The Chemistry 222 Companion

Lab Manual, Problem Sets, Lecture Slides and Learning Resources

Dr. Michael A. Russell
Mt. Hood Community College
Winter 2025

Welcome to Chemistry 222!

My name is **Dr. Michael Russell** and I am pleased that you have decided to take Chemistry 222 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 222 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 PowerPoint 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 222 website (http://mhchem.org/222/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,

Dr. Michael Russell
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AC 2568 - *office*

CHEMISTRY 222 TABLE OF CONTENTS

CH 222 Laboratory Experiments - *http://mhchem.org/lab222*

Labs for Sections 01 and H1:

Labs for Section W1:

CH 222 Problem Sets and Exam Prep Worksheets - *http://mhchem.org/pset222*

CH 222 Lecture Slides - *http://mhchem.org/ho222*

CH 222 Handouts - *http://mhchem.org/ho222*

Chapter Seven:

Chapter Eight:

Chapter Twenty:

Chapter Nine:

Chapter Ten:

Chapter Eleven:

Chapter Twelve:

Chapter Twenty-one:

CH 222 Concept Guides - *http://mhchem.org/ho222*

CH 222 Practice Problem Sets and Quizzes *(with answers)* - *http://mhchem.org/222*

More sample quizzes and exams online - http://mhchem.org/222/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") <u>and</u> 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:

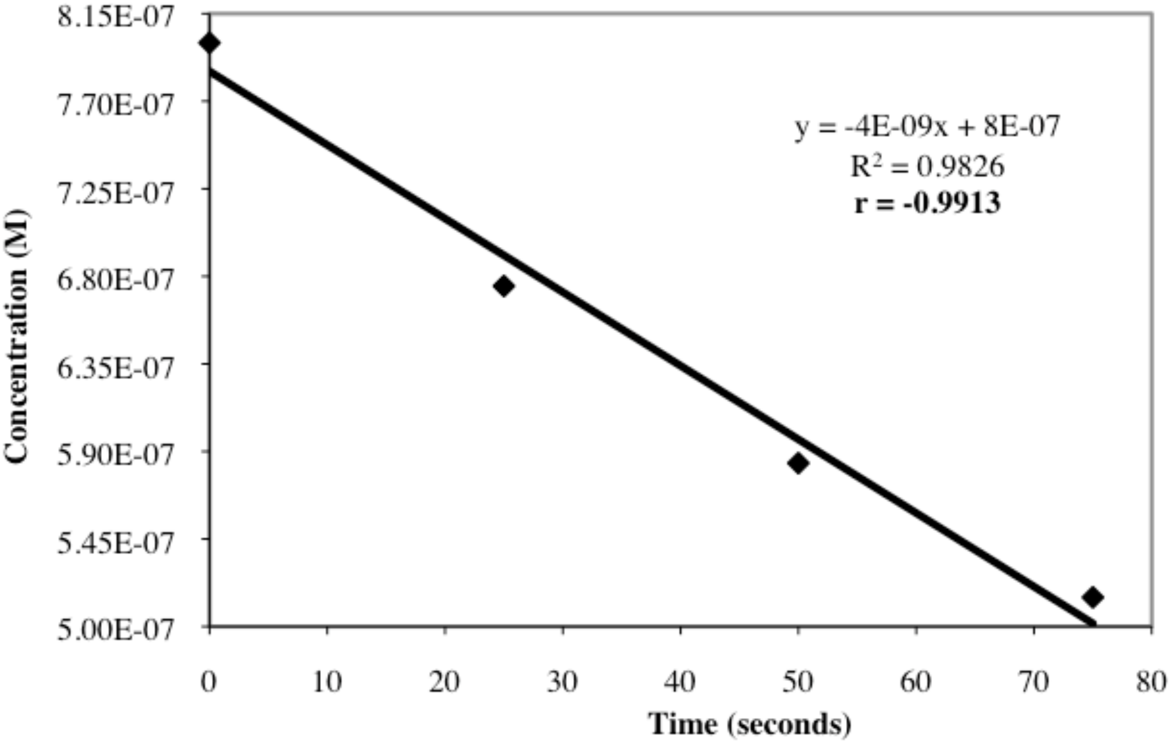
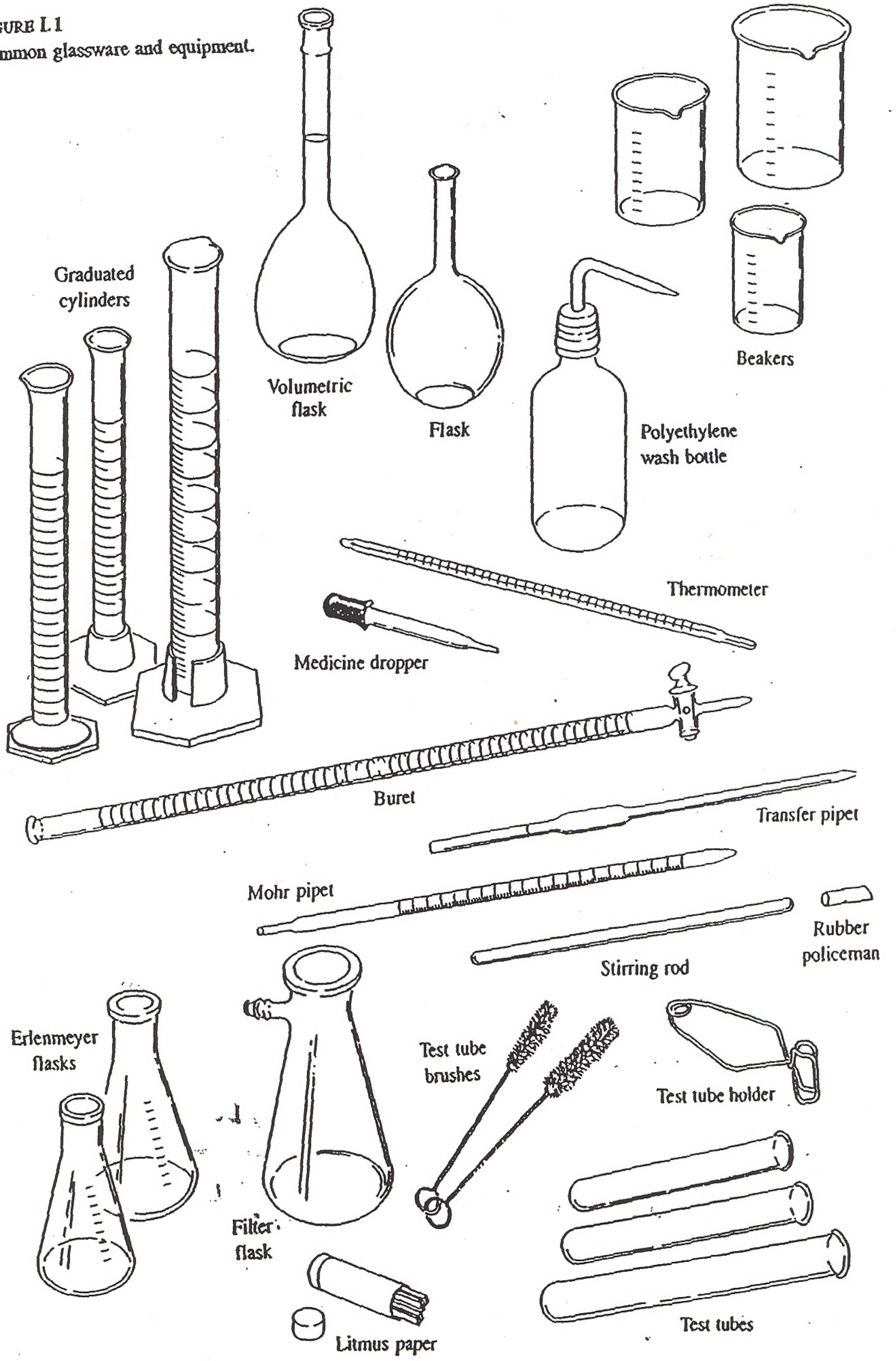
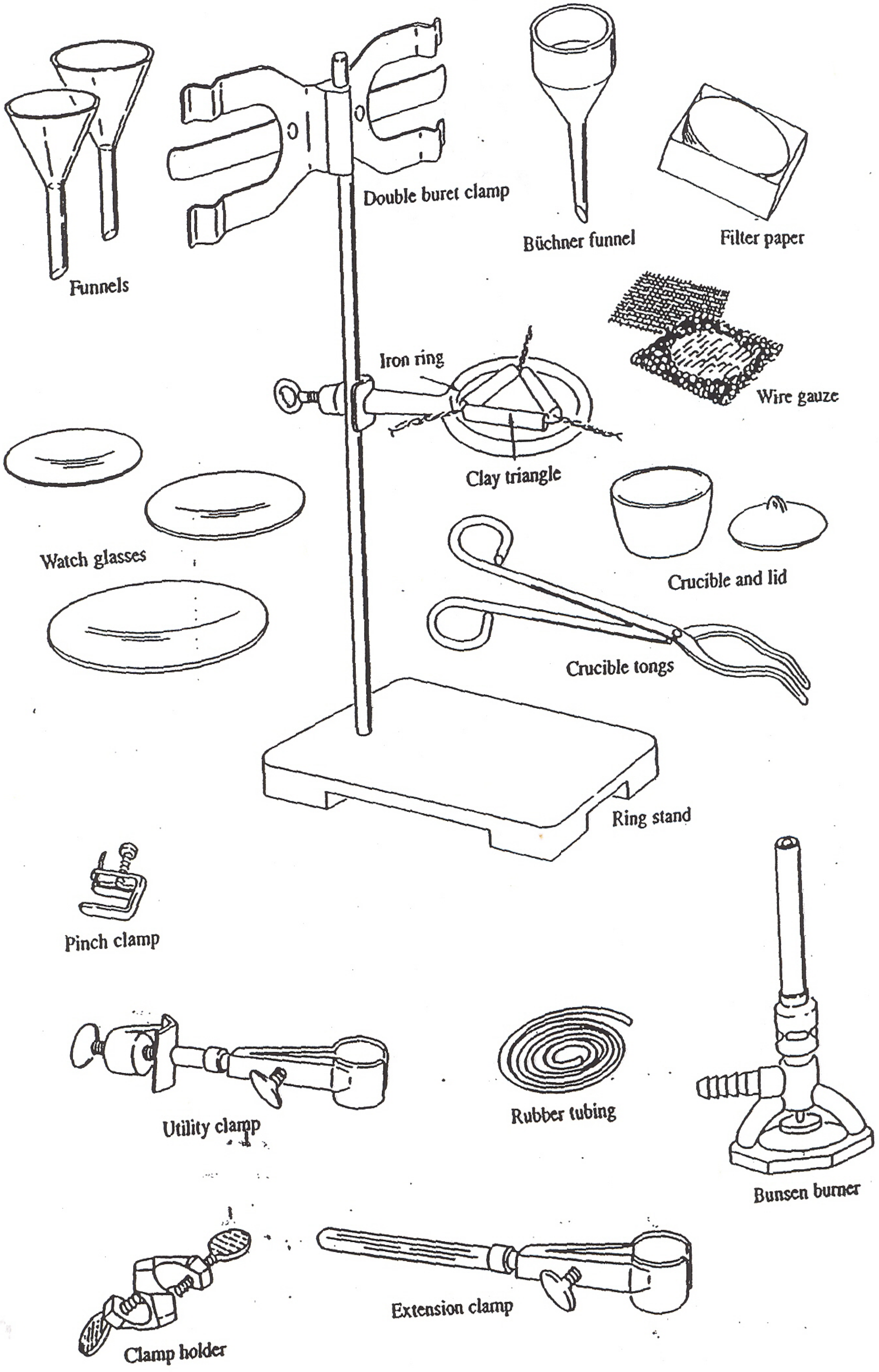

8.15E-07
7.70E-07
7.25E-07
6.80E-07
6.35E-07
5.90E-07
5.45E-07
5.00E-07
Concentration (M)
0
10
20
30
40
50
60
70
80
Time (seconds)
y = -4E-09x + 8E-07
R2 = 0.9826
r = -0.9913

FIGURE I.1
Common glassware and equipment.




