-6}
    c) 10^{-3}
    d) 10^{-2}
    e) 10^3

17. The prefix “milli-” corresponds to what multiplication factor?
    a) 10^{-9}
    b) 10^{-6}
    c) 10^{-3}
    d) 10^{-2}
    e) 10^3

18. The prefix “micro-” corresponds to what multiplication factor?
    a) 10^{-9}
    b) 10^{-6}
    c) 10^{-3}
    d) 10^{-2}
    e) 10^3

19. The prefix “centi-” corresponds to what multiplication factor?
    a) 10^{-9}
    b) 10^{-6}
    c) 10^{-3}
    d) 10^{-2}
    e) 10^3

20. Convert 32.0 cm into nm.
    a) 3.2×10^{-6}
    b) 3.20×10^{-6}
    c) 3.2×10^8
    d) 3.20×10^8
    e) 320.

21. Convert 475 mL into L.
    a) 4.75 L
    b) 0.475 L
    c) 47.5 L
    d) 0.0475 L
    e) .5 L

22. Convert 367 K into °C.
    a) 93.85 °C
    b) 93.9 °C
    c) 94 °C
    d) 90 °C
    e) 640. °C

23. Convert -212.1 °C into K.
    a) -61.05 K
    b) 61.05 K
    c) 61.1 K
    d) 61 K
    e) 60 K

24. Convert 32.1 °C into °F.
    a) 89.78 °F
    b) 89.8 °F
    c) 90. °F
    d) 90 °F
    e) 100 °F

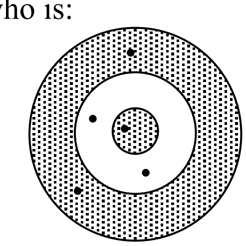
25. You measure the density of a slab of lead as 11.10 g/mL. The accepted value is 11.34 g/mL. The percent error for your measurement is
    a) 2.1 %
    b) 2.4 %
    c) 3.7 %
    d) 5.1 %

26. A sample of ore with a mass of 44.15 g contains aluminum and oxygen. Chemical analysis shows the sample contains 23.0 g of aluminum. The percent oxygen in the sample is
    a) 47.90 %
    b) 47.9 %
    c) 52.1 %
    d) 52.10 %

27. Which one of the following elements is correctly matched with its symbol?
    a) Ag, gold
    b) Ni, nickel
    c) Fl, fluorine
    d) Mg, manganese
    e) H, helium

28. Which one of the following elements is correctly matched with its symbol?
    a) P, potassium
    b) S, sodium
    c) Mn, magnesium
    d) Os, osmium
    e) B, beryllium

29. The marks on the following target represent someone who is:



    a) accurate, but not precise.
    b) precise, but not accurate.
    c) both accurate and precise.
    d) neither accurate nor precise.

30. You need 36.7 g of Fe from a sample that is 36.0% iron by mass. How many grams of the sample will you need?
    a) 101.94 g
    b) 101.9 g
    c) 102 g
    d) 13.212 g
    e) 13.2 g

Answers:

1.	**C**	16.	**A**
2.	**E**	17.	**C**
3.	**B**	18.	**B**
4.	**E**	19.	**D**
5.	**B**	20.	**D**
6.	**B**	21.	**B**
7.	**E**	22.	**C**
8.	**A**	23.	**C**
9.	**C**	24.	**B**
10.	**C**	25.	**A**
11.	**D**	26.	**B**
12.	**B**	27.	**B**
13.	**D**	28.	**D**
14.	**E**	29.	**D**
15.	**C**	30.	**C**

CH 221 Name: ____________________

Dimensional Analysis Worksheet

Directions: You must show all work and it must be presented in a neat and orderly fashion. The numbers in your setups and answers must include proper units and significant figures. You must use proper dimensional analysis technique, which means use one continuous conversion. **Answers appear immediately following the problems.**

---

1. Convert 124.0 days into seconds.

2. Convert 9.75×10^7 fluid ounces of water (density = 0.99998 g/mL) into metric tons.

3. Convert 3.87×10^{-8} km into cm.

4. Convert 67 U.S. quarts into kL.

5. Convert 6.5 pounds into cups if the density of the liquid is 2.03 g/L.

6. Convert 3.409 miles per hour into km per minute.

7. Convert 56.2 m^3 into yd^3.

8. What is the density of a mystery liquid in g per mL if 65.0 fluid ounces weighs 202 mass ounces?

9. A piece of gold leaf (density 19.3 g/cm^3) weighs 1.93 mg. What is the volume in mm^3?

10. What is a better deal, a one gallon gasoline for $2.89 or one liter of gasoline for $0.75? Support you answer using calculations.

11. A car travels at a rate of 65 miles per hour. If the car gets 33.5 miles to the gallon, how many hours can a car travel on 25.0 pounds of fuel? (density of fuel is 6.50 pounds/gallon)

12. The recommended dose of a medication is 5 mg/kg body weight. You have a patient whose weight is 125 pounds. The pharmacy offers three different oils containing 500 mg, 250 mg, and 100 mg of medication. Which pill should you give your patient?

13. The bromine content of the ocean is about 65 grams of bromine per million grams of sea water. How many cubic meters of ocean must be processed to recover 1.0 pounds of bromine if the density of sea water is 1.0×10^3 kg/m^3?

14. An average man is requires about 2.00 mg of fiboflavin (vitamin B2) per day. Cheese contains 5.5µg of riboflavin per gram of cheese. How many pounds of cheese would a man have to eat per day if this is his only source of riboflavin?

15. Alan is going to the Boy Scouts Jamboree in D.C. next summer and he has been asked to bring the smores supply for all the boys going from the district in Oregon. Each giant chocolate bar makes 16 smores. Each boy will be limited to exactly 3 smores. The problem is that he has to buy the chocolate once he gets to D.C. because there will be too many of them and they may melt in the summer heat. On average, the stores only carry 25 of these giant chocolate bars in stock. How many stores will he have to visit if there are 2,225 boys?

16. In the yearly fundraiser at school, kids can earn a hamburger phone if they raise at least $250 in donations. Aaron was able to get all of his family and family friends to pledge enough money that he will earn $35 for each mile he runs. The problem is he runs very slowly, 88 inches per second. How many hours will it take him to run just long enough to earn the hamburger phone?

Answers to the Dimensional Analysis Worksheet:

1. $1.071 * 10^7$ s
2. $2.88 * 10^3$ tons
3. $3.87 * 10^{-3}$ cm
4. $6.3 * 10^{-2}$ kL
5. $6.1 * 10^3$ cups
6. $9.142 * 10^{-2}$ km
7. 73.6 yd^3
8. 2.98 g/mL
9. 0.100 mm^3
10. $0.75/L (which equals $2.84 / gallon)
11. 2.0 hr
12. Use 250 mg pill (answer = 300 mg)
13. 7.0 m^3
14. 0.80 lb
15. 17 stores
16. 1.4 hr

Nomenclature
Self Quiz Name: ______________________________ Lab Section: _______

This is a sample quiz providing examples of nomenclature. Answers are provided at the end of this handout. *Good luck!*

Provide names or formulas for the following compounds:

nitrogen trifluoride	nitrogen monoxide	nitrogen dioxide
dinitrogen tetroxide	dinitrogen monoxide	phosphorus trichloride
phosphorus pentachloride	sulfur hexafluoride	disulfur decafluoride
xenon tetrafluoride	CCl_4	P_4O_{10}
ClF_3	BCl_3	SF_4
HBr(g)	N_2F_2	XeF_3
PI_3	SCl_2	S_2Cl_2
OF_2	NCl_3	$AsCl_5$

Nomenclature
Self Quiz Name: ____ Lab Section: _______

This is a sample quiz providing examples of nomenclature. Answers are provided at the end of this handout. *Good luck!*

Provide a formula for the following combinations of cation and anion:

	Cl^-	NO_3^-	S^{2-}	CO_3^{2-}	N^{3-}	PO_4^{3-}	OH^-
Na^+							
NH_4^+							
Sn^{2+}							
Hg_2^{2+}							
Al^{3+}							
Sn^{4+}							

Provide the formula and name for the following combinations of cations and anions:

Cation	**Anion**	**Formula**	**Name**
Cu^{2+}	OH^-		
Ba^{2+}	SO_4^{2-}		
NH_4^+	$Cr_2O_7^{2-}$		
Ag^+	$C_2H_3O_2^-$		
Fe^{3+}	S^{2-}		

Provide names and/or formulas for the following:

Formula	**Name**
HCl(aq)	
$HBrO_2$(aq)	
H_2SO_4(aq)	
HNO_2(aq)	
HIO(aq)	
HIO_4(aq)	
NaOH	
LiOH	
NH_4OH	
$Mg(OH)_2$	

Formula	**Name**
	hydrobromic acid
	chlorous acid
	sulfurous acid
	hydrosulfuric acid
	nitric acid
	phosphoric acid
	phosphorous acid
	potassium hydroxide
	calcium hydroxide
	dihydrogen monoxide

nitrogen trifluoride **NF_3**	nitrogen monoxide **NO**	nitrogen dioxide **NO_2**
dinitrogen tetroxide **N_2O_4**	dinitrogen monoxide **N_2O**	phosphorus trichloride **PCl_3**
phosphorus pentachloride **PCl_5**	sulfur hexafluoride **SF_6**	disulfur decafluoride **S_2F_{10}**
xenon tetrafluoride **XeF_4**	CCl_4 **carbon tetrachloride**	P_4O_{10} **tetraphosphorus decaoxide**
ClF_3 **chlorine trifluoride**	BCl_3 **boron trichloride**	SF_4 **sulfur tetrafluoride**
HBr(g) **hydrogen monobromide** *(not an acid)*	N_2F_2 **dinitrogen difluoride**	XeF_3 **xenon trifluoride**
PI_3 **phosphorus triiodide**	SCl_2 **sulfur dichloride**	S_2Cl_2 **disulfur dichloride**
OF_2 **oxygen difluoride**	NCl_3 **nitrogen trichloride**	$AsCl_5$ **arsenic pentachloride**

Nomenclature *Self Quiz* - **ANSWERS APPEAR IN BOLD**

Provide a formula for the following combinations of cation and anion:

	Cl^-	NO_3^-	S^{2-}	CO_3^{2-}	N^{3-}	PO_4^{3-}	OH^-
Na^+	**$NaCl$**	**$NaNO_3$**	**Na_2S**	**Na_2CO_3**	**Na_3N**	**Na_3PO_4**	**$NaOH$**
NH_4^+	**NH_4Cl**	**NH_4NO_3**	**$(NH_4)_2S$**	**$(NH_4)_2CO_3$**	**$(NH_4)_3N$**	**$(NH_4)_3PO_4$**	**NH_4OH**
Sn^{2+}	**$SnCl_2$**	**$Sn(NO_3)_2$**	**SnS**	**$SnCO_3$**	**Sn_3N_2**	**$Sn_3(PO_4)_2$**	**$Sn(OH)_2$**
Hg_2^{2+}	**Hg_2Cl_2**	**$Hg_2(NO_3)_2$**	**Hg_2S**	**Hg_2CO_3**	**$(Hg_2)_3N_2$**	**$(Hg_2)_3(PO_4)_2$**	**$Hg_2(OH)_2$**
Al^{3+}	**$AlCl_3$**	**$Al(NO_3)_3$**	**Al_2S_3**	**$Al_2(CO_3)_3$**	**AlN**	**$AlPO_4$**	**$Al(OH)_3$**
Sn^{4+}	**$SnCl_4$**	**$Sn(NO_3)_4$**	**SnS_2**	**$Sn(CO_3)_2$**	**Sn_3N_4**	**$Sn_3(PO_4)_4$**	**$Sn(OH)_4$**

Provide the formula and name for the following combinations of cations and anions:

Cation	**Anion**	**Formula**	**Name**
Cu^{2+}	OH^-	**$Cu(OH)_2$**	**copper(II) hydroxide**
Ba^{2+}	SO_4^{2-}	**$BaSO_4$**	**barium sulfate**
NH_4^+	$Cr_2O_7^{2-}$	**$(NH_4)_2Cr_2O_7$**	**ammonium dichromate**
Ag^+	$C_2H_3O_2^-$	**$AgC_2H_3O_2$**	**silver(I) acetate**
Fe^{3+}	S^{2-}	**Fe_2S_3**	**iron(III) sulfide**

Provide names and/or formulas for the following:

Formula	**Name**
$HCl(aq)$	**hydrochloric acid**
$HBrO_2(aq)$	**bromous acid**
$H_2SO_4(aq)$	**sulfuric acid**
$HNO_2(aq)$	**nitrous acid**
$HIO(aq)$	**hypoiodous acid**
$HIO_4(aq)$	**periodic acid**
NaOH	**sodium hydroxide**
LiOH	**lithium hydroxide**
NH_4OH	**ammonium hydroxide**
$Mg(OH)_2$	**magnesium hydroxide**

Formula	**Name**
$HBr(aq)$	hydrobromic acid
$HClO_2(aq)$	chlorous acid
$H_2SO_3(aq)$	sulfurous acid
$H_2S(aq)$	hydrosulfuric acid
$HNO_3(aq)$	nitric acid
$H_3PO_4(aq)$	phosphoric acid
$H_3PO_3(aq)$	phosphorous acid
KOH	potassium hydroxide
$Ca(OH)_2$	calcium hydroxide
H_2O	dihydrogen monoxide

CH 221 Nomenclature
Self Quiz Name: ______________________________ Lab Section: _______

This is a sample quiz for CH 221 providing examples of nomenclature. Answers are provided at the end of this handout. *Good luck!*

Name the following compounds. Use acid names where appropriate.

1. NaF
2. $MgCl_2$
3. $CaCO_3$
4. $Ca(HCO_3)_2$
5. $KHCO_3$
6. $AgNO_3$
7. $Zn(NO_2)_2$
8. $CdSO_4$
9. $Al_2(SO_3)_3$
10. $Au_3(PO_4)_2$
11. Cu_3PO_3
12. $NaClO_3$
13. $Ba(ClO_2)_2$
14. AgClO
15. $Fe(ClO_4)_3$
16. $Co(BrO_3)_2$
17. $Mn(BrO_2)_2$
18. $Cr(BrO)_3$
19. $Zn(BrO_4)_2$
20. $AgIO_3$
21. $Al(IO_2)_3$
22. NaIO
23. $LiIO_4$

24. $Ti(HSO_3)_2$

25. SnO_2

26. N_2O

27. CO

28. P_2O_5

29. CCl_4

30. SO_3

31. Se_2F_6

32. HCl

33. $HClO_3$

34. $HBrO_2$

35. HIO

36. $HClO_4$

37. H_2SO_4

38. H_2CO_3

39. HCH_3CO_2

40. HNO_3

41. HBr

42. HI

Write the formulas for the following compounds:

43. Sodium chloride

44. Potassium sulfide

45. Calcium fluoride

46. Lead(IV) oxide

47. Lithium nitrate

48. Barium nitrite

49. Gold(II) sulfate

50. Chromium(I) sulfite

51. Manganese(II) bisulfate

52. Iron(III) carbonate

53. Cadmium(II) carbonate

54. Zinc(II) hydrogen carbonate

55. Silver(I) hydrogen carbonate

56. Aluminum perchlorate

57. Chromium(III) iodate

58. Titanium(IV) bromite

59. Zinc(II) hypoiodite

60. Tetraphosphorous decafluoride

61. Sulfur trioxide

62. Trinitrogen hexachloride

63. Carbonic acid

64. Sulfurous acid

65. Perchloric acid

66. Bromic acid

67. Bromous acid

68. Hypoiodous acid

69. Acetic acid

Answers

1. Sodium fluoride
2. Magnesium chloride
3. Calcium carbonate
4. Calcium hydrogen carbonate
5. Potassium hydrogen carbonate
6. Silver(I) nitrate
7. Zinc(II) nitrite
8. Cadmium(II) sulfate
9. Aluminum sulfite
10. Gold(II) phosphate
11. Copper(I) phosphite
12. Sodium chlorate
13. Barium chlorite
14. Silver hypochlorite
15. Iron(III) perchlorate
16. Cobalt(II) bromate
17. Manganese(II) bromite
18. Chromium(III) hypobromite
19. Zinc(II) perbromate
20. Silver(I) iodate
21. Aluminum iodite
22. Sodium hypoiodite
23. Lithium periodate
24. Titanium(II) bisulfite
25. Tin(IV) oxide
26. Dinitrogen monoxide
27. Carbon monoxide
28. Diphosphorus pentaoxide
29. Carbon tetrachloride
30. Sulfur trioxide
31. Diselenium hexafluoride
32. Hydrochloric acid
33. Chloric acid
34. Bromous acid
35. hypoiodous acid
36. Perchloric acid
37. Sulfuric acid
38. Carbonic acid
39. Acetic acid
40. Nitric acid
41. Hydrobromic acid
42. Hydroiodic acid
43. $NaCl$
44. K_2S
45. CaF_2
46. PbO_2
47. $LiNO_3$
48. $Ba(NO_2)_2$
49. $AuSO_4$
50. Cr_2SO_3
51. $Mn(HSO_4)_2$
52. $Fe_2(CO_3)_3$
53. $CdCO_3$
54. $Zn(HCO_3)_2$
55. $AgHCO_3$
56. $Al(ClO_4)_3$
57. $Cr((IO_3)_3$
58. $Ti(BrO_2)_4$
59. $Zn(IO)_2$
60. P_4F_{10}
61. SO_3
62. N_3Cl_6
63. H_2CO_3
64. H_2SO_3
65. $HClO_4$
66. $HBrO_3$
67. $HBrO_2$
68. HIO
69. HCH_3CO_2

CH 221 "Mass, Moles, Atoms" Study Questions

1. What is the molar mass of ammonium sulfate?

2. What is the molar mass of cobalt(II) iodide hexahydrate?

3. Calculate the number of moles in 0.41 g of titanium.

4. What is the mass of $1.0 * 10^9$ carbon atoms?

5. The density of carbon tetrachloride is 1.59 g/mL. How many Cl atoms are present in 55 mL of carbon tetrachloride?

6. The molar mass of cesium is 132.9 g/mol. What is the mass of a single Cs atom?

7. The density of lithium is 0.546 g/cm^3. What volume is occupied by 1.96×10^{23} atoms of Li?

8. What is the mass percentage of oxygen in acetic acid, HCH_3CO_2?

9. Which of the following could be an empirical formula? C_6H_{10}, B_4H_{10}, NO_3, $AsCl_5$.

10. Benzene has an empirical formula of CH. If the molar mass of benzene is 78.11 g/mol, what is the molecular formula for benzene?

11. Toluene is 91.25% C and 8.75% H. Determine the empirical formula for toluene. *Hint:* 8/7 = 1.14

12. The compound azulene is 93.71%C with the remainder hydrogen, and it has a molar mass of 128.16 g/mol. Calculate the empirical formula and molecular formula for azulene. *Hint:* 5/4 = 1.25

Answers appear on the next page

CH 221 "Mass, Moles, Atoms" Study Questions - *Answers*

1. 132.1 g/mol
2. 420.8 g/mol
3. 8.6×10^{-3} mol
4. 2.0×10^{-14} g
5. 1.4×10^{24} atoms
6. 2.207×10^{-22} g
7. 4.14 cm^3
8. 53.29%
9. NO_3 and $AsCl_5$ could be empirical formulas.
10. C_6H_6
11. C_7H_8
12. C_5H_4 (EF) and $C_{10}H_8$ (MF)

CH 221 Name: ____________________

Chemical Reactions Worksheet

Directions: Balance the following chemical reactions using the given information. In addition, ***classify*** each chemical reaction. **Answers appear immediately following the problems.**

---

1. Hypochlorous acid decomposes into water and dichlorine monoxide.

Reaction classification: _______________

---

2. Acetic acid is burned.

Reaction classification: _______________

---

3. Solid magnesium fluoride appears upon mixing magnesium chloride and sodium fluoride.

Reaction classification: _______________

---

4. Phosphorus (P_4) and oxygen produce tetraphosphorus decaoxide.

Reaction classification: _______________

---

5. Calcium and hydrochloric acid create a gas. Identify the gas through the balanced equation.

Reaction classification: _______________

---

6. Calcium hydroxide is added to perchloric acid..

Reaction classification: _______________

***Answers* to the Chemical Reactions Worksheet:**

---

1. Hypochlorous acid decomposes into water and dichlorine monoxide.

$2\ HClO(aq) \rightarrow H_2O(l) + Cl_2O(aq)$ *Classification:* **Decomposition**

2. Acetic acid is burned.

$HC_2H_3O_2(aq) + 2\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(l)$ *Classification:* **Combustion / Burning**

3. Solid magnesium fluoride appears upon mixing magnesium chloride and sodium fluoride.

$MgCl_2(aq) + 2\ NaF(aq) \rightarrow MgF_2(s) + 2\ NaCl(aq)$ *Classification:* **Precipitation**

4. Phosphorus (P_4) and oxygen produce tetraphosphorus decaoxide.

$P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$ *Classification:* **Combination**

5. Calcium and hydrochloric acid create a gas. Identify the gas through the balanced equation.

$Ca(s) + 2\ HBr(aq) \rightarrow CaBr_2(aq) + H_2(g)$ *Classification:* **Single Replacement**

6. Calcium hydroxide is added to perchloric acid..

$Ca(OH)_2(aq) + 2\ HClO_4(aq) \rightarrow Ca(ClO_4)_2(aq) + 2\ H_2O(l)$ *Classification:* **Acid/Base**

Chemical Reactions *answers follow at end*

Balancing Chemical Equations

1. What is the coefficient of oxygen gas after balancing the following equation?

$$\_\_P(s) + O_2(g) \rightarrow \_P_2O_3(s)$$

(a) 1
(b) 2
(c) 3
(d) 5
(e) none of the above

2. What is the coefficient of oxygen gas after balancing the following equation?

$$\_\_P(s) + \_\_O_2(g) \rightarrow \_\_P_2O_5(s)$$

(a) 1
(b) 2
(c) 4
(d) 5
(e) none of the above

3. What is the coefficient of phosphorus after balancing the following equation?

$$\_\_P(s) + O_2(g) \rightarrow \_P_2O_5(s)$$

(a) 1
(b) 2
(c) 4
(d) 5
(e) none of the above

4. What is the coefficient of nitrogen gas after balancing the following equation?

$$\_\_N_2(g) + \_\_H_2(g) \rightarrow \_\_NH_3(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

5. What is the coefficient of hydrogen gas after balancing the following equation?

$$\_\_N_2(g) \ + \ \_\_H_2(g) \rightarrow \_\_NH_3(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

6. What is the coefficient of ammonia gas after balancing the following equation?

$$\_\_N_2(g) \ + \ \_\_H_2(g) \rightarrow \_\_NH_3(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

7. What is the coefficient of chlorine gas after balancing the following equation?

$$\_\_Fe(s) \ + \ \_\_Cl_2(g) \rightarrow \_\_FeCl_3(s)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

8. What is the coefficient of carbon dioxide after balancing the following equation?

$$\_\_KHCO_3(s) \xrightarrow{\Delta} \_\_K_2CO_3(s) \ + \ \_\_H_2O(g) \ + \ \_\_CO_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

9. What is the coefficient of carbon dioxide after balancing the following equation?

$$\_\_Cr_2(CO_3)_3(s) \xrightarrow{\Delta} \_\_Cr_2O_3(s) + \_\_CO_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

10. What is the coefficient of oxygen gas after balancing the following equation?

$$\_\_AgClO_3(s) \xrightarrow{\Delta} \_AgCl(s) + O_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

11. What is the coefficient of oxygen gas after balancing the following equation?

$$\_\_LiNO_3(s) \xrightarrow{\Delta} \_\_LiNO_2(s) + \_\_O_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

12. What is the coefficient of oxygen gas after balancing the following equation?

$$\_\_HgO(s) \xrightarrow{\Delta} \_\_Hg(s) + \_\_O_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

13. What is the coefficient of oxygen gas after balancing the following equation?

$$\_\_H_2O_2(l) \xrightarrow{\Delta} \_H_2O(l) + O_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

14. What is the coefficient of silver metal after balancing the following equation?

$$\_\_Cu(s) + AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

15. What is the coefficient of Cd metal after balancing the following equation?

$$\_\_Al(s) + \_\_Cd(C_2H_3O_2)_2(aq) \rightarrow \_\_Al(C_2H_3O_2)_3(aq) + \_\_Cd(s)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

16. What is the coefficient of nickel metal after balancing the following equation?

$$\_\_Fe(s) + \_\_Ni(NO_3)_2(aq) \rightarrow \_\_Fe(NO_3)_3(aq) + \_\_Ni(s)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

17. What is the coefficient of hydrogen gas after balancing the following equation?

$$\_\_Pb(s) + \_\_HNO_3(aq) \rightarrow \_\_Pb(NO_3)_4(aq) + \_\_H_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

18. What is the coefficient of hydrogen gas after balancing the following equation?

$$\_\_Co(s) + \_\_HCl(aq) \rightarrow \_\_CoCl_3(aq) + \_\_H_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

19. What is the coefficient of hydrogen gas after balancing the following equation?

$$\_\_Mn(s) + \_\_H_2SO_4(aq) \rightarrow \_\_MnSO_4(aq) + \_\_H_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

20. What is the coefficient of sodium metal after balancing the following equation?

$$\_\_Na(s) + \_\_H_2O(l) \rightarrow \_\_NaOH(aq) + \_\_H_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

21. What is the coefficient of water after balancing the following equation?

$$\_\_Li(s) \quad + \quad \_\_H_2O(l) \rightarrow \_\_LiOH(aq) \quad + \quad \_\_H_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

22. What is the coefficient of hydrogen gas after balancing the following equation?

$$\_\_Ca(s) \quad + \quad \_\_H_2O(l) \rightarrow \quad \_\_Ca(OH)_2(aq) \quad + \quad \_\_H_2(g)$$

(a) 1
(b) 2
(c) 3
(d) 4
(e) none of the above

23. What is the coefficient of AgCl after balancing the following equation?

$$\_\_AlCl_3(aq) \quad + \quad \_\_AgNO_3(aq) \rightarrow \quad \_\_Al(NO_3)_3(aq) \quad + \quad \_\_AgCl(s)$$

(a) 1
(b) 2
(c) 3
(d) 6
(e) none of the above

24. What is the coefficient of NaCl after balancing the following equation?

$$\_\_CrCl_3(aq) \quad + \quad \_\_Na_2CO_3(aq) \rightarrow \quad \_\_Cr_2(CO_3)_3(s) \quad + \quad \_\_NaCl(aq)$$

(a) 1
(b) 2
(c) 3
(d) 6
(e) none of the above

25. What is the coefficient of KNO_3 after balancing the following equation?

$$\_\_Au(NO_3)_3(aq) + \quad K_2CrO_4(aq) \rightarrow \_Au_2(CrO_4)_3(s) + \quad KNO_3(aq)$$

(a) 1
(b) 2
(c) 3
(d) 6
(e) none of the above

26. What is the coefficient of water after balancing the following equation?

$$\_\_H_2SO_4(aq) + \_\_NaOH(aq) \rightarrow \_\_Na_2SO_4(aq) + \_\_H_2O(l)$$

(a) 1
(b) 2
(c) 3
(d) 6
(e) none of the above

27. What is the coefficient of water after balancing the following equation?

$$\_\_HC_2H_3O_2(aq) + \_\_Ca(OH)_2(aq) \rightarrow \_\_Ca(C_2H_3O_2)_2(aq) + \_\_H_2O(l)$$

(a) 1
(b) 2
(c) 3
(d) 6
(e) none of the above

28. What is the coefficient of water after balancing the following equation?

$$\_\_H_3PO_4(aq) + \_\_Ba(OH)_2(aq) \rightarrow \_\_Ba_3(PO_4)_2(s) + \_\_H_2O(l)$$

(a) 1
(b) 2
(c) 3
(d) 6
(e) none of the above

Classifying Chemical Reactions

29. Which of the following types of chemical reactions is illustrated below?

$$N_2(g) \quad + \quad H_2(g) \rightarrow \quad NH_3(g)$$

(a) combination
(b) decomposition
(c) single replacement
(d) double replacement
(e) neutralization

30. Which of the following types of chemical reactions is illustrated below?

$$SO_2(g) \quad + \quad O_2(g) \rightarrow \quad SO_3(g)$$

(a) combination
(b) decomposition
(c) single replacement
(d) double replacement
(e) neutralization

31. Which of the following types of chemical reactions is illustrated below?

$$KHCO_3(s) \xrightarrow{\Delta} K_2CO_3(s) \quad + \quad H_2O(g) \quad + \quad CO_2(g)$$

(a) combination
(b) decomposition
(c) single replacement
(d) double replacement
(e) neutralization

32. Which of the following types of chemical reactions is illustrated below?

$$KClO_3(s) \xrightarrow{\Delta} KCl(s) \quad + \quad O_2(g)$$

(a) combination
(b) decomposition
(c) single replacement
(d) double replacement
(e) neutralization

33. Which of the following types of chemical reactions is illustrated below?

$$Zn(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

(a) combination
(b) decomposition
(c) single replacement
(d) double replacement
(e) neutralization

34. Which of the following types of chemical reactions is illustrated below?

$$Sr(s) + H_2O(l) \rightarrow Sr(OH)_2(aq) + H_2(g)$$

(a) combination
(b) decomposition
(c) single replacement
(d) double replacement
(e) neutralization

35. Which of the following types of chemical reactions is illustrated below?

$$AlCl_3(aq) + AgNO_3(aq) \rightarrow Al(NO_3)_3(aq) + AgCl(s)$$

(a) combination
(b) decomposition
(c) single replacement
(d) double replacement
(e) neutralization

36. Which of the following types of chemical reactions is illustrated below?

$$FeBr_3(aq) + AgNO_3(aq) \rightarrow Fe(NO_3)_3(aq) + AgBr(s)$$

(a) combination
(b) decomposition
(c) single replacement
(d) double replacement
(e) neutralization

37. Which of the following types of chemical reactions is illustrated below?

$$H_2SO_4(aq) + NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$$

(a) combination
(b) decomposition
(c) single replacement
(d) double replacement
(e) neutralization

38. Which of the following types of chemical reactions is illustrated below?

$$HClO_4(aq) + KOH(aq) \rightarrow KClO_4(aq) + H_2O(l)$$

(a) combination
(b) decomposition
(c) single replacement
(d) double replacement
(e) neutralization

Combination Reactions

39. What is the predicted product from the following combination reaction?

$$Li(s) + O_2(g) \xrightarrow{\Delta}$$

(a) LiO
(b) Li_2O
(c) LiO_2
(d) Li_2O_3
(e) Li_3O_2

40. What is the predicted product from the following combination reaction?

$$Ca(s) + O_2(g) \xrightarrow{\Delta}$$

(a) CaO
(b) Ca_2O
(c) CaO_2
(d) Ca_2O_3
(e) Ca_3O_2

41. What is the predicted product from the following combination reaction?

$$Al(s) + O_2(g) \xrightarrow{\Delta}$$

(a) AlO
(b) Al_2O
(c) AlO_2
(d) Al_2O_3
(e) Al_3O_2

42. What is the predicted product from the following combination reaction?

$$K(s) + Cl_2(g) \xrightarrow{\Delta}$$

(a) KCl
(b) K_2Cl
(c) KCl_2
(d) KCl_3
(e) K_3Cl

43. What is the predicted product from the following combination reaction?

$$Sr(s) + Br_2(l) \xrightarrow{\Delta}$$

(a) $SrBr$
(b) Sr_2Br
(c) $SrBr_2$
(d) Sr_2Br_3
(e) Sr_3Br_2

44. What is the predicted product from the following combination reaction?

$$Zn(s) + I_2(s) \xrightarrow{\Delta}$$

(a) ZnI
(b) ZnI_2
(c) Zn_2I
(d) Zn_2I_3
(e) Zn_3I_2

45. What is the formula of the predicted product from heating magnesium metal and nitrogen gas?
   (a) MgN
   (b) MgN_2
   (c) Mg_2N
   (d) Mg_2N_3
   (e) Mg_3N_2

46. What is the formula of the predicted product from heating potassium metal and powdered phosphorus?
   (a) KP
   (b) KP_3
   (c) K_3P
   (d) K_2P_3
   (e) K_3P_2

47. What is the formula of the predicted product from heating cadmium metal and powdered sulfur?
   (a) CdS
   (b) Cd_2S
   (c) CdS_2
   (d) Cd_2S_3
   (e) Cd_3S_2

Decomposition Reactions

48. What are the predicted products from the following decomposition reaction?

$$LiHCO_3(s) \xrightarrow{\Delta}$$

   (a) Li, H_2, and CO_2
   (b) Li, H_2O, and CO_2
   (c) Li_2CO_3, H_2, and CO_2
   (d) Li_2CO_3, H_2O, and CO_2
   (e) Li_2CO_3 and H_2O

49. What are the predicted products from the following decomposition reaction?

$$Zn(HCO_3)_2(s) \xrightarrow{\Delta}$$

(a) Zn, H_2, and CO_2
(b) Zn, H_2O, and CO_2
(c) $ZnCO_3$, H_2, and CO_2
(d) $ZnCO_3$, H_2O, and CO_2
(e) $ZnCO_3$ and H_2O

50. What are the predicted products from the following decomposition reaction?

$$Al(HCO_3)_3(s) \xrightarrow{\Delta}$$

(a) Al, H_2, and CO_2
(b) Al, H_2O, and CO_2
(c) $Al_2(CO_3)_3$, H_2, and CO_2
(d) $Al_2(CO_3)_3$, H_2O, and CO_2
(e) $Al_2(CO_3)_3$ and H_2O

51. What are the predicted products from the following decomposition reaction?

$$Fe(HCO_3)_3(s) \xrightarrow{\Delta}$$

(a) Fe, H_2O, and CO_2
(b) $FeCO_3$, H_2, and CO_2
(c) $FeCO_3$, H_2O, and CO_2
(d) $Fe_2(CO_3)_3$, H_2O, and CO_2
(e) $Fe_2(CO_3)_3$, H_2, and CO_2

52. What are the predicted products from the following decomposition reaction?

$$Cu_2CO_3(s) \xrightarrow{\Delta}$$

(a) Cu and CO_2
(b) Cu_2O and CO
(c) Cu_2O and CO_2
(d) CuO and CO
(e) CuO and CO_2

53. What are the predicted products from the following decomposition reaction?

$$PbCO_3(s) \xrightarrow{\Delta}$$

(a) Pb and CO_2
(b) PbO and CO
(c) PbO and CO_2
(d) PbO_2 and CO
(e) PbO_2 and CO_2

54. What are the predicted products from the following decomposition reaction?

$$Fe_2(CO_3)_3(s) \xrightarrow{\Delta}$$

(a) Fe and CO_2
(b) FeO and CO
(c) FeO and CO_2
(d) Fe_2O_3 and CO
(e) Fe_2O_3 and CO_2

55. What are the predicted products from the following decomposition reaction?

$$NaClO_3(s) \xrightarrow{\Delta}$$

(a) Na and CO_2
(b) Na, Cl_2, and O_2
(c) NaCl and H_2O
(d) NaCl and O_2
(e) NaCl and CO_2

56. What are the predicted products from the following decomposition reaction?

$$Zn(ClO_3)_2(s) \xrightarrow{\Delta}$$

(a) Zn and CO_2
(b) Zn, Cl_2, and O_2
(c) $ZnCl_2$ and H_2O
(d) $ZnCl_2$ and O_2
(e) $ZnCl_2$ and CO_2

57. What are the predicted products from the following decomposition reaction?

$$Al(ClO_3)_3(s) \xrightarrow{\Delta}$$

(a) Al and CO_2
(b) Al, Cl_2, and O_2
(c) $AlCl_3$ and H_2O
(d) $AlCl_3$ and O_2
(e) $AlCl_3$ and CO_2

Single-Replacement Reactions

58. What are the products from the following single–replacement reaction?

$$Zn(s) + CuSO_4(aq) \rightarrow$$

(a) Cu and $ZnSO_4$
(b) Cu and $ZnSO_3$
(c) CuO and $ZnSO_4$
(d) CuO and $ZnSO_3$
(e) no reaction

59. What are the products from the following single–replacement reaction?

$$Cd(s) + AgNO_3(aq) \rightarrow$$

(a) Ag and $Cd(NO_3)_2$
(b) Ag and $Cd(NO_2)_2$
(c) Ag_2O and $Cd(NO_3)_2$
(d) Ag_2O and $Cd(NO_2)_2$
(e) no reaction

60. What are the products from the following single–replacement reaction?

$$Al(s) + Pb(NO_3)_2(aq) \rightarrow$$

(a) Pb and $Al(NO_3)_3$
(b) Pb and $Al(NO_2)_3$
(c) PbO and $Al(NO_3)_3$
(d) PbO and $Al(NO_2)_3$
(e) no reaction

61. What are the products from the following single–replacement reaction?

$$Mg(s) + H_2SO_4(aq) \rightarrow$$

(a) MgO and H_2SO_3
(b) MgO and H_2S
(c) $MgSO_4$ and H_2
(d) $MgSO_4$ and H_2O
(e) no reaction

62. What are the products from the following single–replacement reaction?

$$Zn(s) + HNO_3(aq) \rightarrow$$

(a) ZnO and HNO_2
(b) $Zn(NO_2)_2$ and H_2
(c) $Zn(NO_3)_2$ and H_2
(d) $Zn(NO_3)_2$ and H_2O
(e) no reaction

63. What are the products from the following single–replacement reaction?

$$K(s) + H_2O(l) \rightarrow$$

(a) K_2O and H_2
(b) K_2O and H_2O
(c) KOH and H_2
(d) KOH and H_2O
(e) no reaction

64. What are the products from the following single–replacement reaction?

$$Ba(s) + H_2O(l) \rightarrow$$

(a) BaO and H_2
(b) BaO and H_2O
(c) $Ba(OH)_2$ and H_2
(d) $Ba(OH)_2$ and H_2O
(e) no reaction

Solubility Rules

65. Which of the following solid compounds is soluble in water?
    (a) Na_2CO_3
    (b) $CuC_2H_3O_2$
    (c) $AgNO_3$
    (d) all of the above
    (e) none of the above

66. Which of the following solid compounds is soluble in water?
    (a) $CaCO_3$
    (b) $PbSO_4$
    (c) $AlPO_4$
    (d) all of the above
    (e) none of the above

67. Which of the following solid compounds is soluble in water?
    (a) $NiCO_3$
    (b) $PbCrO_4$
    (c) Ag_3PO_4
    (d) CuS
    (e) $Ba(OH)_2$

68. Which of the following solid compounds is insoluble in water?
    (a) $PbCl_2$
    (b) Hg_2I_2
    (c) $BaSO_4$
    (d) all of the above
    (e) none of the above

69. Which of the following solid compounds is insoluble in water?
    (a) Li_2CO_3
    (b) $AgC_2H_3O_2$
    (c) $Cu(NO_3)_2$
    (d) all of the above
    (e) none of the above

70. Which of the following solid compounds is insoluble in water?
   (a) $(NH_4)_2CO_3$
   (b) K_2CrO_4
   (c) $BaSO_4$
   (d) Na_2S
   (e) $Sr(OH)_2$

Double-Replacement Reactions

71. What are the products from the following double–replacement reaction?

$$AgNO_3(aq) \ + \ NaCl(aq) \rightarrow$$

   (a) Ag_3N and $NaClO_3$
   (b) $AgCl$ and $NaNO_2$
   (c) $AgCl$ and $NaNO_3$
   (d) $AgClO_3$ and $NaNO_2$
   (e) $AgClO_3$ and $NaNO_3$

72. What are the products from the following double–replacement reaction?

$$BaCl_2(aq) \ + \ K_2SO_4(aq) \rightarrow$$

   (a) BaS and $KClO_4$
   (b) $BaSO_3$ and KCl
   (c) $BaSO_3$ and $KClO_4$
   (d) $BaSO_4$ and KCl
   (e) $BaSO_4$ and $KClO_4$

73. What are the products from the following double–replacement reaction?

$$AgNO_3(aq) \ + \ Li_3PO_4(aq) \rightarrow$$

   (a) Ag_3P and $LiNO_3$
   (b) Ag_3PO_3 and $LiNO_2$
   (c) Ag_3PO_3 and $LiNO_3$
   (d) Ag_3PO_4 and $LiNO_2$
   (e) Ag_3PO_4 and $LiNO_3$

Neutralization Reactions

74. What are the predicted products from the following neutralization reaction?

$$HCl(aq) \quad + \quad NH_4OH(aq) \rightarrow$$

(a) NH_3Cl and H_2O
(b) NH_3Cl and O_2
(c) NH_4Cl and H_2O
(d) NH_4Cl and O_2
(e) no reaction

75. What are the predicted products from the following neutralization reaction?

$$HC_2H_3O_2(aq) \quad + \quad Ca(OH)_2(aq) \rightarrow$$

(a) $CaCO_3$ and H_2O
(b) $Ca(HCO_3)_2$ and H_2
(c) $Ca(HCO_3)_2$ and H_2O
(d) $Ca(C_2H_3O_2)_2$ and H_2
(e) $Ca(C_2H_3O_2)_2$ and H_2O

76. What are the predicted products from the following neutralization reaction?

$$HNO_3(aq) \quad + \quad Ba(OH)_2(aq) \rightarrow$$

(a) Ba_3N_2 and H_2O
(b) $Ba(NO_2)_2$ and H_2
(c) $Ba(NO_2)_2$ and H_2O
(d) $Ba(NO_3)_2$ and H_2
(e) $Ba(NO_3)_2$ and H_2O