Funnels
Double buret clamp
Büchner funnel
Filter paper
Wire gauze
Iron ring
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 VIII 8	9 VIII 8	10 VIII 8	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

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

- 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 %**
- Deviation$_1$ = | 35.89 – 35.72 | = 0.17
- Deviation$_2$ = | 35.89 – 35.92 | = 0.03
- 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.

Winter 2025 Chemistry 222 *with* Dr. Michael A. Russell

For Section 01, Section H1 and Section W1

CH 222, 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 222 Discord Server: **https://discord.gg/ZhrmjCS**

Chemistry 222 website:
http://mhchem.org/222

Required/Recommended Materials*:*
- * "Chemistry" *by* The OpenStax College (978-1-947172-62-3), *available here for free:* **http://mhchem.org/text/OpenStaxChem.pdf**
- * **Chemistry 222 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. This second term covers molecular bonding and properties, gases, liquids, solids, physical states and changes of state, solutions, kinetics, and nuclear chemistry. ***Prerequisites****:* RD090, WR090 and MTH020, each with a grade of "C" or better, or placement above stated course levels; and CH221 with a grade of "C" or better.

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*
	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**: 90-100% **B**: 80-89% **C**: 67-79% **D**: 57-66% **F**: less than 57%

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 222 Winter 2025

All assignments *can be found on our* ***website*** *(http://mhchem.org/222)*

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	Lab Assignment
1	1/6 - 1/10	Introduction *to the course* **01/H1** *Lab:* “Chromatography” (Lab #1) *due next week in recitation* **W1** *Lab:* “Introduce Yourself" (Lab #1) *due Friday, January 10 by 9 AM via email*
2	1/13 - 1/17	*Due:* Problem set #1 *Chapter 7;* **01, H1:** *due in recitation;;* **W1:** *due 1/15 by 11:59 PM via email* *Due:* Quiz #1; **01, H1:** *take in recitation;* **W1:** *due 1/17 by 9 AM via email* **01/H1** *Lab:* "Lewis Structures (in class)“ (Lab #2) *due next week:* **01:** *due in lecture AC 1303 1/22 at 9 AM (MLK day);* **H1:** *due 1/22 in recitation* **W1** *Lab:* "Lewis Structures (online)“ (Lab #2) *due Wednesday, January 22 by 11:59 PM*
3	1/20 - 1/24	**Monday, January 20: MLK day, all classes, office hours canceled, Sec. 01 schedule changes!** *Due:* Problem set #2 *Chapter 8;* **01:** *due in lecture AC 1303 1/22 at 9 AM (MLK day);* **H1:** *due in recitation;* **W1:** *due 1/22 by 11:59 PM* *Due:* Quiz #2; **01:** *get in lecture AC 1303 1/22, due Friday 1/24 at 9 AM AC 1303 (MLK);* **H1:** *take in recitation;* **W1:** *due 1/24 by 9 AM* **H1** *Lab:* "Valence Bond and Molecular Orbitals (in class)“ (Lab #3) *due next week in recitation* **01/W1** *Lab:* "Valence Bond and Molecular Orbitals Lab (online)“ (Lab #3) **01:** *due 1/27 in recitation AC 2501 at 1:10 PM (MLK day);* **W1:** *due Wednesday, Jan. 29 by 11:59 PM* **January 24, 9 AM:** *Last chance to reserve a Class Presentation compound*
4	1/27 - 1/31	*Due:* Problem set #3 *Chapter 20;* **01, H1:** *due in recitation;* **W1:** *due 1/29 by 11:59 PM* *Due:* Quiz #3; **01, H1:** *take in recitation;* **W1:** *due 1/31 by 9 AM* **01/H1** *Lab:* "Organic Chemistry (in class)“ (Lab #4) *due next week in recitation* **W1** *Lab:* "Organic Chemistry (online)“ (Lab #4) *due Wednesday, February 5 by 11:59 PM*
5	2/3 - 2/7	**EXAM #1** *(Chapters 7, 8 & 20);* **01, H1***: take in recitation;* **W1***: due 2/7 by 9 AM* *Due:* "Exam Prep I”*;* **01, H1:** *due in recitation;* **W1:** *due 2/5 by 11:59 PM* **01/H1** *Lab:* "Molar Mass of a Volatile Liquid (in class)“ (Lab #5) *due next week in recitation* **W1** *Lab:* "Molar Mass of a Volatile Liquid (online)“ (Lab #5) *due Wednesday, February 12 by 11:59 PM*
6	2/10 - 2/14	*Due:* Problem set #4 *Chapter 9 & 10*; **01, H1:** *due in recitation;* **W1:** *due 2/12 by 11:59 PM* *Due:* Quiz #4; **01, H1:** *take in recitation;* **W1:** *due 2/14 by 9 AM*

Week 6 continued on next page

6	*continued*	**01/H1** *Lab:* "Linear Regression / Structure of Solids (in class)“ (Lab #6) *due next week in recitation* **W1** *Lab:* "Linear Regression / Structure of Solids (online)“ (Lab #6) *due Wednesday, February 19 by 11:59 PM* *Due:* **Class Presentation Rough Draft Paper***;* **01, H1**: *due in recitation;* **W1:** *due 2/12 by 11:59 PM*
7	2/17 - 2/21	*Due:* Problem set #5 *Chapter 10 & 11; ;* **01, H1:** *due in recitation;* **W1:** *due 2/19 by 11:59 PM* *Due:* Quiz #5; **01, H1:** *take in recitation;* **W1:** *due 2/21 by 9 AM* **01/H1** *Lab:* "Molar Mass Determination by Freezing Point Depression (in class)“ (Lab #7) *due next week in recitation* **W1** *Lab:* "Molar Mass Determination by Freezing Point Depression (online)“ (Lab #7) *due Wednesday, Feb. 26 by 11:59 PM* **February 21:** *Last day to drop or change grade status this quarter at Mt. Hood Community College*
8	2/24 - 2/28	**EXAM #2** (*Chapters 9-11);* **01, H1***: take in recitation;* **W1***: due 2/28 by 9 AM* *Due:* "Exam Prep II”*;* **01, H1:** *due in recitation;* **W1:** *due 2/26 by 11:59 PM* **01/H1** *Lab:* "Kinetics I - The Iodination of Acetone (in class)“ (Lab #8) *due next week in recitation* **W1** *Lab:* "Kinetics I - The Iodination of Acetone (online)“ (Lab #8) *due Wednesday, March 5 by 11:59 PM*
9	3/3 - 3/7	**CLASS PRESENTATIONS WEEK** **01/H1:** *Class Presentation paper due at time of presentation during recitation* **W1:** *Class Presentation paper and video due Wednesday, March 5 by 11:59 PM*
10	3/10 - 3/14	*Due:* Problem set #6 *Chapter 12 & 21;* **01, H1:** *due in recitation;* **W1:** *due 3/12 by 11:59 PM* *Due:* Quiz #6; **01, H1:** *take in recitation;* **W1:** *due 3/14 by 9 AM* **01/H1** *Lab:* "Kinetics II - The Iodination of Acetone (in class)“ (Lab #9) *due next week Wednesday during Finals* **W1** *Lab:* "Kinetics II - The Iodination of Acetone (online)“ (Lab #9) *due Wednesday, March 12 by 11:59 PM* *All extra credit closes Friday, March 14 at 9 AM*
11	3/17 - 3/19	**Take Home Lab Final** *released by 9 AM Monday, 3/17 for all CH 222 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, March 19 at 8:45 AM in AC 1303. *Due:* **Final Exam Prep** worksheet, **Take Home Lab Final, Kinetics II Lab** *Section H1: Take* **Lecture Final** *tentatively* on Wednesday, March 19 at 1:10 PM in AC 2501. *Due:* **Final Exam Prep** worksheet, **Take Home Lab Final, Kinetics II Lab** *Section W1: Due:* **Lecture Final** *(available Monday March 17)*, **Final Exam Prep** worksheet and **Take Home Lab Final** *on* Wednesday, March 19 by 11:59 PM

Getting Started in Chemistry 222

Welcome to Chemistry 222! I am glad to have you enrolled in CH 222! 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/ZhrmjCS) during office hours. I try to respond to student inquiries within 24 hours

- There are **three sections of CH 222** 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 222 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 222** 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 222!

- **Check your email often during Chemistry 222**. I will be sending weekly reminders as to “what is due this week” in CH 222 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 222 website** is worth exploring. The Chemistry 222 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/222

  You should see the CH 222 website on your screen.

- Check out the **Chemistry 222 Chapter Guides** by selecting "**Chapter Guides**" from the upper left hand corner of the CH 222 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**. You will be giving a five minute presentation this term on a **compound**, and you must reserve your compound choice with me. A written paper also accompanies the presentation on your compound. To reserve your compound, email your preferred element choices directly to the instructor, or: **http://mhchem.org/cp222**

  The "**Class Presentations FAQ**" (available in the syllabus or here: **http://mhchem.org/cp222info**) has more information.

- The **Chemistry 222 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 222! 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 222 CLASS PRESENTATIONS FAQ

FAQ = Frequently Asked Questions

When: **Monday March 3 (section 01) or Wednesday March 5 (sections H1 and W1)**

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

Who: ***Everyone* enrolled in CH 222 (All Sections)**

What topic should I pick?

For CH 222, the topic will be **compounds**. Pick a compound you find interesting and write a report on the topic. Since there are millions of compounds, every student must pick a different compound. **Reserve** your compound using the online form at **http://mhchem.org/cp222**

Once your compound has been chosen, begin researching interesting information on the compound using the library, internet, etc. You will be preparing a paper on the compound 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 a compound early and start researching promptly. Also, if you still have not reserved a compound 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 compound. 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 compound. **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

During 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 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/rd2) The Class Presentation Rough Draft paper is worth 20 points (out of 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 <u>six slides</u>; 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 peer review process on time.

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/cp222info). *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 222 website (http://mhchem.org/cgr2) The rubric will allow you to look at the items deemed most important when grading your Class Presentation.

How do I get started?

<u>*Step 1*</u>*: Reserve your Class Presentation Compound*

Decide on some compounds that interest you, then email the instructor or complete the online web form to reserve your compound: **http://mhchem.org/cp222**

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 compounds in case your first choice has already been claimed; he can also pick one for you if you are uncertain which compound to pick. Reserve your class presentation compound by the end of the third week, **January 24** at **9 AM**. You can see which compounds are still available here: **http://mhchem.org/222av**

<u>*Step 2*</u>*: 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 using the handout (http://mhchem.org/rd2) at the beginning of your paper. Deadline: **Mon., Feb.. 10 (01)** or **Wed. Feb.. 12 (H1 & W1)**

<u>*Step 3*</u>*: 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, March 3.** *Section H1:* **Wednesday, March 5.**

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 Wed., March 5 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/cp222info) or contact the instructor.

Name: ______________________________

CH 222 CLASS PRESENTATIONS ROUGH DRAFT PAPER

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

Lab Section:

Reserved Compound:

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 222 Class Presentation Frequently Asked Questions handout:	**http://mhchem.org/faq2**
The CH 222 Citation Guide:	**http://mhchem.org/cg**

Class Presentation Rough Draft Paper Due Dates:

Section 01:	**February 10 at 1:10 PM**
Section H1:	**February 12 at 1:10 PM**
Section W1:	**February 12 at 11:59 PM**

Section 01: *The final Class Presentation paper is due at the time of your presentation on Monday, March 3.*

Section H1: *The final Class Presentation paper is due at the time of your presentation on Wednesday, March 5.*

Section W1: *The final Class Presentation paper and recorded presentation will be due by 11:59 PM on Wednesday, March 5.*

Staying Connected in Chemistry 222 This Quarter

Success in Chemistry 222 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/ZhrmjCS**) and join the CH 222 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 222 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 222 website (https://mhchem.org/222/classroom/ci.htm).