77. What are the products from the complete neutralization of sulfuric acid with aqueous sodium hydroxide?

(a) $Na_2S(aq)$ and $H_2O(l)$
(b) $NaHSO_3(aq)$ and $H_2O(l)$
(c) $NaHSO_4(aq)$ and $H_2O(l)$
(d) $Na_2SO_3(aq)$ and $H_2O(l)$
(e) $Na_2SO_4(aq)$ and $H_2O(l)$

78. What are the products from the complete neutralization of carbonic acid with aqueous potassium hydroxide?
   (a) K_2CO_3(aq) and H_2O(l)
   (b) $KHCO_3$(aq) and H_2O(l)
   (c) $KHCO_4$(aq) and H_2O(l)
   (d) $KC_2H_3O_2$(aq) and H_2O(l)
   (e) $K_2C_2H_3O_2$(aq) and H_2O(l)

79. What are the products from the complete neutralization of phosphoric acid with aqueous lithium hydroxide?
   (a) LiH_2PO_4(aq) and H_2O(l)
   (b) Li_2HPO_4(aq) and H_2O(l)
   (c) Li_3PO_4(aq) and H_2O(l)
   (d) $LiHPO_4$(aq) and H_2O(l)
   (e) Li_2PO_4(aq) and H_2O(l)

Combustion Reactions

80. Methane, CH_4, can be used as fuel in an automobile to reduce pollution. What is the coefficient of oxygen in the balanced equation for the reaction?

$$\_\_CH_4(g) + \_\_O_2(g) \xrightarrow{\text{spark}} \_\_CO_2(g) + H_2O(g)$$

   (a) 1
   (b) 2
   (c) 3
   (d) 4
   (e) none of the above

81. Ethane, C_2H_6, burns to give carbon dioxide and water. What is the coefficient of oxygen in the balanced equation for the reaction?

$$\_\_C_2H_6(g) + \_\_O_2(g) \xrightarrow{\text{spark}} \_\_CO_2(g) + \_H_2O(g)$$

   (a) 5
   (b) 7
   (c) 10
   (d) 14
   (e) none of the above

82. Propane, C_3H_8, is flammable and used in rural areas where natural gas is not available. What is the coefficient of oxygen in the balanced equation for the combustion of propane?

$$\_\_C_3H_8(g) + \_\_O_2(g) \xrightarrow{\text{spark}} \_\_CO_2(g) + H_2O(g)$$

(a) 1
(b) 5
(c) 7
(d) 10
(e) none of the above

83. Butane, C_4H_{10}, is flammable and used in butane lighters. What is the coefficient of oxygen in the balanced equation for the combustion of butane?

$$\_\_C_4H_{10}(g) + \_\_O_2(g) \xrightarrow{\text{spark}} \_\_CO_2(g) + H_2O(g)$$

(a) 9
(b) 13
(c) 18
(d) 26
(e) none of the above

84. Octane, C_8H_{18}, is a major component in gasoline. What is the coefficient of oxygen in the balanced equation for the combustion of octane?

$$\_\_C_8H_{18}(g) + \_\_O_2(g) \xrightarrow{\text{spark}} \_\_CO_2(g) + H_2O(g)$$

(a) 17
(b) 25
(c) 34
(d) 50
(e) none of the above

85. Ethanol, C_2H_5OH, is made from fermenting grain and can be blended with gasoline to make "gasohol." If the combustion of "gasohol" produces carbon dioxide and water, what is the coefficient of oxygen in the balanced equation?

$$\_\_C_2H_5OH(g) \quad + \quad \_\_O_2(g) \quad \xrightarrow{\text{spark}} \quad \_\_CO_2(g) \quad + \quad \_H_2O(g)$$

(a) 1
(b) 2
(c) 3
(d) 6
(e) none of the above

86. Methanol, CH_3OH, is derived from natural gas and can be blended with gasoline to make "gasohol." If the combustion of "gasohol" produces carbon dioxide and water, what is the coefficient of oxygen in the balanced equation?

$$\_\_CH_3OH(g) \quad + \quad \_\_O_2(g) \xrightarrow{\text{spark}} \quad \_\_CO_2(g) \quad + \quad H_2O(g)$$

(a) 1
(b) 2
(c) 3
(d) 6
(e) none of the above

Answer Key

1. C
2. D
3. C
4. A
5. C
6. B
7. C
8. A
9. C
10. C
11. A
12. A
13. A
14. B
15. C
16. C
17. B
18. C
19. A
20. B
21. B
22. A
23. C
24. D
25. D
26. B
27. B
28. D
29. A
30. A
31. B
32. B
33. C
34. C
35. D
36. D
37. E
38. E
39. B
40. A
41. D
42. A
43. C
44. B
45. E
46. C
47. A
48. D
49. D
50. D
51. D
52. C
53. C
54. E
55. D
56. D
57. D
58. A
59. A
60. A
61. C
62. C
63. C
64. C
65. D
66. E
67. E
68. D
69. E
70. C
71. C
72. D
73. E
74. C
75. E
76. E
77. E
78. A
79. C
80. B
81. B
82. B
83. B
84. B
85. C
86. C

CH 221 Limiting Reactant Example

Hexane (C_6H_{14}) burns in air (O_2) to give CO_2 and H_2O.

- Write a balanced equation for this reaction.
- If 215 g of C_6H_{14} is mixed with 215 g of O_2, what masses of CO_2 and H_2O are produced in the reaction?
- What mass of excess reactant remains at the end of the reaction?
- If 151.3 g of CO_2 are collected, what is the percent yield of CO_2?

$2\ C_6H_{14}(\ell) + 19\ O_2(g) \rightarrow 12\ CO_2(g) + 14\ H_2O(g)$

215 g C_6H_{14} * (mol/86.18 g) * (12 mol CO_2 / 2 mol C_6H_{14}) * 44.01 g/mol = 658 g CO_2

215 g O_2 * (mol/32.00 g) * (12 mol CO_2 / 19 mol O_2) * 44.01 g/mol = 187 g CO_2 **(Theo. yield)**

Excess Reactant = C_6H_{14}, ***Limiting Reactant = O_2***

215 g O_2 * (mol/32.00 g) $\cdot \dfrac{12 \text{ mol } CO_2}{19 \text{ mol } O_2} \cdot \dfrac{44.01 \text{ g}}{1 \text{ mol } CO_2}$ = **187 g CO_2**

215 g O_2 * (mol/32.00 g) $\cdot \dfrac{14 \text{ mol } H_2O}{19 \text{ mol } O_2} \cdot \dfrac{18.02 \text{ g}}{1 \text{ mol } H_2O}$ = **89.2 g H_2O**

215 g O_2 * (mol/32.00 g) $\cdot \dfrac{2 \text{ mol } C_6H_{14}}{19 \text{ mol } O_2} \cdot \dfrac{86.18 \text{ g}}{1 \text{ mol } C_6H_{14}}$ = **60.9 g C_6H_{14} used**

215 g C_6H_{14} available – 60.9 g C_6H_{14} used = **154 g C_6H_{14} remains**

%yield = (151.3 / 187) * 100% = **80.9% CO_2**

Try it yourself:

Calcium oxide and ammonium chloride can be combined to give ammonia (NH_3), water and calcium chloride.

- Write a balanced equation for this reaction.
- If 112 g of calcium oxide is mixed with 224 g of ammonium chloride, what mass of NH_3 should be produced in the reaction?
- What mass of excess reactant remains at the end of the reaction?
- If only 16.3 g of NH_3 are collected, what is the percent yield of NH_3?

Answers appear on the next page.

CH 221 Limiting Reactant Example - *Answers*

Calcium oxide and ammonium chloride can be combined to give ammonia (NH_3), water and calcium chloride.

- Write a balanced equation for this reaction.

  $\mathbf{CaO(s) + 2\ NH_4Cl(aq) \rightarrow 2\ NH_3(g) + H_2O(g) + CaCl_2(s)}$

- If 112 g of calcium oxide is mixed with 224 g of ammonium chloride, what mass of NH_3 should be produced in the reaction?

  Theoretical yield of NH_3 = 68.0 g

- What mass of excess reactant remains at the end of the reaction?

  10. g of excess reactant remains at the end of the reaction.

- If only 16.3 g of NH_3 are collected, what is the percent yield of NH_3?

  Percent yield = 24.0%

Moles, Mass, and Limiting Reactants

Answers follow at end

Interpreting a Chemical Equation

1. How many moles of chlorine gas react with 1 mol of hydrogen gas according to the balanced chemical equation?

$$H_2(g) \quad + \quad Cl_2(g) \quad \rightarrow \quad 2\ HCl(g)$$

   (a) 1 mol
   (b) 2 mol
   (c) 3 mol
   (d) 4 mol
   (e) none of the above

2. Assuming similar conditions, how many liters of chlorine gas react to produce 2 L of hydrogen chloride gas?

$$H_2(g) \quad + \quad Cl_2(g) \quad \rightarrow \quad 2\ HCl(g)$$

   (a) 1 L
   (b) 2 L
   (c) 3 L
   (d) 4 L
   (e) none of the above

3. How many moles of carbon monoxide react with 1 mol of oxygen gas according to the balanced chemical equation?

$$2\ CO(g) \quad + \quad O_2(g) \quad \xrightarrow{\Delta} \quad 2\ CO_2(g)$$

   (a) 1 mol
   (b) 2 mol
   (c) 3 mol
   (d) 4 mol
   (e) none of the above

4. Assuming similar conditions, how many liters of carbon monoxide gas react to produce 2 L of carbon dioxide gas?

$$2\,CO(g) \quad + \quad O_2(g) \quad \xrightarrow{\Delta} \quad 2\,CO_2(g)$$

(a) 1 L
(b) 2 L
(c) 3 L
(d) 4 L
(e) none of the above

5. How many moles of water vapor, H_2O, react with 1 mol of charcoal, C, according to the balanced chemical equation?

$$C(s) \quad + \quad H_2O(g) \quad \xrightarrow{\Delta} \quad CO(g) \quad + \quad H_2(g)$$

(a) 1 mol
(b) 2 mol
(c) 3 mol
(d) 4 mol
(e) none of the above

6. Assuming similar conditions, how many liters of water vapor, H_2O, react to produce 1 L of hydrogen gas?

$$C(s) \quad + \quad H_2O(g) \quad \xrightarrow{\Delta} \quad CO(g) \quad + \quad H_2(g)$$

(a) 1 L
(b) 2 L
(c) 3 L
(d) 4 L
(e) none of the above

7. In an experiment, 5.585 g of iron metal reacts with 3.207 g of yellow sulfur. Using the conservation of mass law, predict the mass of product.

$$Fe(s) \quad + \quad S(s) \quad \xrightarrow{\Delta} \quad FeS(s)$$

(a) 2.198 g
(b) 2.378 g
(c) 4.396 g
(d) 8.792 g
(e) 17.584 g

8. In an experiment, 1.201 g of charcoal reacts with 6.414 g of powdered sulfur. Using the conservation of mass law, predict the mass of product.

$$C(s) + 2\,S(s) \xrightarrow{\Delta} CS_2(g)$$

(a) 4.408 g
(b) 5.213 g
(c) 7.615 g
(d) 8.816 g
(e) 14.029 g

9. In an experiment, 0.520 g of chromium metal reacts with 3.807 g of iodine. Using the conservation of mass law, predict the mass of product.

$$2\,Cr(s) + 3\,I_2(s) \xrightarrow{\Delta} 2\,CrI_3(s)$$

(a) 1.529 g
(b) 3.287 g
(c) 4.327 g
(d) 8.654 g
(e) 12.461 g

10. In an experiment, 0.243 g of magnesium reacts to give 0.403 g magnesium oxide. Using the conservation of mass law, predict the mass of reacting oxygen gas.

$$2\,Mg(s) + O_2(g) \xrightarrow{\Delta} 2\,MgO(s)$$

(a) 0.080 g
(b) 0.160 g
(c) 0.320 g
(d) 0.646 g
(e) 1.292 g

11. In an experiment, 0.327 g of zinc metal reacts to produce 0.407 g of zinc oxide. Using the conservation of mass law, predict the mass of reacting oxygen gas.

$$2\,Zn(s) + O_2(g) \xrightarrow{\Delta} 2\,ZnO(s)$$

(a) 0.040 g
(b) 0.080 g
(c) 0.160 g
(d) 0.734 g
(e) 1.468 g

12. In an experiment, 0.197 g of gold metal reacts to yield 0.303 g of gold(III) chloride. Using the conservation of mass law, predict the mass of reacting chlorine gas.

$$2\ Au(s) \quad + \quad 3\ Cl_2(g) \quad \xrightarrow{\Delta} \quad 2\ AuCl_3(s)$$

(a) 0.035 g
(b) 0.106 g
(c) 0.318 g
(d) 0.167 g
(e) 0.500 g

Mole–Mole Relationships

13. How many moles of iodine vapor react with 1.00 mol of hydrogen gas?

$$\_\_H_2(g) \quad + \quad \_\_I_2(g) \quad \xrightarrow{\Delta} \quad \_\_HI(g)$$

(a) 0.500 mol
(b) 1.00 mol
(c) 2.00 mol
(d) 4.00 mol
(e) none of the above

14. How many moles of hydrogen iodide are produced from 1.00 mol of iodine?

$$\_\_H_2(g) \quad + \quad \_\_I_2(g) \quad \xrightarrow{\Delta} \quad \_\_HI(g)$$

(a) 0.500 mol
(b) 1.00 mol
(c) 2.00 mol
(d) 4.00 mol
(e) none of the above

15. How many moles of hydrogen gas react to yield 1.00 mol of hydrogen iodide?

$$\_\_H_2(g) \quad + \quad \_\_I_2(g) \quad \xrightarrow{\Delta} \quad \_\_HI(g)$$

(a) 0.500 mol
(b) 1.00 mol
(c) 2.00 mol
(d) 4.00 mol
(e) none of the above

16. How many moles of water are produced from 1.00 mol of hydrogen peroxide?

$$\_\_H_2O_2(l) \xrightarrow{\Delta} \_\_H_2O(l) + \_\_O_2(g)$$

(a) 0.500 mol
(b) 1.00 mol
(c) 2.00 mol
(d) 4.00 mol
(e) none of the above

17. How many moles of oxygen are produced from 1.00 mol of hydrogen peroxide?

$$\_\_H_2O_2(l) \xrightarrow{\Delta} \_\_H_2O(l) + \_\_O_2(g)$$

(a) 0.500 mol
(b) 1.00 mol
(c) 2.00 mol
(d) 4.00 mol
(e) none of the above

18. How many moles of hydrogen peroxide decompose to give 1.00 mol of oxygen?

$$\_\_H_2O_2(l) \xrightarrow{\Delta} \_\_H_2O(l) + \_\_O_2(g)$$

(a) 0.500 mol
(b) 1.00 mol
(c) 2.00 mol
(d) 4.00 mol
(e) none of the above

19. How many moles of oxygen gas react with 1.00 mol NO?

$$\_\_NO(g) + O_2(g) \xrightarrow{UV} \_\_NO_2(g)$$

(a) 0.250 mol
(b) 0.500 mol
(c) 1.00 mol
(d) 2.00 mol
(e) none of the above

20. How many moles of nitrogen dioxide gas are produced from 1.00 mol NO?

$$\_\_NO(g) + O_2(g) \xrightarrow{UV} \_\_NO_2(g)$$

(a) 0.250 mol
(b) 0.500 mol
(c) 1.00 mol
(d) 2.00 mol
(e) none of the above

21. How many moles of oxygen gas react to yield 1.00 mol NO_2?

$$\_\_NO(g) + O_2(g) \xrightarrow{UV} \_\_NO_2(g)$$

(a) 0.250 mol
(b) 0.500 mol
(c) 1.00 mol
(d) 2.00 mol
(e) none of the above

22. How many moles of water react with 5.00 mol of lithium metal?

$$\_\_Li(s) + H_2O(l) \rightarrow \_LiOH(aq) + H_2(g)$$

(a) 2.50 mol
(b) 5.00 mol
(c) 10.0 mol
(d) 20.0 mol
(e) none of the above

23. How many moles of hydrogen gas are produced from 5.00 mol of water?

$$\_\_Li(s) + H_2O(l) \rightarrow \_LiOH(aq) + H_2(g)$$

(a) 2.50 mol
(b) 5.00 mol
(c) 10.0 mol
(d) 20.0 mol
(e) none of the above

24. How many moles of lithium metal react to yield 5.00 mol of hydrogen gas?

$$\_\_Li(s) \quad + \quad H_2O(l) \quad \rightarrow \quad \_LiOH(aq) \quad + \quad H_2(g)$$

(a) 2.50 mol
(b) 5.00 mol
(c) 10.0 mol
(d) 20.0 mol
(e) none of the above

25. How many moles of water react with 0.500 mol of calcium metal?

$$\_\_Ca(s) \quad + \quad H_2O(l) \quad \rightarrow \quad \_Ca(OH)_2(aq) \quad + \quad H_2(g)$$

(a) 0.250 mol
(b) 0.500 mol
(c) 1.00 mol
(d) 2.00 mol
(e) none of the above

26. How many moles of hydrogen gas are produced from 0.500 mol of water?

$$\_\_Ca(s) \quad + \quad H_2O(l) \quad \rightarrow \quad \_Ca(OH)_2(aq) \quad + \quad H_2(g)$$

(a) 0.250 mol
(b) 0.500 mol
(c) 1.00 mol
(d) 2.00 mol
(e) none of the above

27. How many moles of calcium metal react to yield 0.500 mol of hydrogen gas?

$$\_\_Ca(s) \quad + \quad H_2O(l) \quad \rightarrow \quad \_Ca(OH)_2(aq) \quad + \quad H_2(g)$$

(a) 0.250 mol
(b) 0.500 mol
(c) 1.00 mol
(d) 2.00 mol
(e) none of the above

28. How many moles of oxygen gas react with 0.100 mol of pentane, C_5H_{12}?

$$\_\_C_5H_{12}(g) + \_O_2(g) \xrightarrow{\text{spark}} \_\_CO_2(g) + \_\_H_2O(g)$$

(a) 0.100 mol
(b) 0.500 mol
(c) 0.600 mol
(d) 0.800 mol
(e) none of the above

29. How many moles of water are produced from 0.100 mol pentane, C_5H_{12}?

$$\_\_C_5H_{12}(g) + \_O_2(g) \xrightarrow{\text{spark}} \_\_CO_2(g) + \_\_H_2O(g)$$

(a) 0.100 mol
(b) 0.500 mol
(c) 0.600 mol
(d) 0.800 mol
(e) none of the above

30. How many moles of oxygen gas react to yield 0.100 mol water?

$$\_\_C_5H_{12}(g) + \_O_2(g) \xrightarrow{\text{spark}} \_\_CO_2(g) + \_\_H_2O(g)$$

(a) 0.100 mol
(b) 0.500 mol
(c) 0.600 mol
(d) 0.800 mol
(e) none of the above

Mass–Mole–Mole–Mass Problems

31. What is the mass of silver metal produced from 6.35 g of copper?

$$\_\_Cu(s) + \_AgNO_3(aq) \rightarrow \_Cu(NO_3)_2(aq) + Ag(s)$$

(a) 0.187 g
(b) 0.540 g
(c) 0.747 g
(d) 1.08 g
(e) 21.6 g

32. What is the mass of copper metal that yields 0.500 g of silver?

$$\_\_Cu(s) \quad + \quad \_AgNO_3(aq) \quad \rightarrow \quad \_Cu(NO_3)_2(aq) \quad + \quad Ag(s)$$

(a) 0.147 g
(b) 0.294 g
(c) 0.425 g
(d) 0.589 g
(e) 1.70 g

33. What is the mass of silver bromide (187.77 g/mol) precipitated from 2.96 g of iron(III) bromide (295.55 g/mol)?

$$\_\_FeBr_3(s) \quad + \quad AgNO_3(aq) \quad \rightarrow \quad \_AgBr(s) \quad + \quad Fe(NO_3)_3(aq)$$

(a) 0.940 g
(b) 0.627 g
(c) 1.88 g
(d) 5.64 g
(e) 3.76 g

34. What is the mass of iron(III) bromide (295.55 g/mol) that yields 0.188 g of silver bromide (187.77 g/mol) precipitate?

$$\_\_FeBr_3(s) \quad + \quad AgNO_3(aq) \quad \rightarrow \quad \_AgBr(s) \quad + \quad Fe(NO_3)_3(aq)$$

(a) 0.0986 g
(b) 0.148 g
(c) 0.296 g
(d) 0.592 g
(e) 0.888 g

35. What is the mass of lead(II) iodide (461.0 g/mol) precipitated from 0.830 g of potassium iodide (166.00 g/mol)?

$$\_\_Pb(NO_3)_2(aq) \quad + \quad \_KI(s) \quad \rightarrow \quad \_PbI_2(s) \quad + \quad \_KNO_3(aq)$$

(a) 0.149 g
(b) 0.598 g
(c) 1.15 g
(d) 2.31 g
(e) 4.61 g

36. What is the mass of potassium iodide (166.00 g/mol) that yields 0.500 g of lead(II) iodide (461.0 g/mol) precipitate?

$$\_\_Pb(NO_3)_2(aq) \quad + \quad \_KI(s) \quad \rightarrow \quad \_PbI_2(s) \quad + \quad \_KNO_3(aq)$$

(a) 0.0900 g
(b) 0.180 g
(c) 0.360 g
(d) 0.694 g
(e) 2.78 g

37. What is the mass of insoluble calcium phosphate (310.18 g/mol) produced from 0.555 g of calcium chloride (110.98 g/mol)?

$$\_\_CaCl_2(s) \quad + \quad Na_3PO_4(aq) \quad \rightarrow \quad \_Ca_3(PO_4)_2(s) \quad + \quad NaCl(aq)$$

(a) 0.0662 g
(b) 0.517 g
(c) 0.596 g
(d) 1.55 g
(e) 4.65 g

38. What is the mass of sodium phosphate (163.94 g/mol) that yields 1.00 g of calcium phosphate (310.18 g/mol) precipitate?

$$\_\_CaCl_2(s) \quad + \quad Na_3PO_4(aq) \quad \rightarrow \quad \_Ca_3(PO_4)_2(s) \quad + \quad NaCl(aq)$$

(a) 0.264 g
(b) 0.358 g
(c) 0.931 g
(d) 1.06 g
(e) 8.38 g

39. What is the mass of aluminum oxide (101.96 g/mol) produced from 3.59 g of iron(II) oxide (71.85 g/mol)?

$$\_\_FeO(l) \quad + \quad \_Al(l) \quad \xrightarrow{\Delta} \quad \_\_Fe(l) \quad + \quad \_Al_2O_3(l)$$

(a) 0.842 g
(b) 1.70 g
(c) 5.10 g
(d) 7.58 g
(e) 15.3 g

40. What is the mass of aluminum metal that reacts to give 1.00 g of iron?

$$\_\_FeO(l) + \_\_Al(l) \xrightarrow{\Delta} \_\_Fe(l) + \_\_Al_2O_3(l)$$

(a) 0.322 g
(b) 0.483 g
(c) 0.725 g
(d) 0.966 g
(e) 1.449 g

41. What is the mass of aluminum oxide (101.96 g/mol) produced from 1.74 g of manganese(IV) oxide (86.94 g/mol)?

$$\_\_MnO_2(l) + Al(l) \xrightarrow{\Delta} \_Mn(l) + \_Al_2O_3(l)$$

(a) 0.988 g
(b) 1.36 g
(c) 2.04 g
(d) 2.22 g
(e) 3.06 g

42. What is the mass of aluminum metal that reacts to give 11.1 g of manganese metal?

$$\_\_MnO_2(l) + Al(l) \xrightarrow{\Delta} \_Mn(l) + \_Al_2O_3(l)$$

(a) 3.64 g
(b) 4.09 g
(c) 5.45 g
(d) 7.27 g
(e) 8.18 g

43. What is the mass of hydrogen gas released from 2.70 g of aluminum metal and hydrochloric acid?

$$\_\_Al(s) + \_HCl(aq) \rightarrow \_AlCl_3(aq) + H_2(g)$$

(a) 0.101 g
(b) 0.135 g
(c) 0.202 g
(d) 0.303 g
(e) 0.606 g

44. What is the mass of aluminum metal that reacts to give 1.00 g of hydrogen gas?

$$\_\_Al(s) + \_HCl(aq) \rightarrow \_AlCl_3(aq) + H_2(g)$$

(a) 4.46 g
(b) 8.90 g
(c) 13.4 g
(d) 20.0 g
(e) 26.7 g

Limiting Reactant Problems

45. Considering the limiting reactant concept, how many moles of C are produced from the reaction of 1.00 mol A and 1.00 mol B?

$$A(g) + 2\,B(g) \rightarrow 3\,C(g)$$

(a) 1.00 mol
(b) 1.50 mol
(c) 2.00 mol
(d) 3.00 mol
(e) none of the above

46. Considering the limiting reactant concept, how many moles of C are produced from the reaction of 1.50 mol A and 3.50 mol B?

$$A(g) + 2\,B(g) \rightarrow 3\,C(g)$$

(a) 1.50 mol
(b) 3.50 mol
(c) 4.50 mol
(d) 5.25 mol
(e) none of the above

47. Considering the limiting reactant concept, how many moles of C are produced from the reaction of 2.00 mol A and 4.50 mol B?

$$A(g) + 3\,B(g) \rightarrow 2\,C(g)$$

(a) 2.00 mol
(b) 3.00 mol
(c) 4.00 mol
(d) 4.50 mol
(e) none of the above

48. Considering the limiting reactant concept, how many moles of copper(I) sulfide are produced from the reaction of 1.00 mol of copper and 1.00 mol of sulfur?

$$2\ Cu(s) \quad + \quad S(s) \quad \xrightarrow{\Delta} \quad Cu_2S(s)$$

(a) 0.500 mol
(b) 1.00 mol
(c) 1.50 mol
(d) 2.00 mol
(e) none of the above

49. Considering the limiting reactant concept, how many moles of copper(I) sulfide are produced from the reaction of 3.00 mol of copper and 1.00 mol of sulfur?

$$2\ Cu(s) \quad + \quad S(s) \quad \xrightarrow{\Delta} \quad Cu_2S(s)$$

(a) 1.00 mol
(b) 1.50 mol
(c) 4.00 mol
(d) 6.00 mol
(e) none of the above

50. Considering the limiting reactant concept, how many moles of cobalt(III) oxide are produced from the reaction of 1.00 mol of cobalt and 1.00 mol of oxygen gas?

$$4\ Co(s) \quad + \quad 3\ O_2(g) \quad \xrightarrow{\Delta} \quad 2\ Co_2O_3(s)$$

(a) 0.500 mol
(b) 0.667 mol
(c) 1.50 mol
(d) 2.00 mol
(e) none of the above

51. Considering the limiting reactant, what is the mass of zinc sulfide (97.46 g/mol) produced from 0.750 g of zinc and 0.750 g of sulfur?

$$Zn(s) \quad + \quad S(s) \quad \xrightarrow{\Delta} \quad ZnS(s)$$

(a) 0.560 g
(b) 1.12 g
(c) 1.50 g
(d) 2.24 g
(e) 2.28 g

52. Considering the limiting reactant, what is the mass of zinc sulfide (97.46 g/mol) produced from 0.750 g of zinc and 0.350 g of sulfur?

$$Zn(s) + S(s) \xrightarrow{\Delta} ZnS(s)$$

(a) 0.530 g
(b) 1.06 g
(c) 1.10 g
(d) 1.12 g
(e) 2.28 g

53. Considering the limiting reactant, what mass of cobalt(III) sulfide (214.07 g/mol) is produced from 0.750 g of cobalt and 0.350 g of sulfur?

$$2\ Co(s) + 3\ S(s) \xrightarrow{\Delta} Co_2S_3(s)$$

(a) 0.779 g
(b) 1.36 g
(c) 2.34 g
(d) 2.72 g
(e) 5.45 g

54. Considering the limiting reactant, what is the mass of iron produced from 75.0 g of ferrous oxide (71.85 g/mol) and 25.0 g of magnesium metal?

$$FeO(s) + Mg(s) \xrightarrow{\Delta} Fe(l) + MgO(s)$$

(a) 28.7 g
(b) 29.1 g
(c) 57.4 g
(d) 58.3 g
(e) 100.0 g

55. Considering the limiting reactant, what is the mass of iron produced from 80.0 g of ferrous oxide (71.85 g/mol) and 20.0 g of magnesium metal?

$$FeO(s) + Mg(s) \xrightarrow{\Delta} Fe(l) + MgO(s)$$

(a) 23.0 g
(b) 31.1 g
(c) 45.9 g
(d) 62.2 g
(e) 100.0 g

56. Considering the limiting reactant, what is the mass of manganese produced from 25.0 g of manganese(IV) oxide (86.94 g/mol) and 25.0 g of aluminum metal?

$$3\ MnO_2(s) + 4\ Al(s) \xrightarrow{\Delta} 3\ Mn(l) + 2\ Al_2O_3(s)$$

(a) 15.8 g
(b) 38.2 g
(c) 47.4 g
(d) 50.9 g
(e) 67.8 g

57. Considering the limiting reactant, what is the volume of NO gas produced from 30.0 L of nitrogen gas and 40.0 L of oxygen gas? (Assume constant conditions.)

$$N_2(g) + O_2(g) \xrightarrow{\Delta} 2\ NO(g)$$

(a) 30.0 L
(b) 40.0 L
(c) 60.0 L
(d) 80.0 L
(e) none of the above

58. Considering the limiting reactant, what is the volume of NO gas produced from 30.0 L of nitrogen gas and 20.0 L of oxygen gas? (Assume constant conditions.)

$$N_2(g) + O_2(g) \xrightarrow{\Delta} 2\ NO(g)$$

(a) 20.0 L
(b) 30.0 L
(c) 40.0 L
(d) 60.0 L
(e) none of the above

59. Considering the limiting reactant, what is the volume of NO_2 gas produced from 3.00 L of NO gas and 2.00 L of oxygen gas? (Assume constant conditions.)

$$2\ NO(g) + O_2(g) \xrightarrow{\Delta} 2\ NO_2(g)$$

(a) 1.00 L
(b) 2.00 L
(c) 3.00 L
(d) 4.00 L
(e) none of the above

60. Considering the limiting reactant, what is the volume of NO gas produced from 40.0 L of ammonia gas and 40.0 L of oxygen gas? (Assume constant conditions.)

$$4\ NH_3(g) \quad + \quad 5\ O_2(g) \quad \xrightarrow{\Delta} \quad 4\ NO(g) \quad + \quad 6\ H_2O(g)$$

(a) 32.0 L
(b) 40.0 L
(c) 50.0 L
(d) 80.0 L
(e) none of the above

61. Considering the limiting reactant, what is the volume of NO gas produced from 50.0 L of ammonia gas and 60.0 L of oxygen gas? (Assume constant conditions.)

$$4\ NH_3(g) \quad + \quad 5\ O_2(g) \quad \xrightarrow{\Delta} \quad 4\ NO(g) \quad + \quad 6\ H_2O(g)$$

(a) 48.0 L
(b) 50.0 L
(c) 60.0 L
(d) 75.0 L
(e) none of the above

62. Considering the limiting reactant, what is the volume of NO gas produced from 60.0 L of ammonia gas and 50.0 L of oxygen gas? (Assume constant conditions.)

$$4\ NH_3(g) \quad + \quad 5\ O_2(g) \quad \xrightarrow{\Delta} \quad 4\ NO(g) \quad + \quad 6\ H_2O(g)$$

(a) 40.0 L
(b) 50.0 L
(c) 60.0 L
(d) 62.5 L
(e) none of the above

Percent Yield

63. Starting with 1.550 g of potassium chlorate, a student releases 0.617 g of oxygen gas. If the calculated mass of oxygen gas is 0.607 g, what is the percent yield?
    (a) 39.2%
    (b) 39.8%
    (c) 98.4%
    (d) 102%
    (e) 255%

64. Starting with 0.657 g of lead(II) nitrate, a student collects 0.925 g of precipitate. If the calculated mass of precipitate is 0.914 g, what is the percent yield?
    (a) 71.0%
    (b) 71.9%
    (c) 98.8%
    (d) 101%
    (e) 139%

65. Starting with 1.56 g of salicylic acid, a student prepares 1.75 g of aspirin. If the calculated mass of aspirin is 1.88 g, what is the percent yield?
    (a) 83.0%
    (b) 89.1%
    (c) 93.1%
    (d) 107%
    (e) 121%

66. The decomposition of 1.500 g of baking soda gave 0.200 L of carbon dioxide gas. If the calculated volume of carbon dioxide gas is 0.210 L, what is the percent yield?
    (a) 63.5%
    (b) 70.0%
    (c) 95.2%
    (d) 105%
    (e) 158%

67. The decomposition of 1.500 g of potassium chlorate evolved 405 mL of oxygen gas. If the calculated volume of oxygen gas is 411 mL, what is the percent yield?
    (a) 65.7%
    (b) 67.3%
    (c) 98.5%
    (d) 101%
    (e) 148%

68. The decomposition of 1.500 g of sodium nitrate produced 195 mL of oxygen gas. If the calculated volume of oxygen gas is 198 mL, what is the percent yield?
    (a) 65.7%
    (b) 68.0%
    (c) 98.5%
    (d) 102%
    (e) 152%

Answer Key

1. A
2. A
3. B
4. B
5. A
6. A
7. D
8. C
9. B
10. B
11. B
12. B
13. B
14. C
15. A
16. B
17. A
18. C
19. B
20. C
21. B
22. B
23. A
24. C
25. C
26. A
27. B
28. D
29. C
30. E
31. E
32. A
33. D
34. A
35. C
36. C
37. B
38. D
39. B
40. A
41. B
42. D
43. D
44. B
45. B
46. C
47. B
48. A
49. A
50. A
51. B
52. B
53. A
54. C
55. C
56. A
57. C
58. C
59. C
60. A
61. A
62. A

Acids, Bases, pH, and Redox - *Answers at end*

1. If 10.0 mL of 0.100 *M* HCl is titrated with 0.200 *M* NaOH, what volume of sodium hydroxide solution is required to neutralize the acid?

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

2. If 20.0 mL of 0.500 *M* KOH is titrated with 0.250 *M* HNO_3, what volume of nitric acid is required to neutralize the base?

$$HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$$

3. If 25.0 mL of 0.100 *M* HCl is titrated with 0.150 *M* $Ba(OH)_2$, what volume of barium hydroxide is required to neutralize the acid?

$$2\ HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2\ H_2O(l)$$

4. If 25.0 mL of 0.100 *M* $Ca(OH)_2$ is titrated with 0.200 *M* HNO_3, what volume of nitric acid is required to neutralize the base?

$$2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow 2\ Ca(NO_3)_2(aq) + 2\ H_2O(l)$$

5. If 20.0 mL of 0.200 *M* H_2SO_4 is titrated with 0.100 *M* NaOH, what volume of sodium hydroxide is required to neutralize the acid?

$$H_2SO_4(aq) + 2\ NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$$

6. If 30.0 mL of 0.100 *M* $Ca(OH)_2$ is titrated with 0.150 *M* $HC_2H_3O_2$, what volume of acetic acid is required to neutralize the base?

$$2\ HC_2H_3O_2(aq) + Ca(OH)_2(aq) \rightarrow Ca(C_2H_3O_2)_2(aq) + 2\ H_2O(l)$$

7. If a 50.0 mL sample of ammonium hydroxide is titrated with 25.0 mL of 0.200 *M* nitric acid to a methyl red endpoint, what is the molarity of the base?

$$NH_4OH(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq) + H_2O(l)$$

8. If a 50.0 mL sample of ammonium hydroxide is titrated with 25.0 mL of 0.200 *M* sulfuric acid to a methyl red endpoint, what is the molarity of the base?

$$2\ NH_4OH(aq) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + 2\ H_2O(l)$$

9. If a 25.0 mL sample of sulfuric acid is titrated with 50.0 mL of 0.200 *M* potassium hydroxide to a phenolphthalein endpoint, what is the molarity of the acid?

$$H_2SO_4(aq) + 2\ KOH(aq) \rightarrow K_2SO_4(aq) + 2\ H_2O(l)$$

10. What is the molarity of a hydrochloric acid solution if 20.00 mL of HCl is required to neutralize 0.424 g of sodium carbonate (105.99 g/mol)?

$$2\,HCl(aq) + Na_2CO_3(aq) \rightarrow 2\,NaCl(aq) + H_2O(l) + CO_2(g)$$

11. What is the molarity of a nitric acid solution if 25.00 mL of HNO_3 is required to neutralize 0.424 g of sodium carbonate (105.99 g/mol)?

$$2\,HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2\,NaNO_3(aq) + H_2O(l) + CO_2(g)$$

12. What is the molarity of a sulfuric acid solution if 30.00 mL of H_2SO_4 is required to neutralize 0.840 g of sodium hydrogen carbonate (84.01 g/mol)?

$$H_2SO_4(aq) + 2\,NaHCO_3(aq) \rightarrow Na_2SO_4(aq) + 2\,H_2O(l) + 2\,CO_2(g)$$

13. What is the molarity of a hydrochloric acid solution if 25.00 mL of HCl is required to neutralize 0.500 g of calcium carbonate (100.09 g/mol)?

$$2\,HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

14. What is the molarity of a sodium hydroxide solution if 40.00 mL of NaOH is required to neutralize 0.900 g of oxalic acid, $H_2C_2O_4$, (90.04 g/mol)?

$$H_2C_2O_4(aq) + 2\,NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\,H_2O(l)$$

15. What is the molarity of a sodium hydroxide solution if 35.00 mL of NaOH is required to neutralize 1.555 g of KHP, that is $KHC_8H_4O_4$ (204.23 g/mol)?

$$KHC_8H_4O_4(aq) + NaOH(aq) \rightarrow KNaC_8H_4O_4(aq) + H_2O(l)$$

16. If a 0.200 g sample of sodium hydroxide (40.00 g/mol) is completely neutralized with 0.100 *M* H_2SO_4, what volume of sulfuric acid is required?

$$H_2SO_4(aq) + 2\,NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\,H_2O(l)$$

17. If 0.900 g of oxalic acid, $H_2C_2O_4$, (90.04 g/mol) is completely neutralized with 0.300 *M* NaOH, what volume of sodium hydroxide is required?

$$H_2C_2O_4(aq) + 2\,NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\,H_2O(l)$$

18. If 1.020 g of $KHC_8H_4O_4$ (204.23 g/mol) is completely neutralized with 0.200 *M* $Ba(OH)_2$, what volume of barium hydroxide is required?

$$2\,KHC_8H_4O_4(aq) + Ba(OH)_2(aq) \rightarrow BaK_2(C_8H_4O_4)_2(aq) + 2\,H_2O(l)$$

19. Glycine is an amino acid that can be abbreviated HGly. If 27.50 mL of 0.120 *M* NaOH neutralizes 0.248 g of HGly, what is the molar mass of the amino acid?

$$HGly(aq) + NaOH(aq) \rightarrow NaGly(aq) + H_2O(l)$$

20. Proline is an amino acid that can be abbreviated HPro. If 33.55 mL of 0.150 *M* NaOH neutralizes 0.579 g of HPro, what is the molar mass of the amino acid?

$$HPro(aq) \;+\; NaOH(aq) \;\rightarrow\; NaPro(aq) \;+\; H_2O(l)$$

21. Lactic acid is found in sour milk and can be abbreviated HLac. If 47.50 mL of 0.275 *M* NaOH neutralizes 1.180 g of HLac, what is the molar mass of the acid?

$$HLac(aq) \;+\; NaOH(aq) \;\rightarrow\; NaLac(aq) \;+\; H_2O(l)$$

22. What is the pH of an aqueous solution if the $[\,H^+\,] = 5.5\text{x}10^{-3}$ *M*?

23. What is the pH of an aqueous solution if the $[\,H^+\,] = 4.2\text{x}10^{-5}$ *M*?

24. What is the pH of an aqueous solution if the $[\,H^+\,] = 7/5\text{x}10^{-8}$ *M*?

25. What is the $[\,H^+\,]$ in an acid rain sample that has a pH = 3.22?

26. What is the $[\,H^+\,]$ in a blood sample that has a pH = 7.30?

27. What is the $[\,H^+\,]$ in a bleach sample that has a pH = 9.55?

28. What is the $[\,OH^-\,]$ in a seawater sample that has a pH = 8.65?

29. What is the $[\,OH^-\,]$ in an ammonia solution that has a pH = 10.20?

30. What is the $[\,OH^-\,]$ in an oven-cleaning solution that has a pH = 12.35?

31. What substance is oxidized in the following redox reaction?

$$Zn(s) \;+\; Cu^{2+}(aq) \;\rightarrow\; Zn^{2+}(aq) \;+\; Cu(s)$$

32. What substance is reduced in the following redox reaction?

$$Co(s) \;+\; 2\,HCl(aq) \;\rightarrow\; CoCl_2(aq) \;+\; H_2(g)$$

33. What substance is oxidized in the following redox reaction?

$$F_2(g) \;+\; 2\,Br^-(aq) \;\rightarrow\; 2\,F^-(aq) \;+\; Br_2(l)$$

34. What substance is oxidized in the following redox reaction?

$$HgCl_2(aq) \;+\; Sn^{2+}(aq) \;\rightarrow\; Sn^{4+}(aq) \;+\; Hg_2Cl_2(s) \;+\; Cl^-(aq)$$