To access these materials (and more!), go to our website (**http://mhchem.org/222)** 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 222** - 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/222

CH 222 Winter 2025:

"Chromatography" *(in class)* ***Lab*** *- Instructions*

Note: **This is the lab for section 01 and H1 of CH 222 only.**

- *If you are taking section W1 of CH 222*, please use this link:
  http://mhchem.org/r/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-1-6 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, January 6 (section 01) *or* **Wednesday, January 8 (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-1-5 through Ia-1-6 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, January 13 (section 01)** *or* **Wednesday, January 15 (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!

Chromatography

Most of the substances we use everyday are mixtures of pure substances. Separating, detecting and identifying all of the components in a mixture is crucial for the successful chemist. Several techniques have been developed to do this all of which depend on the differing chemical and physical properties of the components in the mixture. Chromatography is a separation and identification technique that takes advantage of the difference in solubility of a pure substance in various solvents.

In a chromatographic separation, a mixture is deposited on a solid adsorbing substance called the *stationary phase*. The stationary phase can be a strip of filter paper, a thin film of silica gel on an inert surface, a column of silica gel, or a tube of small beads coated with a high molecular weight oil. A solvent is allowed to flow through the stationary phase either under pressure or by gravity or capillary action. As the solvent passes over the mixture, the components in the mixture dissolve in the solvent. A competition takes place between the adsorption of a component on the stationary phase and the dissolution of the component in the solvent (the *mobile phase*). The affinity of each of the components to the stationary phase or mobile phase will be different leading to a separation.

The name given to the various types of chromatography is based upon the type of stationary phase or the physical state of the mixture. Examples include column chromatography, paper chromatography, thin layer chromatography, vapor-phase chromatography, and high pressure liquid chromatography (HPLC).

In this experiment we will use **paper chromatography** to separate a mixture of metallic ions in an aqueous solution. A piece of filter paper is spotted with a drop of solution containing a mixture of ions, and the paper is allowed to dry. The paper is then suspended in a beaker (or jar) containing a solvent which moves up the paper by capillary action. Because each component of a mixture has its own characteristic affinities, each metal ion will travel up the paper at its own characteristic rate. If the paper is large enough, all the components will be separated and will appear as separate spots. If the components are highly colored the spots will be visible. You can convert weakly colored or colorless spots to highly colored ones by spraying them with substances that react to form colored compounds. The filter paper will now contain a vertical row of colored spots arranged according to their characteristic rate of ascent. The word *chromatography*, which is derived from to Greek words and literally means "written with color," was coined to describe this phenomenon.

The distance traveled by a component ("spot"), $\mathbf{D_x}$, with respect to the distance traveled by the eluting solvent, $\mathbf{D_s}$, is called the **retention factor**, $\mathbf{R_F}$:

$$R_f = \frac{D_x}{D_s}$$

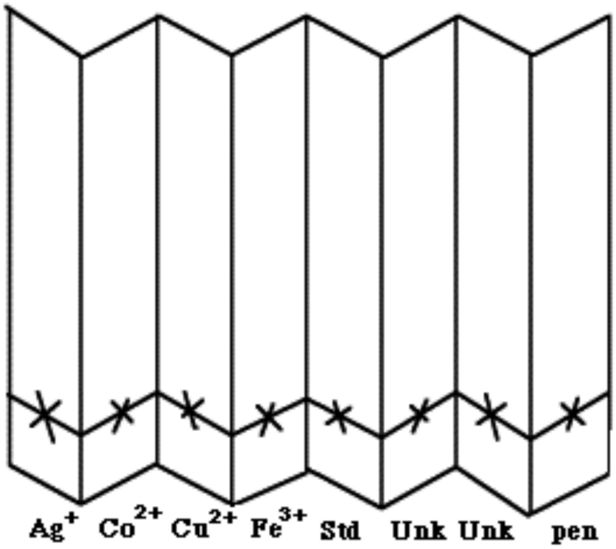
The R_F value is a characteristic property of a given component in a given solvent at a given temperature. Changing solvent or temperature implies a change in R_F. *Note* that the R_F will be a unitless number, and both D_x and D_s must be recorded in the same unit.

In this experiment you will use both the color of the spot and the calculated R_F value to identify the types of metal ions that are present in the solution.

PROCEDURE:

1. Obtain a piece of filter paper about 20 cm long by 12 cm wide. Placing the paper on a clean surface, draw a line using a straight edge and a pencil, 2 cm from one of the long edges.

2. Fold the filter paper so that the line that you have drawn is bisected. In the same manner fold the filter paper in half a second and third time. The line will have been divided into eight equal segments. Refold the paper so that it looks like an accordion, as shown below. Mark the center of each segment of the line with an X using pencil, and label each segment as shown below.



3. Obtain a small capillary tube. Using distilled water, practice spotting a scrap piece of filter paper. The maximum diameter of an acceptable spot is no larger than 0.5 cm.

4. When this procedure is satisfactory, use the appropriate known solutions of silver, cobalt, etc. to spot your chromatographic paper (2 drops each.) The standard (**Std**) should contain a spot from each of the four known solutions (four drops of liquid total). Be sure to allow each spot to dry before spotting it again. Hair dryers will facilitate a complete drying of the spots.

5. Each group will have two unknown (**Unk**) solutions to identify the ions present - be sure to write down their identity (letter) for the lab report. The unknown solutions are less concentrated than the metal ion solutions so you need to spot these two spots three times each.

6. For the last spot (**pen**), place a small dot from a felt pen on the X.

7. Add about 25 mL of the eluting solution (the mobile phase) to a 600 mL beaker and cover with plastic film (alternatively, a jar with a plastic lid can be used.) The eluting solution was made by mixing a solution of HCl with ethanol and butanol, two organic solvents known as alcohols.

8. When all spots are dry, carefully place the paper in the 600 mL beaker (or jar.) It is important that the solvent is below the 2 cm line on the paper. Cover the mouth of the beaker with plastic film (or put a lid on the jar) and allow the solvent to move up the paper.

9. **While the experiment is proceeding**, you can **test the effect of the staining reagent** on the metal ion solutions. Spot a piece of filter paper with each of the four known metal ion solutions and dry the paper as before. Some of the spots will be colored at this point; note the color of the unstained knowns in your lab. In a fume hood, place the filter paper on a paper towel and spray the paper evenly with the staining solution, getting the paper moist but not really wet. The staining reagent is a mixture of solutions of potassium ferrocyanide and potassium iodide and forms colored compounds with the metal ions. Note the **color** of each known metal ion spot in your lab.

10. When the eluting solution has risen to about 4 cm from the top of the chromatographic paper, remove the paper and immediately mark the paper with a line that is the **solvent front**. Dry the paper. Circle any cations that are visible by virtue of their colors. Place the paper on a paper towel and spray it with the staining reagent. Note the colors present. Dry the paper and **circle** the boundary of each spot since they may fade with time.

11. Measure the distance (to the tenth of a millimeter) from the straight line where you applied the spots to the solvent front. This distance is D_s, the distance traveled by the solvent. Measure the distance from the beginning line to the center of each spot. This distance is D_x for each metal cation. Calculate R_f for each of the metal cations. Be sure to include one sample calculation set up in your lab.

12. Report the R_f value for each metal cation as well as its color, the identity of the cations in your unknown solutions and any possible sources for error in this experiment.

13. Dispose of any remaining eluting solution in a waste bottle. Wash your hands before leaving the lab!

Chromatography

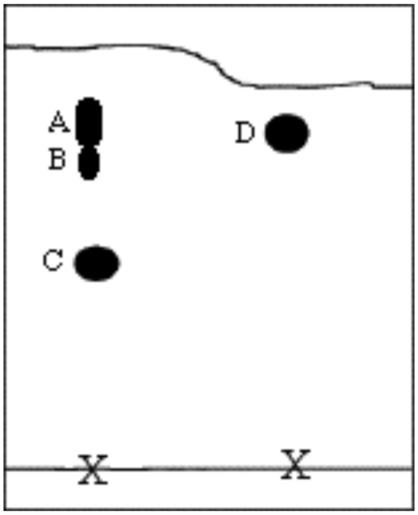
*include first **and** last names* **Name:**

Lab Partner(s):

	Ag^+	Co^{2+}	Cu^{2+}	Fe^{3+}
Known Solutions				
Colors (Dry)	______	______	______	______
Colors (After staining)	______	______	______	______
Distance solvent moved	______	______	______	______
Distance cation moved	______	______	______	______
R_f	______	______	______	______
Known Mixture (Std or Mix)				
Distance solvent moved	______	______	______	______
Distance cation moved	______	______	______	______
R_f	______	______	______	______
Unknown Mixture No. _______				
Color (Dry)	______	______	______	______
Color (After staining)	______	______	______	______
Distance solvent traveled	______	______	______	______
Distance cation traveled	______	______	______	______
R_f	______	______	______	______
Unknown Mixture No. _____				
Color (Dry)	______	______	______	______
Color (After staining)	______	______	______	______
Distance solvent traveled	______	______	______	______
Distance cation traveled	______	______	______	______
R_F	______	______	______	______

Chromatography Postlab Questions

1. Determine the R_f value for each spot on the following chromatogram.



R_f(A) =

R_f(B) =

R_f(C) =

R_f(D) =

2. Based on your calculations, is Compound D more likely to be identical to Compound A, Compound B, or Compound C? Explain.

3. Why are pencils used to mark the chromatographic paper rather than ink?

4. If a 1.0 cm^2 spot of silver solution contains 10. microliters of solution and the solution contains 11 g of Ag^+ ions per liter of solution, what mass of silver ions are in one 1.0 cm^2 spot?

Please note: The instructor will send you email throughout the term, so *please* c*heck 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 222 Winter 2025:
“Lewis Structures *(in class)”* ***Lab** - Instructions*

Note: **This is the lab for section 01 and H1 of CH 222 only.**

- *If you are taking section W1 of CH 222*, please use this link:
http://mhchem.org/r/2b.htm

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-2-7 through Ia-2-12 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, January 13 (section 01) *or* **Wednesday, January 15 (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-2-7 through Ia-2-12 *only* to avoid a point penalty.) ***Due dates:*** **<u>Section 01</u>: Wednesday, January 22 at 9 AM in AC 1303** *(MLK day)*; **<u>section H1:</u> Wednesday, January 22 at 1:10 PM in AC 2501.** 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!

Lewis Structures - The Geometry of Covalent Molecules

The formulas of many covalent compounds, especially those involving only the elements of the first few periods of the periodic table, were brought within the framework of the **octet rule**. The octet rule states that a total of eight **valence** electrons, either shared or unshared, should be in the region of each atom beyond the first period. For hydrogen the desired number is two. Electrons that are shared between two atoms are **bonding electrons** and are to be counted toward the octet of each of the bonding atoms. A single covalent bond consists of a pair of shared electrons, a double bond has two shared pairs, and a triple bond has three shared pairs. Bond distances are shorter and bond energies are greater for multiple bonds than for single bonds.

Structural formulas, such as shown in Figure 1, represent the electron distributions in covalent molecules and ions. These structures are not meant to indicate actual bond angles in threedimensional varieties; they merely show the number of bonds connecting the various atoms. In **Lewis formulas***,* a single line between two atoms represents a **pair** of shared electrons and a **dot** represents an unshared electron.