35. What substance is reduced in the following redox reaction?

$$H^+(aq) \;+\; Fe(s) \;+\; NO_3^-(aq) \;\rightarrow\; Fe^{3+}(aq) \;+\; NO(aq) \;+\; H_2O(l)$$

Acids, Bases, pH, and Redox - *Answers*

1. If 10.0 mL of 0.100 *M* HCl is titrated with 0.200 *M* NaOH, what volume of sodium hydroxide solution is required to neutralize the acid?

$$HCl(aq) \;+\; NaOH(aq) \;\rightarrow\; NaCl(aq) \;+\; H_2O(l)$$

$M_1V_1 = M_2V_2$ (0.100M) (0.010L) = (0.200M)(V_2) V_2 = 0.005 L = **5 mL**

2. If 20.0 mL of 0.500 *M* KOH is titrated with 0.250 *M* HNO_3, what volume of nitric acid is required to neutralize the base?

$$HNO_3(aq) \;+\; KOH(aq) \;\rightarrow\; KNO_3(aq) \;+\; H_2O(l)$$

$M_1V_1 = M_2V_2$ (0.500M) (0.020L) = (0.250M)(V_2) V_2 = 0.040 L = **40 mL**

3. If 25.0 mL of 0.100 *M* HCl is titrated with 0.150 *M* $Ba(OH)_2$, what volume of barium hydroxide is required to neutralize the acid?

$$2\ HCl(aq) \;+\; Ba(OH)_2(aq) \;\rightarrow\; BaCl_2(aq) \;+\; 2\ H_2O(l)$$

$M_1V_1 = M_2V_2$ (0.100M) (0.025L) = (0.150M)(V_2) V_2 = 0.0166 L = 16.6 mL OH^-

But there are 2 OH's per $Ba(OH)_2$ so it takes half this volume = **8.33 mL of $Ba(OH)_2$**

4. If 25.0 mL of 0.100 *M* $Ca(OH)_2$ is titrated with 0.200 *M* HNO_3, what volume of nitric acid is required to neutralize the base?

$$2\ HNO_3(aq) \;+\; Ca(OH)_2(aq) \;\rightarrow\; 2\ Ca(NO_3)_2(aq) \;+\; 2\ H_2O(l)$$

$M_1V_1 = M_2V_2$ (0.100M) (0.025L) = (0.200M)(V_2) V_2 = 0.0125 L = 12.5 mL H+

But it takes 2 HNO_3's per $Ca(OH)_2$ so it takes twice this volume = **25 mL of HNO_3**

5. If 20.0 mL of 0.200 *M* H_2SO_4 is titrated with 0.100 *M* NaOH, what volume of sodium hydroxide is required to neutralize the acid?

$$H_2SO_4(aq) \;+\; 2\ NaOH(aq) \;\rightarrow\; Na_2SO_4(aq) \;+\; 2\ H_2O(l)$$

0.200 M H_2SO_4 = 0.400 M H^+

$M_1V_1 = M_2V_2$ (0.40M) (0.020L) = (0.100M)(V_2) V_2 = 0.080 L = **80 mL NaOH**

6. If 30.0 mL of 0.100 *M* $Ca(OH)_2$ is titrated with 0.150 *M* $HC_2H_3O_2$, what volume of acetic acid is required to neutralize the base?

$$2\ HC_2H_3O_2(aq) + Ca(OH)_2(aq) \rightarrow Ca(C_2H_3O_2)_2(aq) + 2\ H_2O(l)$$

0.100 M $Ca(OH)_2$ = 0.200 M OH^-

$M_1V_1 = M_2V_2$ (0.200M) (0.030L) = (0.150M)(V_2) V_2 = 0.040 L = **40 mL NaOH**

7. If a 50.0 mL sample of ammonium hydroxide is titrated with 25.0 mL of 0.200 *M* nitric acid to a methyl red endpoint, what is the molarity of the base?

$$NH_4OH(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq) + H_2O(l)$$

$M_1V_1 = M_2V_2$ (0.200M) (0.025L) = (M_2)(0.050L) M_2 = 0.100 M NH_4OH

8. If a 50.0 mL sample of ammonium hydroxide is titrated with 25.0 mL of 0.200 *M* sulfuric acid to a methyl red endpoint, what is the molarity of the base?

$$2\ NH_4OH(aq) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + 2\ H_2O(l)$$

0.200 M H_2SO_4 = 0.400 M H^+

$M_1V_1 = M_2V_2$ (0.400M) (0.025L) = (M_2)(0.050L) M_2 = **0.200 M NH_4OH**

9. If a 25.0 mL sample of sulfuric acid is titrated with 50.0 mL of 0.200 *M* potassium hydroxide to a phenolphthalein endpoint, what is the molarity of the acid?

$$H_2SO_4(aq) + 2\ KOH(aq) \rightarrow K_2SO_4(aq) + 2\ H_2O(l)$$

$M_1V_1 = M_2V_2$ (0.200M) (0.050L) = (M_2)(0.025L) M_2 = 0.400 M H^+

But, there are 2 H's per H_2SO_4 so [H_2SO_4] = **0.200M**

10. What is the molarity of a hydrochloric acid solution if 20.00 mL of HCl is required to neutralize 0.424 g of sodium carbonate (105.99 g/mol)?

$$2\ HCl(aq) + Na_2CO_3(aq) \rightarrow 2\ NaCl(aq) + H_2O(l) + CO_2(g)$$

0.424 g/105.99 g/mol = 0.0040 mol Na_2CO_3

Each Na_2CO_3 requires 2 HCl so we need 0.0080 mol HCl

MV = moles (M)(0.020L) = 0.0080 mole HCl M = **0.40 M HCl**

11. What is the molarity of a nitric acid solution if 25.00 mL of HNO_3 is required to neutralize 0.424 g of sodium carbonate (105.99 g/mol)?

$$2\ HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2\ NaNO_3(aq) + H_2O(l) + CO_2(g)$$

0.424 g/105.99 g/mol = 0.0040 mol Na_2CO_3

Each Na_2CO_3 requires 2 HNO_3 so we need 0.0080 mol HNO_3

MV = moles (M)(0.025L) = 0.0080 mole HNO_3 M = **0.32 M HNO_3**

12. What is the molarity of a sulfuric acid solution if 30.00 mL of H_2SO_4 is required to neutralize 0.840 g of sodium hydrogen carbonate (84.01 g/mol)?

$$H_2SO_4(aq) + 2\ NaHCO_3(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l) + 2\ CO_2(g)$$

0.840 g / 84.01 g/mol = 0.010 mol $NaHCO_3$

It takes 2 NaHCO3 per H_2SO_4 so you need 0.005 mol H_2SO_4

MV = moles M(0.030L) = 0.005 moles M = **0.167 M H_2SO_4**

13. What is the molarity of a hydrochloric acid solution if 25.00 mL of HCl is required to neutralize 0.500 g of calcium carbonate (100.09 g/mol)?

$$2\ HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

0.500 g/100.09 g/mol = 0.005 mol $CaCO_3$

Each mole of CaCO3 requires 2 mol HCl so you need 0.005 x 2 = 0.010 mol HCl

MV = moles M(0.025L) = 0.010 mol M = **0.40 M HCl**

14. What is the molarity of a sodium hydroxide solution if 40.00 mL of NaOH is required to neutralize 0.900 g of oxalic acid, $H_2C_2O_4$, (90.04 g/mol)?

$$H_2C_2O_4(aq) + 2\ NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\ H_2O(l)$$

0.900 g / 90.04 g/mol = 0.010 mol Oxalic acid

It takes 2 mole NaOH for every mole of Oxalic acid

so you need 2 x 0.010 mol = 0.02 mol NaOH

MV = moles M(0.040L) = 0.020 mole NaOH **M = 0.50 M NaOH**

15. What is the molarity of a sodium hydroxide solution if 35.00 mL of NaOH is required to neutralize 1.555 g of KHP, that is $KHC_8H_4O_4$ (204.23 g/mol)?

$$KHC_8H_4O_4(aq) + NaOH(aq) \rightarrow KNaC_8H_4O_4(aq) + H_2O(l)$$

1.555g / 204.23 g/mol = 0.00761 mol KHP

1 mole KHP needs 1 mole of NaOH so, 0.00761 mole KHP = 0.00761 mole NaOH 0.00761

mole NaOH / 0.0351 L = **0.2175 M NaOH**

16. If a 0.200 g sample of sodium hydroxide (40.00 g/mol) is completely neutralized with 0.100 *M* H_2SO_4, what volume of sulfuric acid is required?

$$H_2SO_4(aq) + 2\ NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$$

0.200 g NaOH / 40 g/mol = 0.005 mol NaOH

1 mole of H_2SO_4 needs 2 mole NaOH so 0.005 mole NaOH needs 0.0025 mole H_2SO_4

MV = moles (0.100 M H_2SO_4) (V) = 0.0025 mole V = 0.0250 L = **25 mL**

17. If 0.900 g of oxalic acid, $H_2C_2O_4$, (90.04 g/mol) is completely neutralized with 0.300 *M* NaOH, what volume of sodium hydroxide is required? $H_2C_2O_4(aq)$

$$+ 2\ NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\ H_2O(l)$$

0.900 g / 90.04 g/mol = 0.010 mol Oxalic acid

It takes 2 mole NaOH for every mole of Oxalic acid

so you need 2 x 0.010 mol = 0.02 mol NaOH

MV = moles (0.300M) (V) = 0.020 mole NaOH V = 0.0666 L = **66.6 mL**

18. If 1.020 g of $KHC_8H_4O_4$ (204.23 g/mol) is completely neutralized with 0.200 *M* $Ba(OH)_2$, what volume of barium hydroxide is required?

$$2\ KHC_8H_4O_4(aq) + Ba(OH)_2(aq) \rightarrow BaK_2(C_8H_4O_4)_2(aq) + 2\ H_2O(l)$$

1.020g / 204.23 g/mol = 0.0050 mol KHP

2 mole KHP needs 1 mole of $Ba(OH)_2$ so, 0.0050 mole KHP needs 0.0025 mole $Ba(OH)_2$

MV = moles (0.200 M) (V) = 0.0025 mole $Ba(OH)_2$ V = 0.01250 L = **12.5 mL**

19. Glycine is an amino acid that can be abbreviated HGly. If 27.50 mL of 0.120 *M* NaOH neutralizes 0.248 g of HGly, what is the molar mass of the amino acid?

HGly(aq) + NaOH(aq) → NaGly(aq) +

H_2O(l)

MV = moles (0.120 M) (0.02750L) = 0.033 mole NaOH = 0.0033 mole HGly

0.248 g / 0.0033 mole HGly = 75.12 g/mol HGly

20. Proline is an amino acid that can be abbreviated HPro. If 33.55 mL of 0.150 *M* NaOH neutralizes 0.579 g of HPro, what is the molar mass of the amino acid?

HPro(aq) + NaOH(aq) → NaPro(aq) + H_2O(l)

MV = moles (0.150 M) (0.03355L) = 0.005033 mole NaOH = 0.005033 mole HPro

0.579 g / 0.050033 mole HPro = **115.05 g/mol HPro**

21. Lactic acid is found in sour milk and can be abbreviated HLac. If 47.50 mL of 0.275
M NaOH neutralizes 1.180 g of HLac, what is the molar mass of the acid?

HLac(aq) + NaOH(aq) → NaLac(aq) + H_2O(l)

MV = moles (0.275 M) (0.0475L) = 0.01306 mole NaOH = 0.01306 mole HLac

1.180 g / 0.01306 mole HLac = **90.33 g/mol HLac**

22. What is the pH of an aqueous solution if the [H^+] = $5.5x10^{-3}$

M? pH = - log [H^+] pH = - log [$5.5x10^{-3}$] = **2.26**

23. What is the pH of an aqueous solution if the [H^+] = $4.2x10^{-5}$

M? pH = - log [H^+] pH = - log [$4.2x10^{-5}$] = **4.38**

24. What is the pH of an aqueous solution if the [H^+] = $7.5x10^{-8}$

M? pH = - log [H^+] pH = - log [$7.5x10^{-8}$] = **7.12**

25. What is the [H+] in a blood sample that has a pH = 3.22?
[H+] = 10^{-pH} [H+] = $10^{-3.22}$ [H+] = **$6.03x10^{-4}$ M**

26. What is the [H+] in a blood sample that has a pH = 7.30?
[H+] = 10^{-pH} [H+] = $10^{-7.30}$ [H+] = **$5.01x10^{-8}$ M**

25. What is the [H^+] in a bleach sample that has a pH = 9.55?

$[H+] = 10^{-pH}$ $[H+] = 10^{-9.55}$ $[H+] = \mathbf{2.82x10^{-10}\ M}$

26. What is the [OH^-] in a seawater sample that has a pH = 8.65?

$[H+] = 10^{-pH}$ $[H+] = 10^{-8.65}$ $[H+] = 2.24x10^{-9}$ M

$[H+]\ [OH-] = 1x10^{-14}$ $[2.24x10^{-9}\ M]\ [OH-] = 1x10^{-14}$ $[OH-] = \mathbf{4.46x10^{-9}\ M}$

27. What is the [OH^-] in an ammonia solution that has a pH = 10.20?

$[H+] = 10^{-pH}$ $[H+] = 10^{-10.20}$ $[H+] = 6.31x10^{-11}$ M

$[H+]\ [OH-] = 1x10^{-14}$ $[6.31x10^{-11}\ M]\ [OH-] = 1x10^{-14}$ $[OH-] = \mathbf{1.58x10^{-4}\ M}$

28. What is the [OH^-] in an oven-cleaning solution that has a pH = 12.35?

$[H+] = 10^{-pH}$ $[H+] = 10^{-12.35}$ $[H+] = 4.47x10^{-13}$ M

$[H+]\ [OH-] = 1x10^{-14}$ $[4.47x10^{-13}\ M]\ [OH-] = 1x10^{-14}$ [OH-] = **0.0224 M**

29. What substance is oxidized in the following redox reaction?

Zn(s) + $Cu^{2+}(aq)$ → $Zn^{2+}(aq)$ + Cu(s)

30. What substance is reduced in the following redox reaction? Co(s) + 2 ***H***Cl(aq) → $CoCl_2(aq)$ + $H_2(g)$

31. What substance is oxidized in the following redox reaction?

$F_2(g)$ + 2 ***Br^-*** (aq) → 2 F^- (aq) + $Br_2(l)$

32. What substance is oxidized in the following redox reaction?

$HgCl_2(aq)$ + ***Sn^{2+}***(aq) → $Sn^{4+}(aq)$ + $Hg_2Cl_2(s)$ + $Cl^-(aq)$

33. What substance is reduced in the following redox reaction?

$H^+(aq)$ + Fe(s) + ***N***$O_3^-(aq)$ → $Fe^{3+}(aq)$ + NO(aq) + $H_2O(l)$

Self Quiz Name: ________________________________ Lab Section: _______

This is a sample quiz for CH 221 providing examples of energy. Answers are provided at the end of this handout. *Good luck!*

1. Equal masses of two substances, A & B, each absorb 25 Joules of energy. If the temperature of A increases by 4 degrees and the temperature of B increases by 8 degrees, one can say that
   a) the specific heat of A is double that of B.
   b) the specific heat of B is double that of A.
   c) the specific heat of B is negative.
   d) the specific heat of B is triple that of A.

2. If 25 J are required to change the temperature of 5.0 g of substance A by 2.0°C, what is the specific heat of substance A?
   a) 250 J/g°C   c) 10. J/g°C
   b) 63 J/g°C   d) 2.5 J/g°C

3. How much energy is required to change the temperature of 2.00 g aluminum from 20.0°C to 25.0°C? The specific heat of aluminum is 0.902 J/g°C.
   a) 2.3 J   c) 0.36 J
   b) 9.0 J   d) 0.090 J

4. Consider the thermal energy transfer during a chemical process. When heat is transferred to the system, the process is said to be _______ and the sign of ΔH is ________.
   a) exothermic, positive
   b) endothermic, negative
   c) exothermic, negative
   d) endothermic, positive

5. When two solutions react the container "feels hot." Thus,
   a) the reaction is endothermic.
   b) the reaction is exothermic.
   c) the energy of the universe is increased.
   d) the energy of both the system and the surroundings is decreased.

6. The equation for the standard enthalpy of formation of N_2O_3 is
   a) $N_2O(g) + O_2(g) \rightarrow N_2O_3(g)$
   b) $N_2O_5(g) \rightarrow N_2O_3(g) + O_2(g)$
   c) $NO(g) + NO_2(g) \rightarrow N_2O_3(g)$
   d) $N_2(g) + {}^3/_2\, O_2(g) \rightarrow N_2O_3(g)$

7. For the general reaction
   $2\ A + B_2 \rightarrow 2\ AB$, ΔH is +50.0 kJ.
   We can conclude that
   a) the reaction is endothermic.
   b) the surroundings absorb energy.
   c) the standard enthalpy of formation of AB is -50.0 kJ.
   d) the molecule AB contains less energy than A or B_2.

8. Calculate the enthalpy of combustion of C_3H_6:
   $C_3H_6(g) + {}^9/_2\, O_2(g) \rightarrow 3\ CO_2 + 3\ H_2O$
   using the following data:
   $3\ C(s) + 3\ H_2(g) \rightarrow C_3H_6(g)$ $\Delta H° = 53.3$ kJ
   $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H° = -394$ kJ
   $H_2(g) + {}^1/_2\, O_2(g) \rightarrow H_2O(l)$ $\Delta H° = -286$ kJ
   a) -1517 kJ   c) -626 kJ
   b) 1304 kJ   d) -2093 kJ

9. Which one of the following would have an enthalpy of formation value (ΔH_f) of zero?
   a) $H_2O(g)$ c) $H_2O(l)$
   b) $O(g)$ d) $O_2(g)$

10. Calculate the heat of vaporization of titanium (IV) chloride: $TiCl_4(l) \rightarrow TiCl_4(g)$
    using the following enthalpies of reaction:
    $Ti(s) + 2Cl_2(g) \rightarrow TiCl_4(l)$ $\Delta H° = -804.2$ kJ
    $TiCl_4(g) \rightarrow 2Cl_2(g) + Ti(s)$ $\Delta H° = 763.2$ kJ
    a) -1567.4 kJ c) 1165.0 kJ
    b) -783.7 kJ d) 41.0 kJ

11. Calculate the enthalpy of reaction for:
    $D + F \rightarrow G + M$
    using the following equations and data:
    $G + C \rightarrow A + B$ $\Delta H° = +277$ kJ
    $C + F \rightarrow A$ $\Delta H° = +303$ kJ
    $D \rightarrow B + M$ $\Delta H° = -158$ kJ
    a) -132 kJ c) +422 kJ
    b) -422 kJ d) +132 kJ

12. Calculate the standard enthalpy of the reaction for the process
    $3\ NO(g) \rightarrow N_2O(g) + NO_2(g)$
    using the standard enthalpies of formation (in kJ/mol): NO = 90.0; N_2O = 82.1; NO_2 = 34.0
    a) -153.9 kJ c) -26.1 kJ
    b) 206.1 kJ d) 386.0 kJ

13. The standard molar enthalpy of combustion is -1277.3 kJ for the combustion of ethanol.
    $C_2H_5OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$
    Calculate the standard molar enthalpy of formation for ethanol based on the following standard enthalpies of formation:
    $\Delta H°_f\ CO_2 = -393.5$ kJ/mol
    $\Delta H°_f\ H_2O = -241.8$ kJ/mol
    a) -642.7 kJ/mol c) 235.1 kJ/mol
    b) -235.1 kJ/mol d) 642.7 kJ/mol

14. Calculate the amount of heat needed to change 25.0 g ice at 0°C to water at 0°C. The heat of fusion of H_2O = 333 J/g.
    a) 56.5 kJ c) 7.06 kJ
    b) 8.33 kJ d) 463 kJ

15. How many joules are equivalent to 37.7 cal?
    a) 9.01 J c) 1.51 J
    b) 4.184 J d) 158 J

16. What is the value for the specific heat of liquid water?
    a) 2.418 J/g°C c) 1.248 J/g°C
    b) 4.184 J/g°C d) 8.148 J/g°C

Answers:

1.	**A**	9.	**D**
2.	**D**	10.	**D**
3.	**B**	11.	**A**
4.	**D**	12.	**A**
5.	**B**	13.	**B**

6.	**D**	14.	**B**
7.	**A**	15.	**D**
8.	**D**	16.	**B**

Sample Chemistry Question (Ch. 3 - 5) - CH 221

Questions for Chapters Three, Four and Five:

1. A compound contains by weight 41.4% carbon, 3.47% hydrogen, and 55.1% oxygen. A 0.0500 mole sample of this compound weighs 8.71 g. The molecular formula of the compound is:

a. CHO
b. C_3H_3O
c. $C_3H_3O_3$
d. $C_4H_4O_4$
e. $C_6H_6O_6$

2. Many metals react with halogens to give metal halides. For example: Fe(s) + Cl2(g) -> FeCl2(s) If you begin with 10.0 g of iron,

a. You will need 10.0 g of Cl2 for complete reaction and will produce 22.7 g of FeCl2.
b. You will need 12.7 g of Cl2 for complete reaction and will produce 10.0 g of FeCl2.
c. You will need 12.7 g of Cl2 for complete reaction and will produce 22.7 g of FeCl2.
d. You will need 10.0 g of Cl2 for complete reaction and will produce 10.0 g of FeCl2.
e. You will need 10.0 g of Cl2 for complete reaction and will produce 20.0 g of FeCl2.

3. Caffeine has the formula $C_8H_{10}N_4O_2$. If 5.00 mg of caffeine is burned, how many milligrams of CO_2 are produced?

a. 1.13 mg
b. 1.76 mg
c. 2.06 mg
d. 9.06 mg
e. 5.67 mg

4. Calculate the enthalpy of reaction for the combustion of 9.25 g of butane, C_4H_{10}, using the following standard enthalpies of formation: CO_2(g) = -394 kJ/mol; H_2O(g) = - 286 kJ/mol; C_4H_{10}(g) = -484 kJ/mol

a. -401 kJ
b. -802 kJ
c. +401 kJ
d. +25.6 kJ
e. +201 kJ

5. Find the enthalpy of formation for PCl_3 using the following reactions:

P_4 + 10 Cl_2 -> 4 PCl_5 ΔH_{rxn} = -1774.0 kJ
PCl_3 + Cl_2 -> PCl_5 ΔH_{rxn} = -128.8 kJ

a. -447.3 kJ
b. -314.7 kJ
c. +238.0 kJ
d. +998.4 kJ
e. +87.6 kJ

6. Calculate the quantity of heat required to convert 60.1 g of ice at 0 °C to steam at 100.0 °C using the heat of fusion (333 J/g) and heat of vaporization (2260 J/g) for water.

a. 144 kJ
b. 52.3 kJ
c. 312 kJ
d. 180. kJ
e. 460. kJ

Here are the answers to the previous questions:

----

1. A compound contains by weight 41.4% carbon, 3.47% hydrogen, and 55.1% oxygen. A 0.0500 mole sample of this compound weighs 8.71 g. The molecular formula of the compound is:

a. CHO
b. C_3H_3O
c. $C_3H_3O_3$
d. $C_4H_4O_4$
e. $C_6H_6O_6$

Answer: Assume you have 100 g of the compound. This means you will have 41.4 g C, 3.47 g H and 55.1 g O.

Convert these to moles: (41.4 g C / 12.01) = 3.45 mol C; (3.47 g H / 1.008) = 3.45 mol H; (55.1 g O / 16.00) = 3.44 mol O

Now find empirical formula by dividing moles by smallest quantity; note that in this problem, the mol of C, H and O are similar (i.e. all are about 3.45 mol). We can write this as **CHO = Empirical Formula**

To find the molecular formula, we need a molar mass; this can be accomplished by dividing the 8.71 g by 0.0500 mol = 174 g/mol in molecular formula

The empirical formula, CHO, is (12.01 + 1.008 + 16.00) = 29.02 g/mol, and this is roughly 1/6th of 174 g/mol

So final formula is **(e), C6H6O6**

----

2. Many metals react with halogens to give metal halides. For example: Fe(s) + Cl2(g) –> FeCl2(s) If you begin with 10.0 g of iron,

a) You will need 10.0 g of Cl2 for complete reaction and will produce 22.7 g of FeCl2.
b) You will need 12.7 g of Cl2 for complete reaction and will produce 10.0 g of FeCl2.
c) You will need 12.7 g of Cl2 for complete reaction and will produce 22.7 g of FeCl2.
d) You will need 10.0 g of Cl2 for complete reaction and will produce 10.0 g of FeCl2.
e) You will need 10.0 g of Cl2 for complete reaction and will produce 20.0 g of FeCl2.

Answer: First find g of Cl2 required to react with 10.0 g Fe by converting the g of Fe to moles, then to moles of Cl2 and then to g of Cl2:

10.0 g Fe * (mol Fe / 55.85 g Fe) * (1 mol Cl2 / 1 mol Fe) * (70.9 g Cl2 / mol Cl2) = **12.7 g Cl2**

Next, find how much FeCl2 will be made using the 10.0 g of Fe by converting to moles Fe, then to moles of FeCl2 and finally to g of FeCl2:

10.0 g Fe * (mol Fe / 55.85 g Fe) * (1 mol FeCl2 / 1 mol Fe) * (126.8 g FeCl2 / mol FeCl2) = **22.7 g FeCl2**

Alternatively, you could use law of mass action to find g of product by taking 10.0 g of Fe and adding to the 12.7 g of Cl2 needed (as calculated above) to find the 22.7 g of FeCl2.

The answer to this problem is **(c), 12.7 g of Cl2 for complete reaction and will produce 22.7 g of FeCl2.**

----

3. Caffeine has the formula $C_8H_{10}N_4O_2$. If 5.00 mg of caffeine is burned, how many milligrams of CO_2 are produced?

a. 1.13 mg
b. 1.76 mg
c. 2.06 mg
d. 9.06 mg
e. 5.67 mg

Answer: This is a combustion reaction, so write out the equation:

$C_8H_{10}N_4O_2$ + x O_2 -> y CO_2 + z H_2O + a NO_2

We only need the relationship between caffeine and CO_2, and since all the carbon in CO_2 comes from caffeine, there will be 8 mol of CO_2 per mol of caffeine. Therefore:

5.00 mg * (1 g/1000 mg) * (1 mol caffeine / 194.19 g) * (8 mol CO_2 / 1 mol caffeine) * (44.01 g CO_2 / mol CO_2) * (1000 mg / g) = **9.06 mg, answer d**.

----

4. Calculate the enthalpy of reaction for the combustion of 9.25 g of butane, C_4H_{10}, using the following standard enthalpies of formation: CO_2(g) = -394 kJ/mol; H_2O(g) = - 286 kJ/mol; C_4H_{10}(g) = -484 kJ/mol

a. -401 kJ
b. -802 kJ
c. +401 kJ
d. +25.6 kJ
e. +201 kJ

Answer: This is also a combustion reaction, and the equation will be:

C_4H_{10} + $^{13}/_2$ O_2 -> 4 CO_2 + 5 H_2O

The enthalpy of reaction, ΔH_{rxn}, will equal the enthalpies of the products minus the enthalpies of reactants. In this case,

ΔH_{rxn} = (4*$\Delta H_f(CO_2)$ + 5*$\Delta H_f(H_2O)$) - ($\Delta H_f(C_4H_{10})$ + $^{13}/_2$*$\Delta H_f(O_2)$)

Notice the stoichiometric coefficients; heat of formation values (ΔH_f) are calculated per mole, so if more than one mole is present, a coefficient must be added on to account for the energy.

Remember that elements in their standard states have enthalpies equal to zero; hence, $\Delta H(O_2) = 0$. Therefore,

ΔH_{rxn} = (4*(−394) + 5*(-286)) - ((-484) + $^{13}/_{2}$*0) = -3006 + 484 = -2522 kJ/mol butane

To find the enthalpy of reaction for 9.25 g, convert the mass to moles and multiply:

9.25 g butane * (mol butane / 58.123 g) * (-2522 kJ / mol butane) = **-401 kJ, answer a.**

----

5. Find the enthalpy of formation for PCl_3 using the following reactions:

P_4 + 10 Cl_2 -> 4 PCl_5 ΔH_{rxn} = -1774.0 kJ
PCl_3 + Cl_2 -> PCl_5 ΔH_{rxn} = -128.8 kJ

a. -447.3 kJ
b. -314.7 kJ
c. +238.0 kJ
d. +998.4 kJ
e. 87.6 kJ

Answer: The heat of formation reaction for PCl_3 would be written as:

1/4 P_4 + 3/2 Cl_2 -> PCl_3

Heat of formation equations always have one mole of product and reactants which are elements in their standard states. Combining the first given equation with the four times the reverse of the second given equation leads to:

P_4 + 10 Cl_2 -> 4 PCl_5 ΔH_{rxn} = -1774.0 kJ
4 PCl_5 -> 4 PCl_3 + 4 Cl_2 ΔH_{rxn} = -4*(-128.8 kJ) = 515.2 kJ

P_4 + 6 Cl_2 -> 4 PCl_3 ΔH_{rxn} = -1774.0 + 515.2 = -1258.8 kJ

To get the resulting equation in a form comparable to the heat of formation equation for PCl_3, divide by 4 to get:

1/4 P_4 + 3/2 Cl_2 -> PCl_3 ΔH_{rxn} = 0.25*(-1258.8 kJ) = **-314.7 kJ, answer (b)**

6. Calculate the quantity of heat required to convert 60.1 g of ice at 0 °C to steam at 100.0 °C using the heat of fusion (333 J/g) and heat of vaporization (2260 J/g) for water.

a. 144 kJ
b. 52.3 kJ
c. 312 kJ
d. 180. kJ
e. 460. kJ

Answer: There are 3 calculations in this problem, all of which will be combined to find the total heat required to turn the ice into steam. The first calculation will convert the solid ice to liquid water using the heat of fusion. The second calculation will heat the liquid water from 0 to 100 degrees. The third calculation will transform the liquid water into steam (gas) using the heat of vaporization.

To turn the ice to water:

$q = 333 \text{ J/g} * 60.1 \text{ g} = 2.00*10^4 \text{ J}$

To heat the water from 0 to 100 °C:

$q = mC\Delta T = 60.1 \text{ g} * 4.184 \text{ J/gC} * (100 - 0) = 2.51*10^4 \text{ J}$

To turn the liquid water into steam:

$q = 2260 \text{ J/g}*60.1 \text{ g} = 1.36*10^5 \text{ J}$

Therefore, to convert the ice to steam takes $1.36*10^5$ J + $2.51*10^4$ J + $2.00*10^4$ J = **$1.80*10^5$ J, or 180. kJ, answer (d)**.

CH 221 QM / Chapter 6-7
Self Quiz Name: ______________________ Lab Section: _______

This is a sample quiz providing quantum chemistry examples. Answers are provided at the end of this handout. *Good luck!*

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1. What wavelength corresponds to a frequency of 8.22×10^9 Hz?
   a) 0.307 m d) 0.110 m
   b) 0.0365 m e) 27.4 m
   c) 0.122 m

2. A radio station transmits at 110. MHz (110. x 10^6 Hz). What wavelength is this radio wave?
   a) 3.65×10^{-5} m c) 3.81×10^{-5} m
   b) 3.30 m d) 2.73 m

3. Which one of the following is NOT a proper unit for frequency?
   a) Hz c) $m \cdot s^{-1}$
   b) s^{-1} d) $\frac{1}{sec}$

4. Calculate the wavelength of the fourth line in the Balmer series (the n=6 to n=2 transition) of the hydrogen spectrum ($R = 1.097 \times 10^7\ m^{-1}$, $E_n = -Rhc/n^2$)
   a) 0.1233 m d) 4.102×10^{-7} m
   b) 24.37 m e) 36.56 m
   c) 2.735×10^{-7} m

5. What is the relationship between the energy of a photon of light and its frequency?
   a) $E = \nu$ d) $E = \frac{1}{h\nu}$
   b) $E = \frac{h}{\nu}$ e) $E = \frac{\nu}{h}$
   c) $E = h\nu$

6. What is the energy needed to raise an electron in the hydrogen atom from the second energy level to the third energy level? ($R = 1.097 \times 10^7\ m^{-1}$, $E_n = -Rhc/n^2$)
   a) 1.52×10^4 J d) 4.48×10^{-19} J
   b) 3.63×10^{-19} J e) 3.03×10^{-19} J
   c) 2.18×10^{-19} J

7. What is the de Broglie wavelength of an electron moving at 80.0% the speed of light.
   a) 3.03×10^{-12} m c) 3.30×10^{11} m
   b) 2.42×10^{-12} m d) 1.59×10^{-25} m

8. How many orbitals make up the **4d** subshell?
   a) 0 b) 1 c) 3 d) 5 e) 7

9. The value of ℓ that is related to the following orbital is:

   a) 0 b) 1 c) 2 d) 3 e) 4

10. The correct electron configuration for nitrogen is
   a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
   b) $1s^2\ 2s^2\ 2p^6\ 2d^4$
   c) $1s^2\ 2s^2\ 2p^3$
   d) $1s^2\ 2s^2\ 3s^2\ 4s^1$
   e) $1s^2\ 1p^5$

11. The electron configuration of the indicated atom in the ground state is correctly written for which atom?
    a) Ga   [Ar] $3d^{12}\ 4s^2$
    b) Ni   [Ar] $3d^{10}$
    c) Ni   [Ar] $3s^2\ 3p^8$
    d) Fe   [Ar] $3d^6\ 4s^2$

12. Which of the following sets of quantum numbers is possible for a **3d** electron?
    a) $n = 3, \ell = 3, m_\ell = -2, m_s = +\frac{1}{2}$
    b) $n = 2, \ell = 1, m_\ell = +1, m_s = -\frac{1}{2}$
    c) $n = 3, \ell = 1, m_\ell = 0, m_s = -\frac{1}{2}$
    d) $n = 3, \ell = 2, m_\ell = -2, m_s = +\frac{1}{2}$
    e) $n = 4, \ell = 1, m_\ell = +1, m_s = +\frac{1}{2}$

13. In what section of the periodic table is the **4f** subshell being filled?
    a) period 4
    b) transition elements Y to Cd
    c) noble gases
    d) group IA
    e) lanthanides

14. Which one of the following elements has 3 electrons in a **p** subshell?
    a) Sb b) Na c) Sc d) V e) Nd

15. Which of the following distributions of electrons is correct for three electrons in p-subshell?
    a) ↑ ↑ ↑
    b) ↑↓ ↑ __
    c) ↑ ↑ ↓
    d) ↑ ↑↓ __
    e) ↑↑ ↑ __

16. Which of the following particles would be most paramagnetic?
    a) P b) Ga c) Br d) Cl^- e) Na^+

17. Which of the following correctly represents the ionization of an atom?
    a) $Cl(g) + e^- \rightarrow Cl^-(g)$
    b) $Na(g) \rightarrow Na^+(g) + e^-$
    c) $Na(s) - e^- \rightarrow Na^+(g)$
    d) $Cl_2(g) \rightarrow 2\ Cl(g)$

18. Which of the following is likely to have the largest atomic radius?
    a) H b) Mn c) Cl d) Rb e) Ag

19. Which one of the following isoelectronic species has the smallest radius?
    a) Mg^{2+} b) Na^+ c) Ne d) F^- e) O^{2-}

20. Which of the following has the <u>greatest</u> ionization energy?
    a) K b) Ca c) Fe d) Ga e) Br

21. Which of the following has the <u>lowest</u> ionization energy?
    a) Li b) Na c) K d) Rb e) Cs

22. The successive ionization energies for one of the period three elements are listed below. Which element is referred to?

E_1	577.4 kJ/mol
E_2	1,816 kJ/mol
E_3	2,744 kJ/mol
E_4	11,580 kJ/mol
E_5	15,030 kJ/mol

    a) Na b) Mg c) Al d) Si e) P

Answers:

1.	**B**	12.	**D**
2.	**D**	13.	**E**
3.	**C**	14.	**A**
4.	**D**	15.	**A**
5.	**C**	16.	**A**

6.	**E**	17.	**B**
7.	**A**	18.	**D**
8.	**D**	19.	**A**
9.	**B**	20.	**E**
10.	**C**	21.	**E**
11.	**D**	22.	**C**

CH 221

Sample Quiz #1 Name: ____ Lab Section: _______

Be sure to show all work, use the correct number of significant figures, circle final answers and use correct units in all problems.

1. Calculate these expressions. Include the correct number of significant figures and units in the answer. (4 points)

   (9.994 g - 8.33 g) / (1.44 cm^3 – 0.536 cm^3) ____________

   (9.16 * 10^{+3} mL) * (2.3411 * 10^{-6} g) / 12.001 * 10^{-3} g ____________

2. Convert 892.0 ˚C to ˚F. (4 points)

3. Density Problem (8 points)

   a. Calculate the density of "substance A" when its mass = 11.22 g and its volume is 0.244 cm^3. Express the density in units of g / mm^3.

   b. It costs 10.79 cents to make 11.4 cm^3 of "substance A". Calculate how much it would cost to make 8.91 pounds of "substance A" in units of $. (1 pound = 453.59237 g; 100 cents = 1 $)

4. Differentiate between chemical and physical properties. Give at least one example of each. (4 points)

CH 221
Sample Quiz #1   Name: ____   Lab Section: _______
Answers

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1. Calculate these expressions. Include the correct number of significant figures and units in the answer. (4 points)

   1.8 g/cm^3

   1.79 mL

2. Convert 892.0 ˚C to ˚F. (4 points)

   1637.6 °F

3. Density Problem (8 points)

   a. Calculate the density of "substance A" when its mass = 11.22 g and its volume is 0.244 cm^3. Express the density in units of g / mm^3.

   4.60 x 10^{-2} g/mm^3

   b. It costs 10.79 cents to make 11.4 cm^3 of "substance A". Calculate how much it would cost to make 8.91 pounds of "substance A" in units of $. (1 pound = 453.59237 g; 100 cents = 1 $)

   $0.832

4. Differentiate between chemical and physical properties. Give at least one example of each. (4 points)

chemical properties change the relationship with atoms on an atomic scale and are generally irreversible. (reactions with other elements/compounds)

physical properties do not change the relationship within the atoms are can be reversible. (melting, boiling, magnetism)

CH 221
Sample Quiz #2 Name: ______________________________ Lab Section: _______

Be sure to show all work, use the correct number of significant figures, circle final answers and use correct units in all problems.

---

1. Complete the missing information in the following table: (5 points)

Element name	*Element Symbol*	*Atomic Number*	*Number of protons*	*Number of electrons in neutral atom*
				17
			37	
		7		
	Sc			
Potassium				

2. True or false? **Circle** your answer. (5 points)

Marie Curie discovered the charge-to-mass ratio for an electron.	true	false
The proton has a positive charge.	true	false
Gamma rays are deflected easily by clothing.	true	false
John Dalton discovered the neutron.	true	false
A lithium atom usually has three protons and three neutrons in the nucleus.	true	false

3. Short answer. (5 points)

What is the mass number of an atom of cobalt with 30 neutrons? ____

Which particle has the smallest mass: proton, neutron, or electron ____

Which radioactive decay particle is a helium nucleus: alpha, beta, or gamma ____

How many electrons in a neutral atom of iodine? ____

How many protons are present in $^{79}_{34}\text{Se}$? ____

4. Calculate the atomic number and mass number for an atom with 38 protons, 41 neutrons and 36 electrons. What element is it? What is the atom's symbol? Give the symbol for this isotope in the form $^{A}_{Z}\text{X}$. Is this atom electrically neutral? Explain. (5 points)

CH 221
Sample Quiz #2 Name: ____ Lab Section: _______
Answers

1. Complete the missing information in the following table: (5 points)

Element name	*Element Symbol*	*Atomic Number*	*Number of protons*	*Number of electrons in neutral atom*
chlorine	**Cl**	**17**	**17**	17
rubidium	**Rb**	**37**	37	**37**
nitrogen	**N**	7	**7**	**7**
scandium	Sc	**21**	**21**	**21**
Potassium	**K**	**19**	**19**	**19**

2. True or false? **Circle** your answer. (5 points)

   False

   True

   False

   False

   False

3. Short answer. (5 points)

   57

   electron

   alpha

   53

   34

4. Calculate the atomic number and mass number for an atom with 38 protons, 41 neutrons and 36 electrons. What element is it? What is the atom's symbol? Give the symbol for this isotope in the form ${}^{A}_{Z}X$. Is this atom electrically neutral? Explain. (5 points)

   Strontium, Sr, ${}^{79}_{38}Sr$ Not neutral, different number of electrons and protons

CH 221
Sample Quiz #3   Name: ____   Lab Section: _______

Be sure to show all work, use the correct number of significant figures, circle final answers and use correct units in all problems.