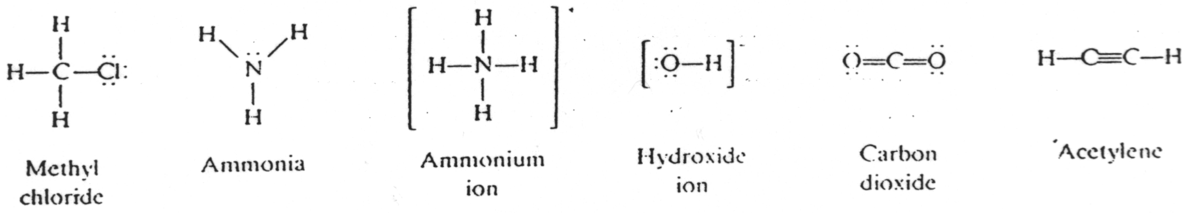
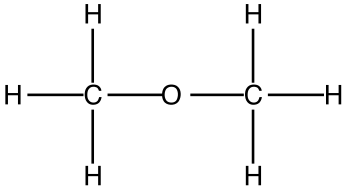

Methyl chloride
Ammonia
Ammonium ion
Hydroxide ion
Carbon dioxide
Acetylene

Figure 1

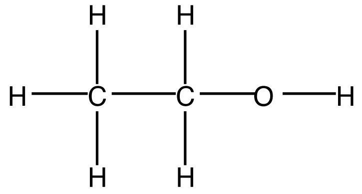
Two lines constitute a *double bond,* and *three lines* constitute a *triple bond.* The total number of electrons shown in such a molecular structure is equal to the sum of the numbers of valence (outershell) electrons in the free atoms: 1 for H, 4 for C, 5 for N, 6 for O, and 7 for Cl. For an ionic structure, one additional electron must be added to this sum for each unit of negative charge on the whole ion, as in OH^-, and one electron must be subtracted from the sum for each unit of positive charge on the ion, as in NH_4^+. The number of pairs of electrons shared by an atom is called its **covalence**.

The covalence of hydrogen is always one. The covalence of oxygen is practically always one or two. The covalence of carbon is four in almost all its stable compounds. Thus each carbon is expected to form either four single bonds, a double bond and two single bonds, two double bonds, or a single and a triple bond. Although the octet rule is not a rigid rule of chemical bonding, it is obeyed for C, N, O, and F in almost all their compounds. The octet is exceeded commonly for elements in the third and higher periods of the periodic table.

Isomerism: For some molecules with a given molecular formula, it is possible to satisfy the octet rule with different atomic arrangements. A simple example would be

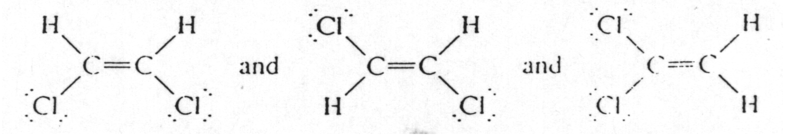


and



The two molecules are called **isomers** of each other, and the phenomenon is called **isomerism**. Although the molecular formulas of both substances are the same, C_2H_6O, their properties differ because of the different atomic arrangement. Since several different types of isomerism can exist especially in organic molecules, these are called **structural isomers**.

When **double bonds** are present, isomerism can occur in very small molecules. For example:



These isomers (called **geometric** isomers) result from the fact that there is no rotation around a double bond, as in single bonds.

Resonance: Sometimes more than one satisfactory structure can be written and there is no reason to select one over another. In such cases a single structural formula is inadequate to represent a substance correctly, and several such diagrams must be written. The true structure is then said to be a *resonance hybrid* of the several diagrams.

For example, experiment has shown that the two terminal oxygens in ozone are equivalent; that is, they are equidistant from the central oxygen. If only one of the resonance diagrams in Fig.2(a) were written, it would appear that one of the terminal oxygens is bonded more strongly to the central oxygen by a double bond than is the other (by a single bond) and that the more strongly bonded atom should be closer to the central atom. The hybrid of the two ozone structures gives equal weight to the extra bonding of the two terminal oxygen atoms. Similarly, the three resonance structures of carbonate in Fig.2(b) are needed to account for the experimental fact that all three oxygens are equidistant from the central carbon.

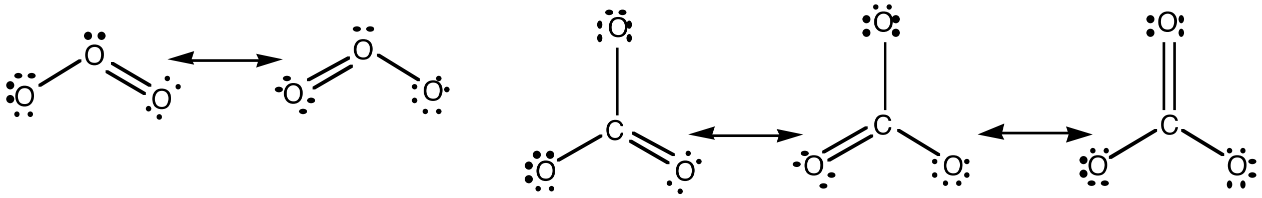
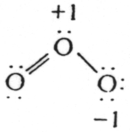


Figure 2: a) ozone *(left)* and b) carbonate *(right)*

The total bond energy of a substance for which resonance structures are written is greater than would be expected if there were only one formal Lewis structure. This additional stabilization is called **resonance energy**. It arises from the principle that is responsible for covalent bond energy, the **delocalization** of electrons about the atoms forming the bond. As a result of resonance in ozone, for example, the electrons constituting the second pair of the double bond are delocalized around the 3 oxygen atoms. The writing of two or more resonance structures is a way of overcoming the inability of a single valencebond structure to show this delocalization.

Formal Charge: Although a molecule as a whole is electrically neutral, it is a matter of much current interest to know whether there are local charges which can be identified with particular parts of a molecule, the algebraic sum of which would equal zero. In an ion, the algebraic sum would equal the charge of the ion as a whole. In one approximate method of apportioning charges within a molecule or ion, the shared electrons in a covalent bond are arbitrarily divided equally between the two atoms forming the bond. Unshared valence electrons on an atom are assigned exclusively to that atom. Each atom is then assigned a **formal charge** which is equal to the number of valence electrons possessed by that atom in the neutral free state minus the number of valence electrons assigned to it in the structure. These charges may be written near the atoms on the structural diagrams.

Figure 3 shows a single resonance structure for ozone. The central oxygen is assigned just five electrons (two in the unshared pair plus half of the three pairs in the bonds); this atom, being one electron short of the complement of six in a free oxygen atom, is thus assigned a formal charge of **+1**. The terminal oxygen connected by a single bond is assigned 7 (6 in the unshared pairs plus half of one pair in the bond); having one electron more than a neutral oxygen atom, this atom is assigned a formal charge of **-1**. The other terminal oxygen has no formal charge because six electrons are assigned to it (four in the unshared pairs plus half of the two shared pairs).

Figure 3:  *Formal charges on ozone*

A rule useful in choosing one Lewis structure over another is that **structures which minimize formal charge separation are favored**. Especially to be avoided are formal charges of magnitude greater than 1 and structures in which appreciable formal charges of the same sign are located on adjacent atoms.

Dipole moments and electronegativity: There are some experimental procedures that give information about the actual distribution of charges within a molecule (as distinct from the arbitrary assignment of formal charges). One such is the measurement of **dipole moment**. An electric dipole is a neutral object that has a positive charge of magnitude q and a separately located, equal but opposite negative charge. The rotation of a dipole by an electric field is dependent upon and a measure of the dipole moment, defined as the product of q and the distance, d, separating the positive and negative charges.

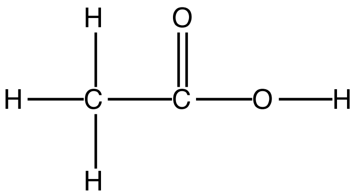
In a covalent diatomic molecule, the dipole moment would be expected to be zero if the bonding electrons were shared truly equally by the two atoms. This is indeed the case in molecules of the type XX where two identical atoms are bonded. In the more general type XY, two different kinds of atoms are bonded, and a dipole moment is usually observed. This is explained by hypothesizing that one of the atoms, say Y, has a greater attraction for the shared electrons in the bond than does X. Y is said to have a greater **electronegativity** than X. Electronegativity correlates with other atomic properties; in general, atoms with high ionization energies and/or high electron affinities tend to have high electronegativities. The most electronegative elements in order of decreasing electronegativity are **F > O > N = Cl** C is more electronegative than H. Metals are less electronegative than nonmetals.

HOW TO COMPLETE THIS LAB:

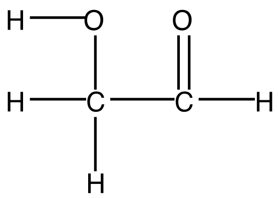
In this experiment you will be constructing models of molecules and ions containing atoms covalently bonded. If you have access to a model kit, use it! The model kit consists of plastic colored balls where each type of ball represents a certain species of atom; see the front cover of the model box to determine each atom type.

In assembling a molecular model of a molecule you should use a systematic approach keeping in mind the valence of the atoms in the molecule. Let us illustrate the recommended procedure with the following molecule, $\mathbf{C_2H_4O_2}$:

1. Draw a correct Lewis structure, remembering that carbon will *always* have four bonds and oxygen usually just two. Unless otherwise informed assume the carbon atoms form a carbon - carbon bond. Count up the total number of valence electrons (24 e^-). Those that are not bonding are nonbonding electrons. This leads to the following two isomers (other isomers exist as well):



and



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acetic acid

hydroxy acetaldehyde (unstable)

2. Connect the balls and sticks to according to the Lewis structure. Use two bonds to represent the carbon-oxygen double bond.

3. Determine the electron pair geometry (EPG) and molecular geometry (MG) of the molecule from the molecular model. For acetic acid we say that the molecule is tetrahedral at one carbon, trigonal planar at the other carbon and tetrahedral / bent at the oxygen.

4. Determine whether any isomers exist by changing the location of the atoms; however you must maintain the correct valence of each atom.

5. Resonance will most often exist when the central atom forms a double bond to one atom and a single bond to another atom of the same element.

PROCEDURE:

Use the procedure outlined above to construct Lewis structures of the molecules and ions listed below using the attached worksheet. Neatness counts! Be sure to **include all lone pair electrons** on your drawings. For *each* molecule:

i) Draw the Lewis structure, including all lone pair electrons.
ii) Describe the electron pair geometry and molecular geometry around the central atom. If more than one central atom, put a star (*) by the central atom you will use in your answer.
iii) Determine if the molecule is polar or nonpolar and list any bond angles.
iv) Draw the Lewis structures of isomers and resonance forms, if any.

1. CH_4
2. CH_2Cl_2
3. CH_4O
4. H_2O
5. BF_3
6. NH_4^{+1}
7. CH_2O
8. NH_3
9. H_2O_2
10. HNO_3
11. C_2H_2
12. SO_2
13. C_2H_4
14. $C_2H_2Cl_2$
15. $C_2H_4Br_2$
16. SCN^-
17. NO_3^{-1}
18. SO_4^{2-}
19. CO_2
20. SeF_4
21. $AsCl_3Br_2$
22. PCl_5
23. SiF_6^{2-}
24. $SiF_4Cl_2^{2-}$
25. $SiF_3Cl_3^{2-}$

Also, on the last page, construct six possible isomers of $C_6H_{12}O_2$.

Name: ____

Lewis Structures Worksheet - *Hand drawn versions only, no computer generated structures*

Be sure to include all lone pair electrons for full credit! Draw all resonance forms and isomers (if any).

Molecule/Ion	Lewis Structure	Geometry and Polarity	Resonance/Isomers? (draw)
CH_4		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
CH_2Cl_2		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
CH_4O *hint: put a star next to your central atom*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
H_2O		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
BF_3		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	

The Geometrical Structure Of Covalent Molecules Worksheet - *Continued*

Molecule/Ion	Lewis Structure	Geometry and Polarity	Resonance/Isomers? (draw)
NH_4^{+1}		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
CH_2O		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
NH_3		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
H_2O_2		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
HNO_3 *hint: trigonal planar*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	

The Geometrical Structure Of Covalent Molecules Worksheet - *Continued*

Molecule/Ion	Lewis Structure	Geometry and Polarity	Resonance/Isomers? (draw)
C_2H_2 *hint: triple bond*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
SO_2		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
C_2H_4 *hint: double bond*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
$C_2H_2Cl_2$		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
$C_2H_4Br_2$		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	

The Geometrical Structure Of Covalent Molecules Worksheet - *Continued*

Molecule/Ion	Lewis Structure	Geometry and Polarity	Resonance/Isomers? (draw)
SCN^{-1} *C in middle*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
NO_3^{-1}		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
SO_4^{-2}		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
CO_2		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
SeF_4		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	

The Geometrical Structure Of Covalent Molecules Worksheet - *Continued*

Molecule/Ion	Lewis Structure	Geometry and Polarity	Resonance/Isomers? (draw)
$AsCl_3Br_2$ *hint: trigonal bipyramid*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
PCl_5		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
SiF_6^{2-} *hint:octahedral*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
$SiF_4Cl_2^{2-}$		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
$SiF_3Cl_3^{2-}$		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	

Finally, **construct six possible isomers of $C_6H_{12}O_2$ in the space provided below**. For this question, carbon will always have four bonds and no lone pairs, and oxygen should always have two bonds (two single or one double) and two lone pairs.