---

1. How many moles are in 10.227 kg of Mo? (3 points)

2. How many moles of boron can be found in $3.212 * 10^7$ atoms of boron? (3 points)

3. List all the elements in the periodic table that exist naturally as diatomic molecules. (4 points)

4. Write the correct formula for each of the following compounds. (4 points)

potassium bromide	________	chromium(II) oxide	______
calcium hydroxide	________	Dinitrogen trioxide	______

5. Write the correct name for each of the following compounds. (6 points)

Li_2S ________________________

$Zr(NO_3)_2$ ________________________

$AsCl_5$ ________________________

SrO ________________________

K_2CO_3 ________________________

SeF_4 ________________________

CH 221
Sample Quiz #3 Name: ____ Lab Section: _______
Answers

1. How many moles are in 10.227 kg of Mo? (3 points)

   106.57 mol Mo

2. How many moles of boron can be found in $3.212*10^{7}$ atoms of boron? (3 points)

   $5.334*10^{-17}$ mol B

3. List all the elements in the periodic table that exist naturally as diatomic molecules. (4 points)

   H N F O I Cl Br

4. Write the correct formula for each of the following compounds. (4 points)

   KBr
   $Ca(OH)_2$
   CrO
   N_2O_3

5. Write the correct name for each of the following compounds. (6 points)

   lithium sulfide
   zirconium(II) nitrate
   arsenic pentachloride
   strontium oxide
   potassium carbonate
   selenium tetrafluroride

CH 221
Sample Quiz #4 Name: ____ Lab Section: _______

Be sure to show all work, use the correct number of significant figures, circle final answers and use correct units in all problems.

---

1. What mass of oxygen, O_2, is required to react completely with 37.1 grams of pentane, C_5H_{12}? (4 points)

$$\mathbf{C_5H_{12}(g) + 8\ O_2(g) \rightarrow 5\ CO_2(g) + 6\ H_2O(g)}$$

2. The reaction of 20.0 g H_2 with 30.0 g O_2 yields 12.4 g H_2O. What is the limiting reactant? What is the theoretical yield in grams? What is the percent yield of this reaction? (6 points) $\mathbf{2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}}$

3. A mass of 2.052 g of a metal carbonate, MCO_3, is heated to give the metal oxide and 0.4576 g CO_2.

$$MCO_3(s) \rightarrow MO(s) + CO_2(g)$$

What is the identity of the metal? (4 points)

4. Fill in the missing stoichiometric coefficients. **Blank entries** will be **considered** to be **zero**. All stoichiometric coefficients must be whole numbers. (6 points)

___ $Pb(NO_3)_2(aq)$ + ___ $LiCl(aq)$ $\rightarrow$ ___ $PbCl_2(s)$ + ___ $LiNO_3(aq)$

___ $C_6H_6(l)$ + ___ $O_2(g)$ $\rightarrow$ ___ $CO_2(g)$ + ___ $H_2O(g)$

___ $N_2(g)$ + ___ $H_2(g)$ $\rightarrow$ ___ $NH_3(g)$

CH 221
Sample Quiz #4 Name: ____ Lab Section: _______
Answers

1. What mass of oxygen, O_2, is required to react completely with 37.1 grams of pentane, C_5H_{12}? (4 points)

$$\mathbf{C_5H_{12}(g) + 8\ O_2(g) \rightarrow 5\ CO_2(g) + 6\ H_2O(g)}$$

132 g O_2

2. The reaction of 20.0 g H_2 with 30.0 g O_2 yields 12.4 g H_2O. What is the limiting reactant? What is the theoretical yield in grams? What is the percent yield of this reaction? (6 points) $\mathbf{2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}}$

LR = O_2
TY = 33.8 g
% yield = 36.7%

3. A mass of 2.052 g of a metal carbonate, MCO_3, is heated to give the metal oxide and 0.4576 g CO_2.

$$MCO_3(s) \rightarrow MO(s) + CO_2(g)$$

What is the identity of the metal? (4 points)

M = Barium

4. Fill in the missing stoichiometric coefficients. **Blank entries** will be **considered** to be **zero**. All stoichiometric coefficients must be whole numbers. (6 points)

1, 2, 1, 2
2, 15, 12, 6
1, 3, 2

CH 221
Sample Quiz #5 Name: ____ Lab Section: _______

Be sure to show all work, use the correct number of significant figures, circle final answers and use correct units in all problems.

---

1. Complete the following problems.
   a. Write the net ionic equation for the following reaction: (3 points)

$$\mathbf{Ba(NO_3)_2(aq) + K_2SO_4(aq) \rightarrow 2\ KNO_3(aq) + BaSO_4(s)}$$

   b. Write the spectator ion(s) in the reaction in #1a. (2 points)

2. Hydrazine, N_2H_4, a base like ammonia, can react with an acid such as sulfuric acid as shown below. What mass of hydrazine reacts with 155 mL of 0.310 M H_2SO_4? (5 points)

$$\mathbf{2\ N_2H_4(aq) + H_2SO_4(aq) \rightarrow 2\ N_2H_5^{+}(aq) + SO_4^{2-}(aq)}$$

3. If 25 J are required to change the temperature of 5.0 g of substance A by 2.0 K, what is the specific heat of substance A? (4 points)

4. Determine ΔH for the following reaction,

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

given the thermochemical equations below. (6 points)

$N_2(g) + O_2(g) \rightarrow 2\ NO(g)$	$\Delta H = +180.8$ kJ
$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$	$\Delta H = -906.2$ kJ
$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$	$\Delta H = -483.6$ kJ

CH 221
Sample Quiz #5   Name: ____   Lab Section: _______
Answers

1. Complete the following problems.
   a. Write the net ionic equation for the following reaction: (3 points)

$$\mathbf{Ba(NO_3)_2(aq) + K_2SO_4(aq) \rightarrow 2\ KNO_3(aq) + BaSO_4(s)}$$

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

   b. Write the spectator ion(s) in the reaction in #1a. (2 points)

K^+ and NO_3^{-1}

2. Hydrazine, N_2H_4, a base like ammonia, can react with an acid such as sulfuric acid as shown below. What mass of hydrazine reacts with 155 mL of 0.310 M H_2SO_4? (5 points)

$$\mathbf{2\ N_2H_4(aq) + H_2SO_4(aq) \rightarrow 2\ N_2H_5^+(aq) + SO_4^{2-}(aq)}$$

3.08 g

3. If 25 J are required to change the temperature of 5.0 g of substance A by 2.0 K, what is the specific heat of substance A? (4 points)

C = 2.5 J/(gK)

4. Determine ΔH for the following reaction,

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

given the thermochemical equations below. (6 points)

$$N_2(g) + O_2(g) \rightarrow 2\ NO(g) \qquad \Delta H = +180.8\text{ kJ}$$
$$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g) \qquad \Delta H = -906.2\text{ kJ}$$
$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g) \qquad \Delta H = -483.6\text{ kJ}$$

$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$ $\qquad$ $\Delta H = -91.5$ kJ

CH 221
Sample Quiz #6   Name: ____   Lab Section: _______

Be sure to show all work, use the correct number of significant figures, circle final answers and use correct units in all problems.

---

1. **Circle** the color of light with the lowest energy: (1 point)

   orange   blue   red   green   yellow

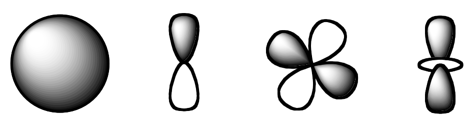
2. **Circle** the color of light with the smallest frequency: (1 point)

   orange   blue   red   green   yellow

3. Match the names on the left with their scientific discovery on the right: (1 point each)

Einstein	_____	a) received Nobel Prize for the photoelectric effect
Planck	_____	b) light is quantized; radiant energy consists of packets called photons
Schrödinger	_____	c) worked for Nazi Germany, the uncertainty principle
Heisenberg	_____	d) developed quantum mechanics
de Broglie	_____	e) 7 page thesis, particles as waves

4. Use the diagram below to answer the following questions: (2 points)



I   II   III   IV

   The number that corresponds to a *p* orbital is: _____

   The number that corresponds to an orbital with no planar nodes is: _____

5. Use the information below to answer the following questions. (2 points) *Use 1s, 2p, 3d, etc. for your answers.*

   What type of orbital is designated $n = 2$, $l = 0$, $m_l = 0$? ______

   What type of orbital is designated $n = 4$, $l = 1$, $m_l = -1$? ______

6. If the de Broglie wavelength of an electron is 112 nm, what is its velocity in m/s? The mass of an electron is $9.11*10^{-31}$ kg. (5 points) *(Note to physics fans: no relativity in this problem!)*

7. If an FM radio station operates at a frequency of 92.3 megahertz (MHz, or 92.3×10^6 Hz), calculate the **wavelength** of its signal in meters and the **energy** of one photon in Joules. (4 points) (+1 bonus if you can name the Portland-area radio station using this information! ☺)

CH 221
Sample Quiz #6   Name: ____   Lab Section: _______
Answers

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1. **Circle** the color of light with the lowest energy: (1 point)

    red

2. **Circle** the color of light with the smallest frequency: (1 point)

    red

3. Match the names on the left with their scientific discovery on the right: (1 point each)

    a

    b

    d

    c

    e

4. Use the diagram below to answer the following questions: (2 points)

    II

    I

5. Use the information below to answer the following questions. (2 points) *Use 1s, 2p, 3d, etc. for your answers.*

    What type of orbital is designated $n = 2$, $l = 0$, $m_l = 0$? __**2s**__

    What type of orbital is designated $n = 4$, $l = 1$, $m_l = -1$? __**4p**__

6. If the de Broglie wavelength of an electron is 112 nm, what is its velocity in m/s? The mass of an electron is $9.11*10^{-31}$ kg. (5 points) *(Note to physics fans: no relativity in this problem!)*

    6490 m/s

7. If an FM radio station operates at a frequency of 92.3 megahertz (MHz, or 92.3×10^{6} Hz), calculate the **wavelength** of its signal in meters and the **energy** of one photon in Joules. (4 points) (+1 bonus if you can name the Portland-area radio station using this information! ☺)

    3.25 m
    6.12×10^{-26} J

CH 221 Sample Exam
Exam I Name: ______________________ Lab Section: ______

Part I: Multiple Choice Questions (100 Points) There is *only* one best answer for each question.

---

1. At 0 °C, a bottle contains 325 mL of water in its liquid state. What is the volume of the water after it freezes (at 0 °C)? The densities of liquid water and ice at 0 °C are 1.000 g/mL and 0.917 g/mL, respectively.

   a. 27.0 mL
   b. 298 mL
   c. 325 mL
   d. 354 mL
   e. 391 mL

2. The radius of a helium atom is 31 pm. What is the radius in nanometers? ($p = 10^{-12}$)

   a. 3.1×10^{-9} nm
   b. 3.1×10^{-6} nm
   c. 3.1×10^{-5} nm
   d. 3.1×10^{-3} nm
   e. 3.1×10^{-2} nm

3. The density of liquid mercury is 13.5 g/cm^3. What mass of mercury (in kg) is required to fill a hollow cylinder having an inner diameter of 2.00 cm to a height of 25.0 cm? ($V = \pi r^2 h$)

   a. 1.06 kg
   b. 4.24 kg
   c. 0.171 kg
   d. 1.71×10^{-4} kg
   e. 4.24×10^{-6} kg

4. The output of a plant is 4335 pounds of ball bearings per work week (five days). If each ball bearing weighs 0.0113 g, how many ball bearings does the plant make in a single day? (453.6 g = 1 pound)

   a. 3.84×10^{5}
   b. 7.67×10^{4}
   c. 867
   d. 3.48×10^{7}
   e. 2.91×10^{6}

5. The density of mercury is 13.6 g/cm^3. The density of mercury is __________ kg/m^3.

   a. 1.36×10^{-2}
   b. 1.36×10^{4}
   c. 1.36×10^{8}
   d. 1.36×10^{-5}
   e. 1.36×10^{-4}

6. The dimensions of a rectangular solid are 8.45 cm long, 4.33 cm wide and 2.85 cm high. If the density of the solid is 9.43 g/cm^3, what is the mass?

   a. 1.12 g
   b. 11.1 g
   c. 154 g
   d. 896 g
   e. 983 g

7. How many protons, neutrons, and electrons are in a neutral oxygen-18 atom?

   a. 6 protons, 8 neutrons, 4 electrons
   b. 6 protons, 10 neutrons, 8 electrons
   c. 8 protons, 8 neutrons, 8 electrons
   d. 8 protons, 10 neutrons, 8 electrons
   e. 8 protons, 10 neutrons, 18 electrons

8. Which of the following atoms contains the largest number of neutrons?

   a. ${}^{31}_{15}P$
   b. ${}^{30}_{14}Si$
   c. ${}^{37}_{17}Cl$
   d. ${}^{32}_{16}S$
   e. ${}^{34}_{16}S$

9. All of the following statements are true EXCEPT

   a. for any neutral element, the number of protons and electrons are equal.
   b. electrons and protons have equal mass, but opposite charges.
   c. the mass number is the sum of the number of protons and neutrons.
   d. the atomic number equals the number of protons.
   e. isotopes of an element have identical atomic numbers.

10. You have 4.15 g of each of the following elements: Ca, Cu, Ce, Cs, Cf. Which sample contains the largest number of atoms?

    a. Ca
    b. Cu
    c. Ce
    d. Cs
    e. Cf

11. Pennies minted after 1983 are composed of 97% zinc and 3.0% copper and have a mass of 2.46 g. How many moles of copper are in a penny?

    a. 0.0012 mol
    b. 0.014 mol
    c. 0.038 mol
    d. 0.040 mol
    e. 25 mol

12. What mass of He contains the same number of atoms as 5.0 g Kr?

    a. 0.24 g
    b. 0.80 g
    c. 1.2 g
    d. 5.0 g
    e. 1.0×10^2 g

13. The molar mass of cesium is 132.9 g/mol. What is the mass of a single Cs atom?

   a. 2.207×10^{-22} g
   b. 1.249×10^{-26} g
   c. 2.763×10^{-23} g
   d. 4.531×10^{21} g
   e. $1.329 \text{ x } 10^{-23}$ g

14. Identify the ions present in $KHCO_3$.

   a. $KHCO_3$ is not ionic.
   b. KH^+, and CO_3^{-1}
   c. K^+, H^+, C^{4+}, and O^{2-}
   d. KH^{2+} and CO_3^{2-}
   e. K^+ and HCO_3^{-1}

15. What is the molar mass of cobalt(II) iodide hexahydrate?

   a. 212.8 g/mol
   b. 293.9 g/mol
   c. 312.7 g/mol
   d. 420.8 g/mol
   e. 465.1 g/mol

16. How many oxygen atoms are in 1.50 mol of SO_3?

   a. 7.71×10^{21} atoms
   b. 1.12×10^{22} atoms
   c. 3.01×10^{22} atoms
   d. 9.03×10^{23} atoms
   e. 2.71×10^{24} atoms

17. If 1.00 g of an unknown molecular compound contains 4.55×10^{21} molecules, what is its molar mass?

   a. 44.0 g/mol
   b. 66.4 g/mol
   c. 72.1 g/mol
   d. 98.1 g/mol
   e. 132 g/mol

18. Which of the following quantities of compounds contains the largest total number of atoms?

   a. 1.0 mole of H_3PO_4
   b. 2.0 moles of H_2SO_3
   c. 3.0 moles of $HClO_4$
   d. 4.0 moles of H_2S
   e. 5.0 moles of HBr

19. What is the mass percent of each element in dichloromethane, CH_2Cl_2?

   a. 10.06% C, 60.24% H, 29.70% Cl
   b. 20.00% C, 20.00% H, 60.00% Cl
   c. 24.10% C, 3.11% H, 72.79% Cl
   d. 33.87% C, 0.22% H, 65.91% Cl
   e. 14.14% C, 2.37% H, 83.48% Cl

20. A molecule is found to contain 47.35% C, 10.60% H, and 42.05% O. What is the empirical formula for this molecule?

    a. C_2H_6O
    b. $C_3H_4O_2$
    c. $C_3H_8O_2$
    d. $C_4H_6O_2$
    e. $C_4H_8O_3$

21. A 2.000 g sample of $CoCl_2 \cdot xH_2O$ is dried in an oven. When the anhydrous salt is removed from the oven, its mass is 1.565 g. What is the value of x?

    a. 1
    b. 2
    c. 3
    d. 4
    e. 6

22. Benzene, an organic solvent, has the empirical formula CH. If the molar mass of benzene is 78.11 g/mol, what is the molecular formula of benzene?

    a. C_4H_{30}
    b. C_5H_{18}
    c. C_6H_6
    d. C_7H_8
    e. C_2H_2

23. What is the common name for NH_3?

    a. ammonia
    b. nitrogen trihydride
    c. trihydrogen nitride
    d. ammonium
    e. nitrous

24. How many oxygen atoms are in 0.20 g CO_2?

    a. 2.4×10^{23} oxygen atoms
    b. 2.7×10^{21} oxygen atoms
    c. 5.5×10^{21} oxygen atoms
    d. 1.2×10^{23} oxygen atoms
    e. 9.7×10^{10} oxygen atoms

25. What formula represents the binary compound formed by magnesium and phosphorus?

    a. MgP
    b. Mg_2P
    c. MgP_3
    d. Mg_3P_2
    e. Mg_2P_3

Part II: Short Answer / Calculation. *Show all work!*

1. The density of chromium is 7.19 g/cm^3 at room temperature. How many atoms are in a cube of pure chromium that has an edge of 3.47 inches? (1 inch = 2.54 cm, edge^3 = Volume)

2. A molecule with a molecular weight of 180.18 g/mol is analyzed and found to contain 40.00% carbon, 6.72% hydrogen and 53.28% oxygen. Find the empirical formula and molecular formula for this compound.

3. Silver has two stable isotopes, one (Ag-109) with the exact mass of 108.9047 amu and an abundance of 48.18%. Determine the identity and exact mass of the second isotope. (The atomic mass of silver = 107.87).

4. Write the correct name for each of the following compounds.

Li_2CrO_4	______________________________
$K_2C_2O_4$	______________________________
H_3As	______________________________
NCl_3	______________________________
NH_4ClO	______________________________
SBr_4	______________________________
$Ca(OH)_2$	______________________________
H_2O	______________________________
$TiNO_2$	______________________________
SO_3	______________________________

CH 221 Sample Exam
Exam I Name: ____ Lab Section: ______
Answers

<u>Part I</u>: Multiple Choice Questions

1. D
2. E
3. A
4. D
5. B
6. E

7. D
8. C
9. B
10. A
11. A
12. A

13. A
14. E
15. D
16. E
17. E
18. C
19. E

20. C
21. B
22. C
23. A
24. C
25. D

<u>Part II</u>: Short Answer / Calculation.

1. 5.70×10^{25} atoms
2. EF = CH_2O, MF = $C_6H_{12}O_6$
3. ^{107}Ag
4. *names:*
    a. lithium chromate
    b. potassium oxalate
    c. hydroarsenic acid or trihydrogen monoarsenide
    d. nitrogen trichloride
    e. ammonium hypochlorite
    f. sulfur tetrabromide
    g. calcium hydroxide
    h. water or dihydrogen monoxide
    i. titanium(I) nitrite
    j. sulfur trioxide

CH 221 Sample Exam
Exam II Name: ________________________________ Lab Section: ______

Part I: Multiple Choice Questions (100 Points) There is *only* one best answer for each question.