CH 222 Winter 2025: **"Valence Bond and Molecular Orbitals** *(in class)"* ***Lab*** *- Instructions*

Note: **This is the lab for section H1 of CH 222 *only*.**

- *If you are taking section 01 or section W1 of CH 222*, please use this link: http://mhchem.org/r/3b.htm

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Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-3-3 through Ia-3-12 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 Wednesday, January 22 (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-3-5 through Ia-3-12 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Wednesday, January 29 (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!

VALENCE BOND (VB) THEORY

and

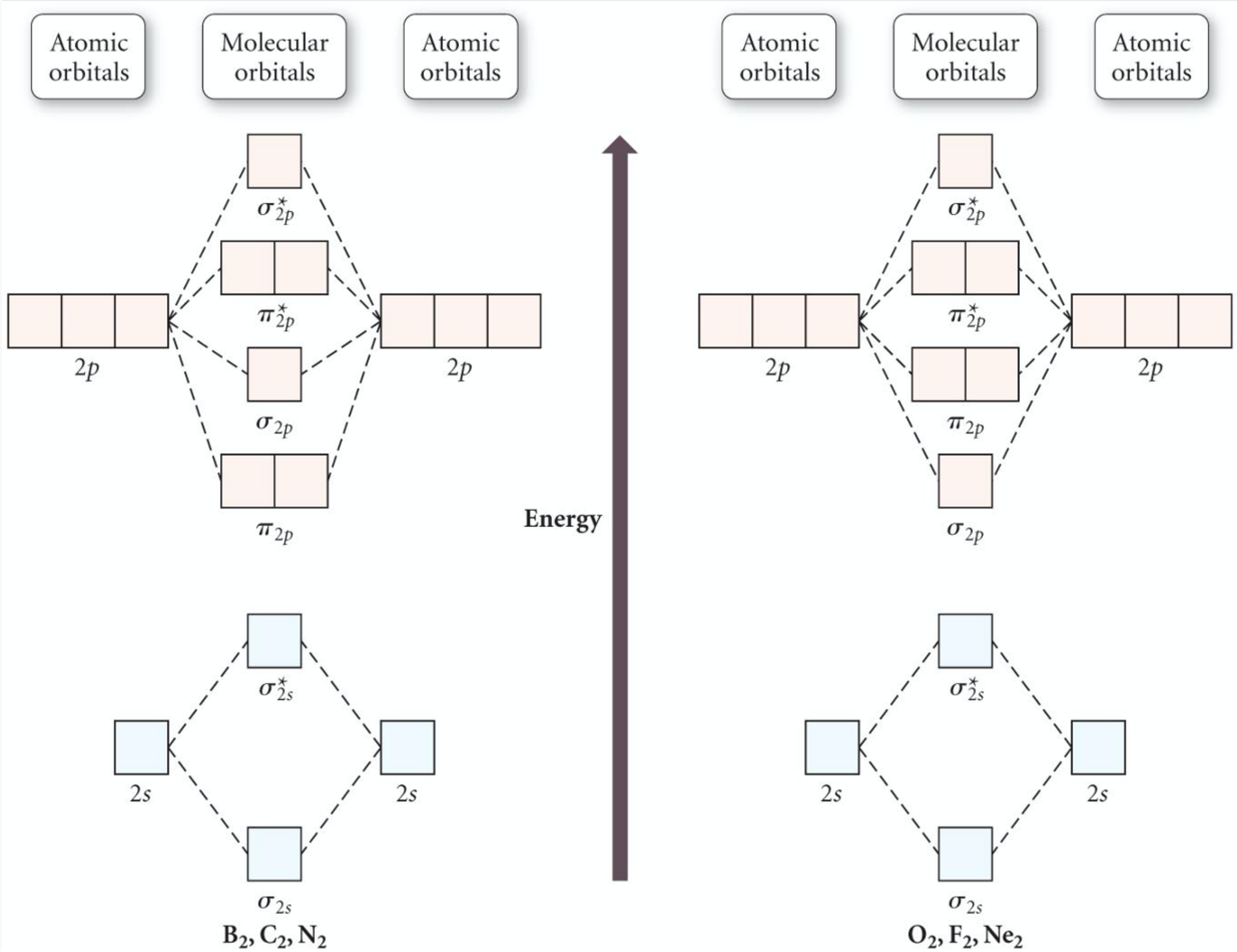
MOLECULAR ORBITAL (MO) THEORY

LAB

This lab contains a series of handouts that you should complete. Neatness counts!

Useful handouts from the Chemistry 222 website (http://mhchem.org/222):

- Geometry and Polarity Guide (http://mhchem.org/geopo/)
- MO Diagram - B_2 through N_2 (http://mhchem.org/MO)
- MO Diagram - O_2 through Ne_2 (http://mhchem.org/MO)



Molecular orbital diagrams for B, C, N (left) and O, F and Ne (right).
Notice that the 1s interactions are ***not*** *included in these diagrams.*
Use [core electrons] *if not showing the 1s interactions in your molecular orbital diagram.*

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VALENCE BOND THEORY and MOLECULAR ORBITAL THEORY LAB - ***Worksheet*** **Name:**

Hand drawn versions only, no computer generated structures

Complete the sections below by providing the appropriate information in the spaces provided. *Neatness counts!*

Part One: Valence Bond Theory *Complete the following sections using Valence Bond Theory.*

Molecule/Ion	Lewis Structure (& Isomers, if any)		
SeI_2		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:***	***Hybridization:*** ***Bond Order:*** ***Resonance Forms? (Y/N)***

Molecule/Ion	Lewis Structure (& Isomers, if any)		
$AsCl_5$		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:***	***Hybridization:*** ***Bond Order:*** ***Resonance Forms? (Y/N)***

Molecule/Ion	Lewis Structure (& Isomers, if any)		
CO_3^{2-}		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:***	***Hybridization:*** ***Bond Order:*** ***Resonance Forms? (Y/N)***

Molecule/Ion	Lewis Structure (& Isomers, if any)		
ClO_3^{-1}		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:***	***Hybridization:*** ***Bond Order:*** ***Resonance Forms? (Y/N)***

Molecule/Ion	Lewis Structure (& Isomers, if any)		
ClO_4^{-1}		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:***	***Hybridization:*** ***Bond Order:*** ***Resonance Forms? (Y/N)***

Molecule/Ion	Lewis Structure (& Isomers, if any)		
$XeOF_4$		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:***	***Hybridization:*** ***Bond Order:*** ***Resonance Forms? (Y/N)***

Part Two: Molecular Orbital Theory *Complete the following sections using Molecular Orbital Theory. Draw a complete Molecular Orbital diagram to answer these questions (include all 1s and 2s interactions, no short hand notation) and provide the missing information.*

Molecule / Ion: **Li_2**

Molecular Orbital Diagram:

Bond Order: ______

Number of sigma bonds: ______

Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic**

Should this molecule exist? *(Circle)* **Yes** *or* **No**

Molecule / Ion: **Be_2**

Molecular Orbital Diagram:

Bond Order: ______

Number of sigma bonds: ______

Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic**

Should this molecule exist? *(Circle)* **Yes** *or* **No**

Molecule / Ion: **B_2**

Molecular Orbital Diagram:

Bond Order: ______

Number of sigma bonds: ______

Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic**

Should this molecule exist? *(Circle)* **Yes** *or* **No**

Molecule / Ion: **N_2**

Molecular Orbital Diagram:

Bond Order: ______

Number of sigma bonds: ______

Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic**

Should this molecule exist? *(Circle)* **Yes** *or* **No**

Molecule / Ion: **F_2**

Molecular Orbital Diagram:

Bond Order: ______

Number of sigma bonds: ______

Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic**

Should this molecule exist? *(Circle)* **Yes** *or* **No**

Molecule / Ion: **Ne_2**

Molecular Orbital Diagram:

Bond Order: ______

Number of sigma bonds: ______

Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic**

Should this molecule exist? *(Circle)* **Yes** *or* **No**

Part Three: Theory Comparison *Complete the following sections using both Valence Bond (VB) Theory and Molecular Orbital (MO) Theory.* ***Shorthand notation for MO Diagrams is recommended****, include 1s and 2s interactions.*

Molecule / Ion: **CN^{-1}** *(Use the MO Diagram for B, C and N on this problem)*

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes** **No**

(MO) Paramagnetic? *(circle)* **Yes** **No**

Molecule / Ion: **NO^{+1}** *(Use the MO Diagram for O, F and Ne on this problem)*

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes No**

(MO) Paramagnetic? *(circle)* **Yes No**

Molecule / Ion: **NO** *(Use the MO Diagram for O, F and Ne on this problem)*

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes No**

(MO) Paramagnetic? *(circle)* **Yes No**

Molecule / Ion: **O_2**

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes** **No**

(MO) Paramagnetic? *(circle)* **Yes** **No**

Molecule / Ion: **OF^{-1}**

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes** **No**

(MO) Paramagnetic? *(circle)* **Yes** **No**

Molecule / Ion: **Ne_2^{+2}**

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes No**

(MO) Paramagnetic? *(circle)* **Yes No**

CH 222 Winter 2025:

"Organic Chemistry" *(in class)* ***Lab - Instructions***

Note: **This is the lab for section 01 and H1 of CH 222 only.**

- *If you are taking section W1 of CH 222*, please use this link:
  http://mhchem.org/r/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-4-13 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, January 27 (section 01) *or* **Wednesday, January 29 (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-4-9 through Ia-4-13 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, February 3 (section 01)** *or* **Wednesday, February 5 (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!

Organic Chemistry

The classification of chemical compounds in to the general areas of organic and inorganic derives from the use of the "mineral, vegetable and animal" designation by the early workers in chemistry. Those compounds derived from living systems were termed **organic** (about 1777) whereas those derived from mineral sources were termed **inorganic**. In modern times, organic compounds are classified as compounds of carbon containing either carbon-carbon or carbon-hydrogen bonds or both. Originally, organic compounds were thought to be imbued with a "vital essence" attainable only from God. Thus, it was believed that organic compounds could be prepared from sources that had once lived, as this would be the only way that this vital essence could be obtained by man. In 1828, Freidrich Wöhler prepared the organic compound urea (found in human urine) from entirely nonliving sources, thereby destroying the theory of organic vitalism. Since Wöhler's time, approximately 5 million organic compounds have been synthesized and characterized, many of which are not found in nature.

Why are there so many organic compounds? The reason is that carbon atoms have the ability to link to other carbon atoms (concatenate) to produce chains or rings of almost infinite size. Other elements do not concatenate nearly as well due to such factors as poor orbital overlap and lone pair-lone pair electronic repulsions. Other elements can also combine with carbon to form hetero-species, including hydrogen, oxygen, nitrogen, sulfur and the halogens.

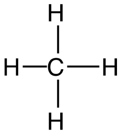
The distinction between the organic and inorganic disciplines is not very sharp. The bonding of metals to carbon has resulted in the large, important and fast growing area of **organometallic** chemistry. Organometallic compounds containing metals and metalloids such as lithium, magnesium, copper, iron, boron, silicon, and other elements play major roles as synthetic reagents.

The purpose of this experiment is to prepare models of the more common organic compound types to enhance their three dimensional nature. Bond angles between atom groupings will become apparent as you build the models.

Saturated Hydrocarbons: The Alkanes

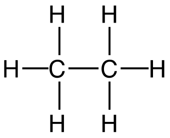
Alkanes are **hydrocarbons** (compounds with hydrogen and carbon atoms) linked with single bonds. All the carbon atoms are sp^3 hybridized and are tetrahedrally bonded to four other carbon or hydrogen atoms. Members of this class have the general formula $\mathbf{C_nH_{2n+2}}$, where **n** is an integer. Examples of this class of organic compound include methane (CH_4, n=1), ethane (C_2H_6, n=2) and propane (C_3H_8, n=3). Structures of the first four straight chain compounds are shown below.

methane



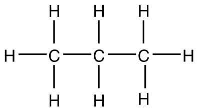
CH_4

ethane



CH_3CH_3

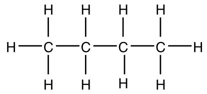
propane



$H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$

$CH_3CH_2CH_3$

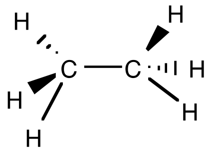
butane



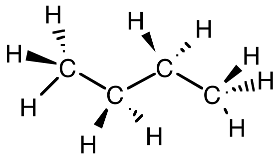
$H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$

$CH_3CH_2CH_2CH_3$

These representations do not show the actual geometrical structure of these compounds. Due to the sp^3 hybridization of the carbon atoms, all the bond angles in the molecules are close to 109.5 ° (tetrahedral) and thus, the carbon chain is nonlinear. The **wedge-dash** notation can be used to represent the nonlinear nature of these molecules. Examples of ethane and butane are given below using the wedge-dash notation.



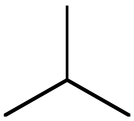
ethane



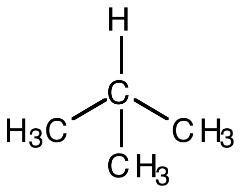
butane

As you prepare models of the alkanes, note that each carbon atom can rotate about its respective carbon-carbon bond. Some conformations (arrangements of the bonds and groups relative to each other) are more stable than others, since in these arrangements there is less interference (steric repulsion) between the hydrogen atoms attached to nearby carbons.

An interesting aspect of the alkanes is that **structural isomers** can exist in compounds having more than three carbon atoms. Structural isomers are species that have the same molecular formula but possess different physical properties due to different arrangements of the carbon backbone. The two compounds below are an example of an isomeric pair of hydrocarbons. Both are butanes and have the formula C_4H_{10}, but their structures are different.



$CH_3CH_2CH_2CH_3$

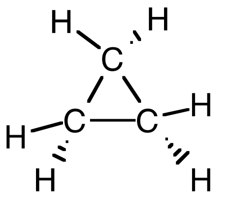


n-butane

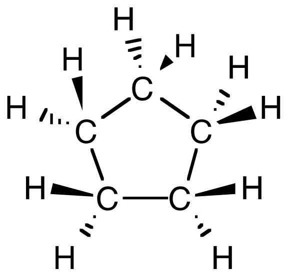
methylpropane (isobutane)

This figure relates several new points. First is the use of the **skeleton formula**. These structures represent the carbon backbone without the hydrogen atoms being shown and are often used as a shorthand method of representing the structures. The second point is the use of substituents in naming compounds. We can see that in methylpropane, one of the CH_3 groups is a "twig" off of the main "branch" of the compounds. Such twigs are called **alkyl groups**. In the figure, the **methyl group** (CH_3) name is derived from the hydrocarbon methane (CH_4) having lost a hydrogen. The number of possible isomers increases rapidly as the number of carbon atoms increases in a compound. For example, the pentane system (five carbons) has three isomers, the heptane system (seven carbons) has nine isomers, and the decane system (ten carbons) has seventy-five isomers!

Alkanes can also exist as **cyclic** hydrocarbons where the carbon atoms are arranged in rings. The general molecular formula is $\mathbf{C_nH_{2n}}$, where **n** is an integer. These carbons are named identically to the alkanes except for the additional prefix *cyclo*. The structures for two cyclic hydrocarbons are given below.



cyclopropane



cyclopentane

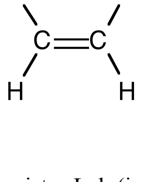
In the smaller rings such as cyclopropane and cyclobutane, smaller bond angles (60° and 90°, respectively) are evident. Such angles are seen as being strained from their normal tetrahedral angle, and these compounds have less stability than their larger counterparts. The most stable cyclic compounds contain rings of five and six carbons (cyclopentane and cyclohexane, respectively).

Unsaturated Hydrocarbons: The Alkenes

Alkenes are hydrocarbons in which there are one or more carbon-carbon double bonds, C=C. The carbon atoms attached to the double bond are sp^2 hybridized. One of the double bonds is a **sigma** bond (oriented along the internuclear axis) and the other is a **pi** bond (oriented perpendicular to the internuclear axis). Members of this class have the general formula $\mathbf{C_nH_{2n}}$, where **n** is an integer (note that this is the same formula as for the cycloalkanes.) Examples of alkenes include ethene (C_2H_4, n=2, also known as ethylene), propene (C_3H_6, n=3, also known as propylene) and butene (C_4H_8, n=4). Structures of several of the alkenes are shown below. Alkenes are named in the same fashion as the alkanes except that the ending *-ene* replaces the ending *-ane*. In cases where more than one structural isomer exists (such as butane, below), it is necessary to indicate the

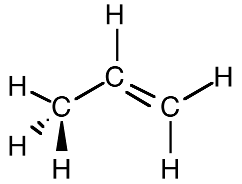
location of the double bond by numbering the carbon atoms in the longest chain containing the double bond, and then giving it the lowest possible number.

ethene (ethylene)



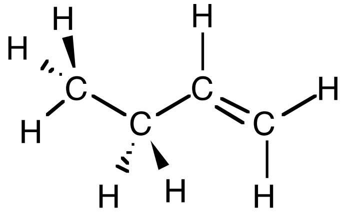
$CH_2=CH_2$

propene (propylene)



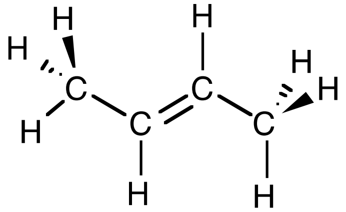
$CH_3CH{=}CH_2$

1-butene
or but-1-ene



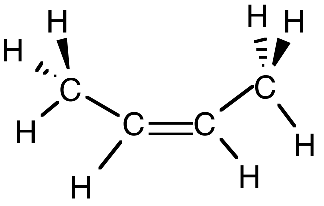
$CH_3CH_2CH{=}CH_2$

trans-2-butene
or trans-but-2-ene



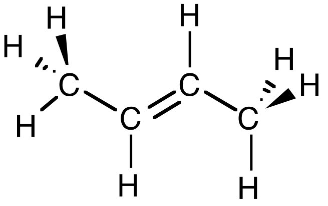
$CH_3CH{=}CHCH_3$

Many kinds of isomerism exist in organic compounds. Since carbon atoms are free to rotate around the C-C single bonds but <u>not</u> around a C=C double bond (this disrupts the overlap of the pi part of the double bond), it is possible to have two separate geometrical isomers of 2-butene. These are known as the *cis*- and *trans*- isomers and are shown below:



cis-2-butene

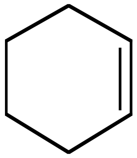
cis-but-2-ene



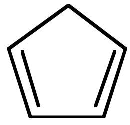
trans-2-butene

trans-but-2-ene

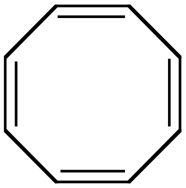
Cyclic structures containing C=C bonds are also possible, and three examples follow.



cyclohexene



1,3-cyclopentadiene



1,3,5,7-cyclooctatetraene

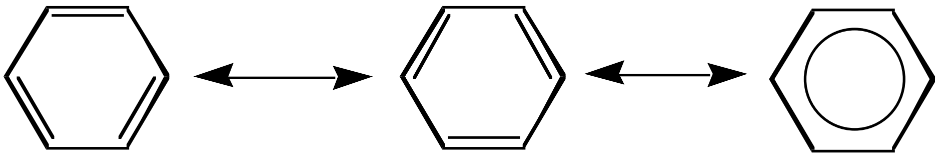
Unsaturated Hydrocarbons: The Alkynes

Hydrocarbons containing a carbon-carbon triple bond are named **alkynes**. The two carbon atoms forming the triple bond are joined by one sigma and two pi bonds and are sp hybridized. The alkynes have the general molecular formula $\mathbf{C_nH_{2n-2}}$. Several representatives of this class are shown below. Alkynes are names in the same manner as alkenes except that the ending *-yne* replaces the ending *-ene*.

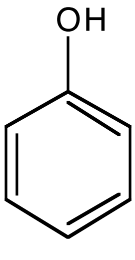
$H\text{-}C{\equiv}C\text{-}H$	$CH_3\text{-}C{\equiv}C\text{-}H$	$CH_3\text{-}C{\equiv}C\text{-}CH_3$
ethyne (acetylene)	propyne	2-butyne

Aromatic Hydrocarbons

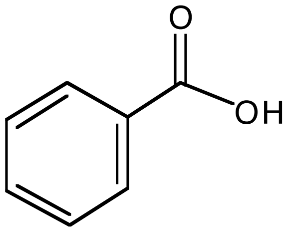
Benzene, C_6H_6, is the most crucial member of this important class of hydrocarbons. The molecule contains a ring of six sp^2 hybridized carbon atoms with the unhybridized p orbitals perpendicular to the ring system. The six electrons in the pi parts of the bonds are **delocalized** (spread evenly over the six carbon nuclei). Such molecules are often unusually stable. There are two equally valid ways of representing the structure of benzene. These two ways are called resonance forms, and the molecule is a **resonance hybrid** with the "true structure" of benzene lying midway between the two resonance forms. As an example, a mule is a genetic hybrid descendant of a male donkey and a female horse. The mule does not change back and forth, being a donkey half the time and a horse the other half. Thus, the properties of a resonance hybrid (such as benzene) are fixed.



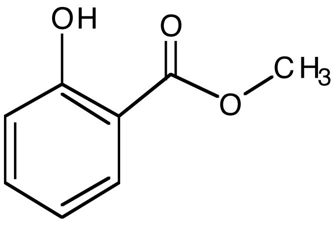
There are many familiar hydrocarbon species which are based on the benzene structure that contain various functional groups. A few representative examples are given below. Note that the delocalized pi electrons can be represented by a circle in the ring. This is an alternate representation of the ring often used by chemists.