---

1. When methanol undergoes complete combustion, the products are carbon dioxide and water according to the equation below. What are the respective coefficients when the equation is balanced with the smallest whole numbers?

$$\_\_\ CH_3OH(l) + \_\_\ O_2(g) \rightarrow \_\_\ CO_2(g) + \_\_\ H_2O(g)$$

a. 1, 1, 1, 1
b. 1, 2, 1, 2
c. 2, 2, 2, 4
d. 2, 3, 2, 4
e. 2, 4, 6, 4

2. Write a balanced chemical equation for the combustion of pentane, C_5H_{12}.

a. $C_5H_{12}(g) + 8\ O_2(g) \rightarrow 5\ CO_2(g) + 6\ H_2O(g)$
b. $C_5H_{12}(g) \rightarrow 5\ C(s) + 6\ H_2(g)$
c. $C_5H_{12}(g) + 9\ O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)$
d. $C_5H_{12}(g) + 11\ O_2(g) \rightarrow C_5O_{10}(g) + 6\ H_2O(g)$
e. $C_5H_{12}(g) + 11\ O_2(g) \rightarrow 5\ CO_2(g) + 12\ H_2O(g)$

3. Aluminum reacts with oxygen to produce aluminum oxide. If 5.0 moles of Al react with excess O_2, how many moles of Al_2O_3 can be formed? The reaction: $4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$

a. 1.0 mol
b. 2.0 mol
c. 2.5 mol
d. 5.0 mol
e. 10.0 mol

4. Nitroglycerine decomposes violently according to the **balanced** chemical equation below. How many total moles of gases are produced from the decomposition of 1.00 mol $C_3H_5(NO_3)_3$?

$$4\ C_3H_5(NO_3)_3(l) \rightarrow 12\ CO_2(g) + 6\ N_2(g) + 10\ H_2O(g) + O_2(g)$$

a. 4.00 mol
b. 6.50 mol
c. 7.25 mol
d. 16.5 mol
e. 29.0 mol

5. The reaction of water and coal at a high temperature produces a mixture of hydrogen and carbon monoxide gases. This mixture is known as synthesis gas (or syngas). What mass of hydrogen gas can be formed from the reaction of 51.3 g of carbon with excess water?

$$C(s) + H_2O(g) \rightarrow H_2(g) + CO(g)$$

a. 4.31 g
b. 8.61 g
c. 17.2 g
d. 1.20×10^2 g
e. 306 g

6. If 0.250 moles of bromine and 0.600 moles of ammonia react according to the equation below, what is the maximum amount of ammonium bromide (in moles) produced?

$$3\ Br_2(l)\ +\ 8\ NH_3(g) \rightarrow 6\ NH_4Br(s)\ +\ N_2(g)$$

   a. 0.250 mol
   b. 0.450 mol
   c. 0.500 mol
   d. 0.600 mol
   e. 0.800 mol

7. The reaction of 10.0 g $H_2(g)$ with 10.0 g $O_2(g)$ yields 8.43 g $H_2O(g)$. What is the percent yield of this reaction?

   a. 9.43%
   b. 27.3%
   c. 42.2%
   d. 66.8%
   e. 74.6%

8. A mass of 2.052 g of a metal carbonate, MCO_3, is heated to give the metal oxide and 0.4576 g CO_2. What is the identity (*hint:* molar mass!) of the metal? The balanced equation: **$MCO_3(s) \rightarrow MO(s) + CO_2(g)$**

   a. Cu
   b. Mg
   c. Ca
   d. Ba
   e. Co

9. Which one of the following compounds is a nonelectrolyte when dissolved in water?

   a. CH_3CH_2OH
   b. $ZnBr_2$
   c. LiCl
   d. $Ca(NO_3)_2$
   e. KOH

10. Which of the following compounds will be soluble in water: LiOH, $Mg(OH)_2$, $Cu(OH)_2$, and $Fe(OH)_3$?

   a. LiOH only
   b. LiOH and $Mg(OH)_2$
   c. $Cu(OH)_2$ and $Fe(OH)_3$
   d. $Mg(OH)_2$ and $Cu(OH)_2$
   e. LiOH, $Mg(OH)_2$, and $Fe(OH)_3$

11. A white solid is either $Pb(NO_3)_2$ or $Zn(NO_3)_2$. If an aqueous solution is prepared, which reagent will allow you to distinguish between the two compounds?

   a. KBr
   b. HNO_3
   c. CH_3CO_2H
   d. NH_4ClO_4
   e. $LiNO_3$

12. What is the net ionic equation for the reaction of aqueous lead(II) nitrate with aqueous sodium bromide?

    a. $Pb(NO_3)_2(aq) + 2\ NaBr(aq) \rightarrow PbBr_2(aq) + 2\ NaNO_3(s)$
    b. $Na^{+}(aq) + NO_3^{-1}(aq) \rightarrow NaNO_3(s)$
    c. $Pb^{2+}(aq) + 2\ Br^{-1}(aq) \rightarrow PbBr_2(s)$
    d. $Pb^{2+}(aq) + 2\ Na^{+1}(aq) \rightarrow PbNa_2(s)$
    e. $Pb(NO_3)_2(aq) + 2\ NaBr(aq) \rightarrow PbBr_2(s) + 2\ NaNO_3(aq)$

13. What are the spectator ions in the reaction between aqueous perchloric acid and aqueous potassium hydroxide?

    a. H^{+1}, ClO_4^{-1}, K^{+1}, and OH^{-1}
    b. H^{+1} and OH^{-1}
    c. K^{+1} and ClO_4^{-1}
    d. H^{+1} and ClO_4^{-1}
    e. K^{+1} and OH^{-1}

14. Which species in the reaction undergoes reduction? $\mathbf{Sn(s) + 2\ H^{+1}(aq) \rightarrow Sn^{2+}(aq) + H_2(g)}$

    a. Sn
    b. H^{+1}
    c. Sn^{2+}
    d. H_2
    e. No compound is reduced.

15. What is the oxidation number of iodine in IO_3^{-1}?

    a. -1
    b. 0
    c. +3
    d. +5
    e. +7

16. Which of the following are oxidation-reduction reactions?
    1. $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$
    2. $Pb(ClO_4)_2(aq) + 2\ KI(aq) \rightarrow PbI_2(s) + 2\ KClO_4(aq)$
    3. $CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$

    a. 1 only
    b. 2 only
    c. 1 and 2
    d. 1 and 3
    e. 2 and 3

17. What is the mass of solute in 225 mL of 5.91×10^{-2} M KIO_3?

    a. 0.0133 g
    b. 0.0562 g
    c. 0.263 g
    d. 1.51 g
    e. 2.85 g

18. What is the pH of 0.27 M HNO_3?

    a. 0.57
    b. 1.31
    c. 1.86
    d. 2.70
    e. 13.43

19. Which one of the following statements is INCORRECT?

    a. In an exothermic process heat is transferred from the system to the surroundings.
    b. The greater the heat capacity of an object, the more thermal energy it can store.
    c. The SI unit of specific heat capacity is joules per gram per kelvin.
    d. The specific heat capacity has a positive value for liquids and a negative value for gases.
    e. When heat is transferred from the system to the surroundings, q is negative.

20. If 136 J is required to change the temperature of 8.75 g of nickel by 35.0 K, what is the specific heat capacity of mercury?

    a. 0.0294 J/g·K
    b. 0.311 J/g·K
    c. 0.417 J/g·K
    d. 0.444 J/g·K
    e. 2.25 J/g·K

21. If the same amount of heat is added to 5.00 g samples of each of the metals below, which metal will experience the smallest temperature change?

Metal	Specific Heat Capacity (J/g·K)
Al	0.897
Au	0.129
Cu	0.385
Fe	0.449
K	0.753

    a. Al
    b. Au
    c. Cu
    d. Fe
    e. K

22. If 50.0 g H_2O at 13.6 °C is combined with 85.0 g H_2O at 93.7 °C, what is the final temperature of the mixture?

    a. 26.1 °C
    b. 40.0 °C
    c. 56.1 °C
    d. 64.0 °C
    e. 80.1 °C

23. The heat of vaporization of benzene, C_6H_6, is 30.8 kJ/mol at its boiling point of 80.1 °C. How much heat is required to vaporize 128 g benzene at its boiling point?

   a. 4.04 kJ
   b. 18.8 kJ
   c. 19.3 kJ
   d. 50.5 kJ
   e. 4.04×10^3 kJ

24. The balanced thermochemical equation for the combustion of hexane is shown below. What is the enthalpy change for the combustion of 2.50 g of C_6H_{14}?

$$C_6H_{14}(g) + {}^{19}/_2\, O_2(g) \rightarrow 6\, CO_2(g) + 7\, H_2O(g) \qquad \Delta H° = -4163 \text{ kJ}$$

   a. -121 kJ
   b. -1.66×10^3 kJ
   c. -1.04×10^4 kJ
   d. -1.43×10^5 kJ
   e. -3.59×10^5 kJ

25. Which of the following chemical equations corresponds to the standard molar enthalpy of formation of N_2O?

   a. $NO(g) + {}^1/_2\, N_2(g) \rightarrow N_2O(g)$
   b. $N_2(g) + {}^1/_2\, O_2(g) \rightarrow N_2O(g)$
   c. $2\, N(g) + O(g) \rightarrow N_2O(g)$
   d. $N_2(g) + O(g) \rightarrow N_2O(g)$
   e. $2\, N_2(g) + O_2(g) \rightarrow 2\, N_2O(g)$

Part II: Short Answer / Calculation. *Show all work!*

1. Determine the heat of reaction for the oxidation of iron,

$$\mathbf{4\, Fe(s) + 3\, O_2(g) \rightarrow 2\, Fe_2O_3(s)}$$

given the thermochemical equations below.

$$2\, Fe(s) + 6\, H_2O(l) \rightarrow 2\, Fe(OH)_3(s) + 3\, H_2(g) \qquad \Delta H = +321.8 \text{ kJ}$$
$$2\, H_2O(l) \rightarrow 2\, H_2(g) + O_2(g) \qquad \Delta H = +571.7 \text{ kJ}$$
$$Fe_2O_3(s) + 3\, H_2O(l) \rightarrow 2\, Fe(OH)_3(s) \qquad \Delta H = +288.6 \text{ kJ}$$

Part II: Short Answer / Calculation *(continued) Show all work!*

2. Benzoic acid contains C, H, and O atoms. When 1.500 g benzoic acid is burned in oxygen, 3.784 g CO_2 and 0.6639 g H_2O are produced. The molar mass of the compound is found to be 122.12 g/mol. Use this information to find the **empirical formula and molecular formula** of benzoic acid.

3. The standard enthalpy change for the combustion of 1 mole of benzene is -3267.4 kJ.

$$C_6H_6(l) + 15/2\ O_2(g) \rightarrow 6\ CO_2(g) + 3\ H_2O(l)$$

Calculate ΔH_f° for benzene based on the following standard molar enthalpies of formation.

molecule	ΔH_f^o (kJ/mol)
$CO_2(g)$	-393.5
$H_2O(l)$	-285.8

Extra Credit Question: Use the above problem to calculate the energy (in kJ) released upon burning 15.0 g of benzene.

CH 221 Sample Exam
Exam II Name: ____ Lab Section: ______
Answers

<u>Part I</u>: Multiple Choice Questions

1. D
2. A
3. C
4. C
5. B

6. B
7. E
8. D
9. A
10. A
11. A

12. C
13. C
14. B
15. D
16. A
17. E

18. A
19. D
20. D
21. A
22. D

23. D
24. A
25. B

<u>Part II</u>: Short Answer / Calculation.

1. -1648.7 kJ
2. EF = $C_7H_6O_2$, MF = $C_7H_6O_2$
3. +49.0 kJ
4. *-627 kJ (extra credit)*

CH 221 Sample Final Exam Name: ____ Lab Section: ______

There is *only* one best answer for each question. Good luck!

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1. If a 21.00 gram sample of a Cu-Zn-Ni alloy contains 7.75 g Cu and 10.58 g Ni, what is the percent composition of Zn?

   a. 36.9%
   b. 2.67%
   c. 50.4%
   d. 20.7%
   e. 12.7%

2. What is the atomic symbol for an element with 39 protons and 50 neutrons?

   a. $^{89}_{11}Na$
   b. $^{50}_{39}Y$
   c. $^{89}_{39}Y$
   d. $^{89}_{50}Sn$
   e. $^{227}_{89}Ac$

3. Rubidium (Rb) has two naturally occurring isotopes. The average mass of Rb is 85.4678 u. If 72.15% of Rb is found as Rb-85 (84.9117 u), what is the mass of the other isotope?

   a. 0.56 u
   b. 85.68 u
   c. 86.91 u
   d. 86.02 u
   e. 83.47 u

4. When strongly heated, boric acid breaks down to boric oxide and water. What mass of boric oxide is formed from the decomposition of 15.0 g $B(OH)_3$? $2\ B(OH)_3(s) \rightarrow B_2O_3(s) + 3\ H_2O(g)$

   a. 7.50 g
   b. 15.0 g
   c. 8.44 g
   d. 16.9 g
   e. 33.8 g

5. Which formula represents the compound formed by aluminum and carbonate ions?

   a. $AlCO_3$
   b. $Al(CO_3)_2$
   c. $Al(CO_3)_3$
   d. $Al_2(CO_3)_3$
   e. $Al_3(CO_3)_2$

6. What is the correct formula for barium nitrate?

   a. $Ba(NO_3)_2$
   b. BNO2
   c. $Ba(NO_2)_2$
   d. BaN
   e. $BaNO_3$

7. What is the correct formula for cobalt(III) oxide?

   a. CoO
   b. Co_3O
   c. Co_3O_2
   d. Co_2O_3
   e. CoO_3

8. Which of the following formulas is not correct?

   a. $Al_2(SO_4)_3$
   b. $NaClO_3$
   c. Ba_2O_3
   d. $Mg(NO_3)_2$
   e. KH_2PO_4

9. What is the molar mass of cobalt(II) iodide hexahydrate?

   a. 212.8 g/mol
   b. 293.9 g/mol
   c. 312.7 g/mol
   d. 420.8 g/mol
   e. 465.1 g/mol

10. Ammonia is prepared by reacting nitrogen and hydrogen gases at high temperature according to the ***unbalanced*** chemical equation: **__ $N_2(g)$ + __ $H_2(g)$ → __ $NH_3(g)$** What are the respective coefficients when the equation is balanced with the smallest whole numbers?

    a. 1, 1, 1
    b. 1, 3, 1
    c. 1, 3, 2
    d. 2, 1, 2
    e. 2, 3, 2

11. When methanol undergoes complete combustion, the products are carbon dioxide and water: **__ $CH_3OH(l)$ + __ $O_2(g)$ → __ $CO_2(g)$ + __ $H_2O(g)$** What are the respective coefficients when the equation is balanced with the smallest whole numbers?

    a. 1, 1, 1, 1
    b. 1, 2, 1, 2
    c. 2, 2, 2, 4
    d. 2, 3, 2, 4
    e. 2, 4, 6, 4

12. What is the net ionic equation for the reaction of aqueous lead(II) nitrate with aqueous sodium bromide?

    a. $Pb(NO_3)_2(aq) + 2\ NaBr(aq) \rightarrow PbBr_2(aq) + 2\ NaNO_3(s)$
    b. $Na^+(aq) + NO_3^{-1}(aq) \rightarrow NaNO_3(s)$
    c. $Pb^{2+}(aq) + 2\ Br^{-1}(aq) \rightarrow PbBr_2(s)$
    d. $Pb^{2+}(aq) + 2\ Na^+(aq) \rightarrow PbNa_2(s)$
    e. $Pb(NO_3)_2(aq) + 2\ NaBr(aq) \rightarrow PbBr_2(s) + 2\ NaNO_3(aq)$

13. Which of the following are oxidation-reduction reactions?

    1. $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$
    2. $Pb(ClO_4)_2(aq) + 2\ KI(aq) \rightarrow PbI_2(s) + 2\ KClO_4(aq)$
    3. $CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$

    a. 1 only
    b. 2 only
    c. 1 and 2
    d. 1 and 3
    e. 2 and 3

14. What is the oxidation number of each atom in potassium nitrate, KNO_3?

    a. K = +1, N = -3, O = -2
    b. K = +1, N = +5, O = -2
    c. K = +1, N = -3, O = +2
    d. K = -1, N = +3, O = -2
    e. K = 0, N = 0, O = 0

15. If 0.3000 g of impure soda ash (Na_2CO_3) is titrated with 17.66 mL of 0.1187 M HCl, what is the percent purity of the soda ash?

    $$Na_2CO_3(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + H_2O(l) + CO_2(g)$$

    a. 11.11%
    b. 22.22%
    c. 57.91%
    d. 37.03%
    e. 74.06%

16. If 2.891 g $MgCl_2$ is dissolved in enough water to make 500.0 mL of solution, what is the molarity of the magnesium chloride solution?

    a. 5.782×10^{-3} M
    b. 1.518×10^{-2} M
    c. 6.073×10^{-2} M
    d. 0.5505 M
    e. 5.782 M

17. How many liters of 0.1107 M KCl(aq) contain 15.00 g of KCl?

    a. 0.02227 L
    b. 0.5502 L
    c. 1.661 L
    d. 1.818 L
    e. 123.8 L

18. When 27.0 g of an unknown metal at 88.4 °C is placed in 115 g H_2O at 21.0 °C, the final temperature of the water is 23.7 °C. What is the specific heat capacity of the metal?

    a. 0.34 J/g·K
    b. 0.51 J/g·K
    c. 0.74 J/g·K
    d. 0.94J/g·K
    e. 1.4 J/g·K

19. Calculate the amount of heat required to change 50.0 g ice at –20.0 °C to steam at 135 °C. (Heat of fusion = 333 J/g; heat of vaporization = 2260 J/g; specific heat capacities: ice = 2.09 J/g·K, steam = 1.84 J/g·K)

    a. 4.18 kJ
    b. 32.4 kJ
    c. 78.8 kJ
    d. 135 kJ
    e. 156 kJ

20. Hydrazine, N_2H_4, is a liquid used as a rocket fuel. It reacts with oxygen to yield nitrogen gas and water: **$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2\ H_2O(l)$** The reaction of 3.80 g N_2H_4 evolves 73.7 kJ of heat. Calculate the enthalpy change per mole of hydrazine combusted.

    a. -8.74 kJ/mol
    b. -19.4 kJ/mol
    c. -2.80×10^2 kJ/mol
    d. -622 kJ/mol
    e. -8.98×10^3 kJ/mol

21. Which of the following chemical equations corresponds to the standard molar enthalpy of formation of N_2O?

    a. $NO(g) + 1/2\ N_2(g) \rightarrow N_2O(g)$
    b. $N_2(g) + 1/2\ O_2(g) \rightarrow N_2O(g)$
    c. $2N(g) + O(g) \rightarrow N_2O(g)$
    d. $N_2(g) + O(g) \rightarrow N_2O(g)$
    e. $2\ N_2(g) + O_2(g) \rightarrow 2\ N_2O(g)$

22. Determine ΔH for the reaction: **$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$** given the thermochemical equations below.

    $N_2(g) + O_2(g) \rightarrow 2\ NO(g)$ $\Delta H = +180.8$ kJ
    $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$ $\Delta H = -906.2$ kJ
    $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$ $\Delta H = -483.6$ kJ

    a. -1209.0 kJ
    b. -1189.0 kJ
    c. -756.5 kJ
    d. -241.8 kJ
    e. -91.5 kJ

23. An argon ion laser emits light at 488 nm. What is the frequency of this radiation?

    a. $4.07 \times 10^{-19}\ s^{-1}$
    b. $1.63 \times 10^{-15}\ s^{-1}$
    c. $1.46 \times 10^{2}\ s^{-1}$
    d. $2.05 \times 10^{6}\ s^{-1}$
    e. $6.14 \times 10^{14}\ s^{-1}$

24. A microwave oven emits radiation with an energy of 3.98×10^{-23} J/photon. What is the frequency of this radiation?

    a. $1.67 \times 10^{-11}\ s^{-1}$
    b. $6.67 \times 10^{-7}\ s^{-1}$
    c. $2.00\ s^{-1}$
    d. $1.50 \times 10^{6}\ s^{-1}$
    e. $6.01 \times 10^{10}\ s^{-1}$

25. What is the energy of a mole of photons of red light with a wavelength of 632 nm?

    a. 189 kJ
    b. 252 kJ
    c. 314 kJ
    d. 515 kJ
    e. 756 kJ

26. For a neutron (mass = 1.675×10^{-27} kg) moving with a velocity of 5.2×10^{3} m/s, what is the de Broglie wavelength?

    a. 7.6×10^{-11} m
    b. 4.5×10^{-9} m
    c. 2.1×10^{-6} m
    d. 486 m
    e. 1.3×10^{10} m

27. What type of orbital is designated n = 3, l = 2, m_l = -1 and $m_s = +^1/_2$?

    a. 3s
    b. 3p
    c. 3d
    d. 2f
    e. 2d

28. Which of the following is a possible set of quantum numbers for an electron in an atom?

    a. n = 1, l = 1, m_l = 1
    b. n = 2, l = 0, m_l = -1
    c. n = 0, l = 0, m_l = 0
    d. n = 3, l = 1, m_l = -1
    e. n = 4, l = 5, m_l = -2

29. What is the maximum number of orbitals that can be identified with the following quantum numbers: n = 3, l = 1, m_l = 0?

    a. 0
    b. 1
    c. 3
    d. 5
    e. 7

30. Which of the following particles would be most paramagnetic?

    a. Se
    b. Cd
    c. Ar
    d. He
    e. Ca

31. Place the following atoms in order of increasing atomic radii: Se, O, S, and As.

    a. O < S < Se < As
    b. O < S < As > Se
    c. As < Se < S < O
    d. Se < As < S < O
    e. S < As < O < Se

32. What is the ground state electron configuration for Cr^{3+}?

    a. [Ar]
    b. [Ar]$3d^34s^2$
    c. [Ar]$3d^44s^1$
    d. [Ar]$3d^3$
    e. [Ar]$3d^74s^2$

CH 221 Sample Final Exam
Name: ____ Lab Section: ______
Answers

Part I: Multiple Choice Questions

1. E
2. C
3. C
4. C
5. D
6. A

7. D
8. C
9. D
10. C
11. D
12. C

13. A
14. B
15. D
16. C
17. D
18. C

19. E
20. D
21. B
22. E
23. E
24. E
25. A

26. A
27. C
28. D
29. B
30. A
31. A
32. D