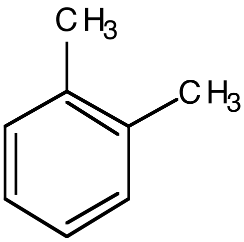
phenol



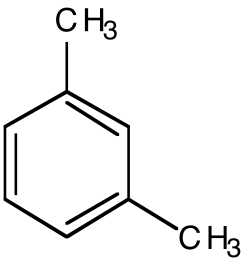
benzoic acid



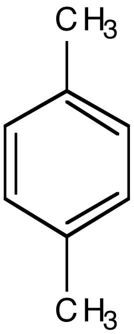
oil of wintergreen



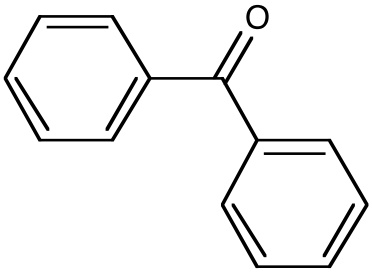
1,2-dimethylbenzene
(ortho-xylene)



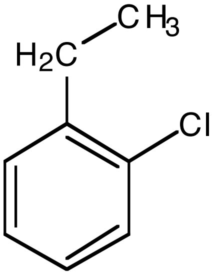
1,3-dimethylbenzene
(meta-xylene)



1,4-dimethylbenzene
(para-xylene)

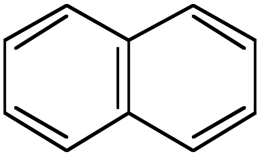


benzophenone

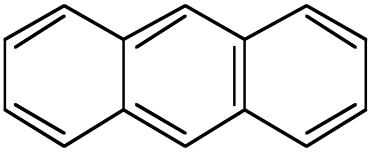


2-chloroethylbenzene

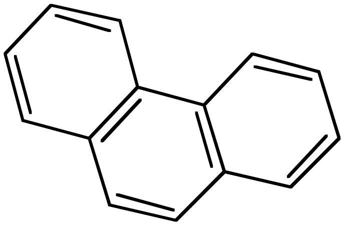
There are also many important aromatic hydrocarbons that contain more than one fused benzene ring. Several are illustrated here.



naphthalene



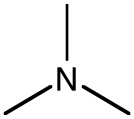
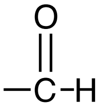
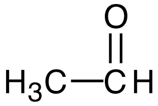
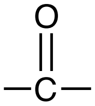
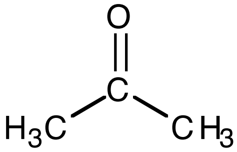
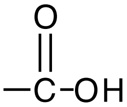
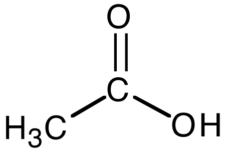
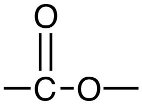
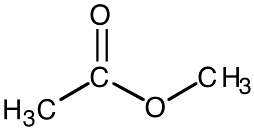
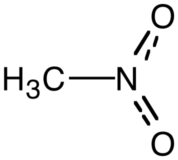
anthracene



phenanthrene

Hydrocarbons Containing Functional Groups

The basic types of hydrocarbon compounds outlined above may have one or more of their hydrogen atoms replaced by a **functional group**. The substituted benzenes earlier illustrated a number of functionalities ($-CH_3$, $-OH$, etc.) attached to the aromatic ring. Additional examples are shown in the table below.

Functional Group	Class of Compound	Example	Name
$-OH$	alcohol	H_3C-CH_2-OH	ethanol (ethyl alcohol)
$-O-$	ether	$H_3C-O-CH_3$	dimethyl ether
	amine	H_2N-CH_3	methylamine
 $-\overset{\overset{O}{\|}}{C}-H$	aldehyde	 $H_3C-\overset{\overset{O}{\|}}{C}H$	ethanal (acetaldehyde)
 $-\overset{\overset{O}{\|}}{C}-$	ketone		propanone (acetone)
 $-\overset{\overset{O}{\|}}{C}-OH$	carboxylic acid		ethanoic acid (acetic acid)
 $-\overset{\overset{O}{\|}}{C}-O-$	ester		methyl acetate
$-NO_2$	nitro		nitromethane
-X (X = F, Cl, Br, I)	haloalkane	H_3C-CH_2-Cl	chloroethane (ethyl chloride)

Name: ____

Organic Chemistry Worksheet - *Hand drawn versions only, no computer generated structures*

Answers to this lab should be provided in this worksheet. Neatness and attention to detail will count.

---

1. **Alkanes and Cycloalkanes**
   a. Write the structural formula and name for each straight-chain alkane with between one and eight carbons.

   b. What are the geometries, bond angles and hybridizations of the carbon atoms in the alkanes? What pattern do you notice for the relationship of carbon to hydrogen? (i.e. if an alkane has *n* carbons, how many hydrogen atoms are on the molecule?)

   c. Draw and name the five structural isomers of the hexane molecule. Give a suitable name to each of the hexane isomers. For example, one of the isomers is named 3-methylpentane. (Remember that one of the isomers is *n*-hexane!)

   d. Draw the structure for cyclohexane. What is the molecular formula of cyclohexane? What is the relationship between the number of carbon atoms and hydrogen atoms? Is cyclohexane an isomer of hexane? Explain.

2. **Alkenes and Alkynes**
   a. Draw the structure and molecular formula for ethene. What is the geometry and bond angle around the double bond? What is the hybridization of the carbon atoms? What is the relationship between the number of carbon atoms and hydrogen atoms?

   b. Construct four alkene isomers with four carbons each. Draw and name the structural formulas. Be sure to include *cis* and *trans* labels when appropriate.

   c. Draw the structure and molecular formula for ethyne. What is the geometry and bond angle around the triple bond? What is the hybridization of the carbon atoms? What is the relationship between the number of carbon atoms and hydrogen atoms? How many sigma and pi bonds are in ethyne?

   d. Construct models of the three five carbon alkyne isomers. Draw and name the isomers.

   e. Draw cyclobutyne. Using your knowledge of bond angles and hybridization, would you expect cyclobutyne to be a stable compound? Explain.

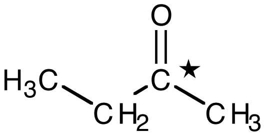
3. **Aromatic Compounds**
   a. Sketch the line structure of the aromatic compound benzene, C_6H_6. Comment on the fact that measurement of the C-C bond distances in benzene shows that all are the same length (1.397 Angstroms) and that the C-C bond angles are all 120°. Is benzene a flat molecule?

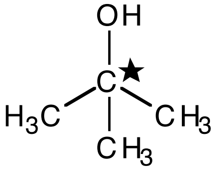
   b. Benzene was at one time used extensively as a solvent. However, since it has been found to be carcinogenic, it has largely been replaced by toluene. Toluene, which is formally known as methylbenzene, contains an aromatic ring substituted with a methyl group and has a molecular formula of C_7H_8. Sketch a possible structure for toluene. Do you think toluene would have a higher or lower boiling point than benzene? Explain.

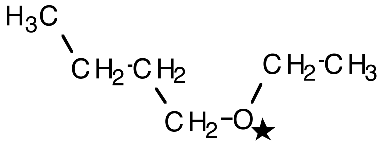
   c. Draw two resonance structures for naphthalene. (*Hint*: look in this lab for the naphthalene structure!)

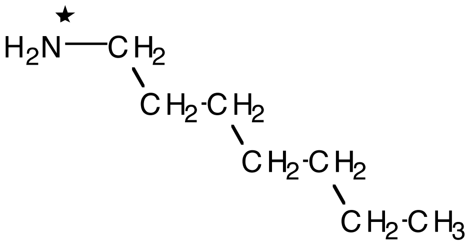
   d. There are three dichlorobenzene isomers. Draw and name them. Which one would you expect not to have a dipole moment? Explain.

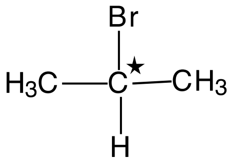
4. Identify the functional group and/or class of compound in each of the following molecules. List the hybridization around the marked atom in each compound. Finally, name the compound.

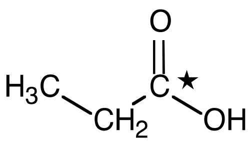


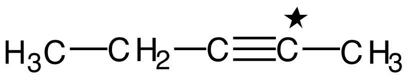












5. Draw structures for acetone, formaldehyde and acetic acid. What are the bond angles around the C=O carbon atom in each of these compounds?

6. Draw the structures for the following compounds:

   a. 2,3-difluorohexane

   b. dimethyl ether

   c. butanone

   d. dimethylamine

   e. pentan-2-ol

   f. 2,3-diethyltoluene

   g. *cis,trans*-octa-2,6-diene

   h. 2-hexyne

   i. 3-nitrotoluene

   j. 4-bromobenzoic acid

   k. 2,4,6-trinitrotoluene

   l. 2-chlorophenol

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CH 222 Winter 2025:
“Molar Mass of a Volatile Liquid *(in class)*” *Lab* - *Instructions*

Note: **This is the lab for section 01 and H1 of CH 222 only.**

- *If you are taking section W1 of CH 222*, please use this link: http://mhchem.org/r/5b.htm

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Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-5-2 through Ia-5-9 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, February 3 (section 01) *or* **Wednesday, February 5 (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. *Bring your safety glasses to this lab!*

Step Three:

Complete the lab work and calculations on your own, then **turn it in** (pages Ia-5-5 through Ia-5-9 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, February 10 (section 01)** *or* **Wednesday, February 12 (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!

Molar Mass of a Volatile Liquid

The experimental determination of the molar mass of gases and vapors is one important application of the Ideal Gas Law. To determine the molar mass of a gas or vapor, we need to determine the mass (g, below) of the gas sample under known conditions of temperature (**T**), Volume (**V**) and pressure (**P**). Assuming the gas obeys the Ideal Gas Law:

$$PV = nRT$$

and if pressure, temperature and volume are determined experimentally and **R** is a constant (**0.082057 L atm mol^{-1} K^{-1}**), we can solve the equation for moles (**n**) of gas. The molar mass (grams/mole, or **MM**) of the gas or vapor can be determined based on the mass of the liquid remaining (**g**) in the flask and the number of moles calculated from the ideal gas law; or you can use:

$$MM = gRT/PV$$

In this experiment, the molar mass of a volatile liquid will be determined using the ideal gas law. A small amount of the liquid is introduced into a weighed flask. The flask is then placed in boiling water where the liquid will completely vaporize, driving out any air and filling the flask with vapor at barometric pressure and the temperature of the boiling water. The mass of the vapor can be determined by cooling the flask to condense the vapor. (Note: ***the thermometer does not need a separate clamp*** (as shown in Figure One, below); you can place the thermometer directly in the boiling water when a reading is required.)

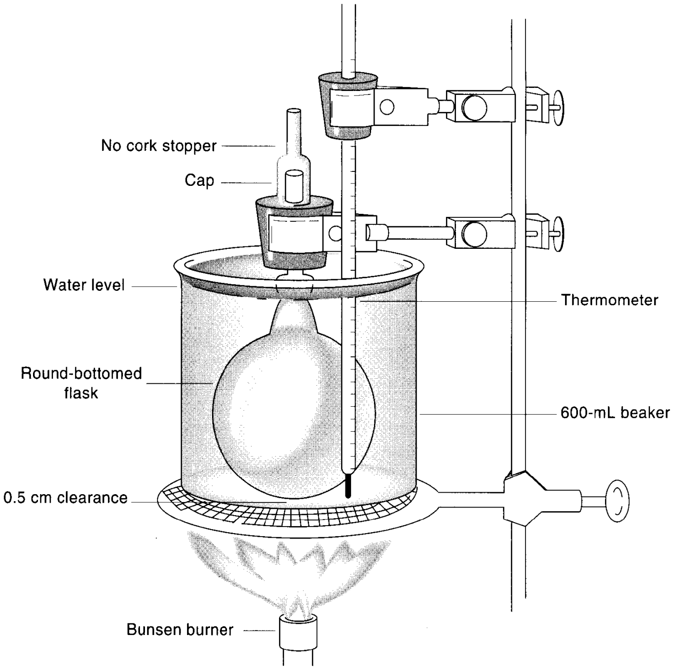

Figure One

PROCEDURE:

Obtain a round bottom flask fitted with a glass cap with a capillary opening, and a cork ring should support the entire apparatus at all times. If necessary, wash the flask with a touch of acetone, then dry the flask and cap and record the mass of the empty, dry flask and cap to the nearest 0.001g using the cork ring for support.

Obtain an unknown liquid and record the identification number. Pour approximately 5 mL of your unknown sample into the flask. (Note: the exact volume is not important; the mass of the final condensed vapor will be determined.) Assemble the apparatus as shown in Figure 1. It is important to have the round bottom flask immersed as deeply as possible in the 1000 mL water bath for uniform heating. Optionally, add 2-3 boiling chips to the water in the 1000 mL beaker to avoid 'hard' bubbles. Heat the water to the boiling point. Watch the liquid level in your flask; the level should gradually disappear as vapor expands and pushes first air from the flask, then as excess unknown vapor. Once all the liquid has disappeared, continue to heat the flask for 5 minutes to allow the vapor to reach the temperature of the boiling water. Measure and record the temperature of the boiling water. Also record the barometric pressure.

Turn off the burner and wait for the water to stop boiling (about 30 seconds). Remove the flask from the beaker of water and allow the flask to cool to room temperature and the vapor to condense. (You can speed up this process by running tap water over the stoppered flask.) Dry the outside of the flask and weigh the flask with the condensed vapor.

Repeat the procedure using another 5 mL sample of your liquid unknown. **Do not empty or clean the flask between trials**; just pour another 5 mL into the flask

Once the two trials are complete (**and not before**!), find the volume of the flask. Fill the flask and cap with water. Determine the mass of the flask filled with water, then determine the temperature of the water used to fill the flask. Using the *Handbook of Chemistry and Physics*, determine the density of the water to at least six sig figs. If the *Handbook* is difficult to read, use this link as an alternative: http://mhchem.org/den

Rinse the flask with a touch of acetone, then blow dry. Return the flask to the cart.

Using the density of water and the mass of water in the flask, calculate the volume of the flask. Be sure to subtract the empty mass of the flask first! When converting to atmospheres, use **1013 mbar = 1 atm** (other conversions will not receive full credit), use **273.15** to convert to Kelvin temperatures, and **use the complete version of R** (listed on the previous page.)

Determine the **average molar mass** of your liquid sample and **parts per thousand**

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Molar Mass of a Volatile Liquid

Record all masses to the milligram (0.001 g)

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

DATA:

empty flask (g): _____________________ *Include flask, stopper and cork ring*

Unknown number: _____________________

Trial I	**Trial II**
Boiling water temperature (°C): _____________________	Boiling water temperature (°C): _____________________
flask + condensed vapor (g): _____________________	flask + condensed vapor (g): _____________________
barometric pressure (mbar): _____________________	barometric pressure (mbar): _____________________

flask filled with water (g): _____________________

Density of water (g/mL): _____________________ *from Handbook of Chemistry & Physics* or link

Density temperature (°C): _____________________ *"cold" water temperature*

Part A Calculations: *Molar Mass Determination of a Volatile Liquid*

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

1. Using your data, determine the temperature of the gas in Kelvin for both Trials.

Trial I Temperature (K): ______________

Trial II Temperature (K): ______________

2. Using your data, determine the pressure of the gas in atmospheres for both Trials.

Trial I Pressure (atm): ______________

Trial II Pressure (atm): ______________

3. Using your data, determine the mass of the unknown liquid remaining at the end of the experiment for both Trials.

Trial I mass of liquid (g): ______________

Trial II mass of liquid (g): ______________

4. Using your data, calculate the volume (L) that the gas occupies in the flask (*hint:* use only the **mass of water** in the flask; this will be the same for both Trial I and Trial II.)

Volume of flask (L): ______________

5. What is the value of R for this lab? Report to five significant figures and include units.

Value of R with units: _______________________________

6. Using the previously recorded and calculated values, **determine the molar mass** of the gas **for Trial I and Trial II**. **Calculate** the **average molar mass** for both trials. **Determine** the **average deviation** and the **parts per thousand** using your molar mass values. *Hint:* do not average any values until you have calculated the molar mass values for the unknown liquid. *Show all work!*

molar mass (Trial I) (g/mol): ______________

molar mass (Trial II) (g/mol): ______________

average molar mass (g/mol): ______________

average deviation: ______________

Parts per thousand: ________________

Postlab Questions:

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

1. Use the data below to find the moles of unknown and the molar mass of the unknown. *Show all work!*
    - mass of an empty flask and stopper = 55.441 g
    - 5 mL of unknown added and heated; boiling water bath at 100.˚C, all unknown liquid vaporized
    - mass of the flask, stopper and condensed vapor = 56.039 g.
    - volume of the flask = 215.9 mL
    - barometric pressure = 1003 mbar

moles unknown: _____________ **Molar mass of unknown (g/mol)**: __________

2. Determine if each of the following procedural errors would **increase**, **decrease**, or have **no effect** on the molar mass calculations in this experiment. **Explain** your reasoning.

   i. The flask was not dried before the final weighing with the condensed vapor inside.

      Effect on molar mass:

      Explain:

   ii. The flask was removed from the boiling water before the vapor had reached the temperature of the boiling water. All the liquid had vaporized.

      Effect on molar mass:

      Explain:

   iii. The volume of the cap was not measured with the volume of the flask.

      Effect on molar mass:

      Explain:

   iv. The student uses 7.5 mL of liquid sample.

      Effect on molar mass:

      Explain:

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CH 222 Winter 2025: **“Linear Regression & The Crystal Structures of Solids** *(in class)”* ***Lab*** *- Instructions*

Note: **This is the lab for section 01 and H1 of CH 222 only.**

- *If you are taking section W1 of CH 222*, please use this link:
  http://mhchem.org/r/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-6-16 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, February 10 (section 01) *or* **Wednesday, February 12 (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-6-11 through Ia-6-16 *only* <u>with computer generated graphs</u> to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, February 17 (section 01)** *or* **Wednesday, February 19 (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!

Linear Regression & The Crystal Structures of Solids

This lab will demonstrate the power of a linear regression analysis while graphing linear data. We will also explore the techniques used to analyze several common types of metals in crystallography.

An **equation** is a mathematical model used to describe the relationship between variables. We will focus on **linear** equations in this lab which use a horizontal (X) axis (the **independent** variable, the variable we input when we make a measurement) and a vertical (Y) axis (the **dependent** variable, the number we measure after we set the X value.) If the plotted data points form a straight line, this means we have a **linear equation**, and we can use: $\mathbf{y = mx + b}$ where **y** is the vertical axis value, **x** is the horizontal axis value, **m** is the slope of the line, and **b** is the y-intercept.

Computer programs and calculators can perform a **linear regression** analysis by plotting the "best fit" line through the data and then writing the slope-intercept equation. The **correlation coefficient** (with the symbol "**r**") is a measure of how well the regression line fits with the observed data. A ***perfect*** fit produces a correlation coefficient of either +1.000 (positive slope) or -1.000 (negative slope), depending on if the line slopes up (a positive slope) or down (a negative slope.) The closer the correlation coefficient is to +/-1.000, the better the regression line expresses the data (the better fit for the data.) Note that R^2 is slightly different than r!

Observing the **crystals** of an ordinary substance (such as table salt) using a magnifying glass, one sees many planes at right angles within the solid. This occurs in many common solids, and the regularity we see implies a deeper regularity in the arrangement of atoms or ions in the solid. The atomic nuclei are present in remarkably symmetrical arrays that continue for millions of units in three dimensions. Substances having a regular arrangement of atom-size particles in the solid are called **crystalline**, and the solid material consists of **crystals**. This lab deals with some of the simpler arrays in which atoms or ions occur in crystals and what these arrays can tell us about such properties as atomic sizes, densities of solids, and the efficiency of packing of particles.

Procedural Notes for the Linear Regression Lab: *Complete the problems using the worksheets at the end*

Each linear regression problem will require one or more **computer generated graphs** that will be stapled to the back of the worksheets. Recommended programs to graph your date include **Microsoft Excel** (**free** for MHCC students; see https://www.mhcc.edu/OfficeInstall/), **Apple Numbers** (free with a Mac computer, https://www.apple.com/numbers/), or **Google Sheets** (https://www.google.com/sheets). ***Note*** that Excel Online (the online version of Excel) and iPad/iPhone/Droid versions of these programs will generally not perform linear regressions, so try to use the "full" computer version instead. Calculators will perform linear regressions, but printing from a calculator might be difficult.

Use a **X-Y scatter plot** when graphing these data sets. **The computer program will analyze the data and perform the linear regression analysis** for you. Each program is different, but generally the user selects the actual data points on the X-Y Scatter plot and either right-clicks or control-clicks on the data to see a new menu…. you wish to "Add a Trendline" and "Display the R^2 value". If an equation appears with an R^2 value, you have performed your linear regression. ***Help*** on performing the linear regression lab can be acquired in the Learning Success Center / AVID Center at MHCC. You can also search YouTube for videos (i.e. search "linear regression Excel 2019" and almost inevitably a helpful video appears.)

Converting R^2 to r is not difficult. Take the square root of R^2 to get r. If the value of the slope is a negative number, then the value of r will also be negative…. watch for this in this lab! Calculators will often give both R^2 and r values when linear regression techniques are applied. Information on making an acceptable graph in this class can be found here: https://mhchem.org/lab

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The Linear Regression Problems:

<u>Problem 1</u>: The Relationship Between Celsius and Fahrenheit

In 1724, the German scientist Gabriel Fahrenheit developed a temperature scale based on phenomenon he thought could be easily repeated in laboratories around the world. For his zero degree point, Fahrenheit chose the coldest mixture of ice, water, and salt that he could produce in his laboratory. For ninety-six degrees, he chose what he believed to be normal body temperature. Fahrenheit wanted a temperature scale that could be divided into twelfths. On this scale, pure water freezes at 32 degrees, and pure water boils at sea level at 212 degrees.

A few years later, in 1742, the Swedish scientist Anders Celsius developed a different temperature scale. This scale used pure water as its standard. Zero degrees was the temperature where pure water froze, and one hundred degrees was the temperature where pure water boiled at sea level. Because Celsius had one hundred degrees between the two reference points on his temperature scale, it was called the *centigrade* scale. Recently this was renamed the Celsius scale in honor of Anders Celsius.

A student measures the following data points in the laboratory using two thermometers:

Temperature (°C)	20.0	40.0	60.0	80.0	100.0
Temperature (°F)	67.6	104.8	141.1	175.0	211.1

1. Construct and print a graph of degrees Fahrenheit (y) as a function of temperature in degrees Celsius (x).
2. Using your calculator, determine the mathematical equation of °F as a function of °C as well as the correlation coefficient, r. Record r to at least four significant figures.
3. Using the actual equation: °F = 1.8°C + 32 *and* your experimental equation, convert 29.0 °C to °F. Calculate **percent error** = (difference / actual value) x 100% Comment on discrepancies.

<u>Problem 2</u>: Solubility of Lead(II) Nitrate in Water

The solubility of lead(II) nitrate in water was measured as a function of temperature. The solubility is given in units of grams of lead(II) nitrate per 100 grams of water.

Temperature (°C)	20.0	40.0	60.0	80.0	100.0
Solubility (g / 100 g water)	56.9	74.5	93.4	114.1	131.1

1. Graph and print the data; temperature will be the independent (x) variable.
2. Determine the equation of the best-fit line. Record the equation and correlation coefficient.
3. What is the solubility of lead(II) nitrate at 47.0 °C?

Problem 3: Colorimetry

The colors in the visible spectrum of light are shown by a rainbow. Colored substances absorb segments of the visible spectrum of light. Pink solutions, for example, are pink because they absorb green light and transmit all other colors of the visible spectrum. If light of the particular color absorbed is passed through a sample, the amount of light absorbed will be related to the number of absorbing molecules in the light beam. Dilute solutions absorb little light, concentrated solutions absorb more. Typically the amount of light transmitted through the solution is measured; *transmittance* is inversely proportional to *absorbance*. The following data was obtained for the transmittance of 525 nm light by solutions containing different concentrations of permanganate ion.

Concentration (mg/100 mL)	1.00	2.00	3.00	4.00
Transmittance *(unitless)*	0.418	0.149	0.058	0.0260

1. **Convert the Transmittance values to Absorbance** using the following equation: **A = log (1/T)**, where A = Absorbance and T = Transmittance. **Use 3 sig figs for your absorbance values**.
2. Graph and print the Absorbance (y) versus Concentration (x) data. Perform a linear regression analysis. Record the equation and the correlation coefficient.
3. Predict the absorbance of 2.50 mg permanganate ion / 100 mL solution.

Problem 4: Kinetics

The branch of chemistry that studies the rate or speed of reactions is called *kinetics*. One must often plot concentration versus time data in a variety of mathematical formats to find a linear relationship; this assists in finding the *order of reaction*. We shall explore this topic more in CH 222. The following data was collected at 25.6 °C while measuring the disappearance of NH_3:

Concentration [NH_3] (mol/L)	$8.00 * 10^{-7}$	$6.75 * 10^{-7}$	$5.84 * 10^{-7}$	$5.15 * 10^{-7}$
Time (h)	0	25.0	50.0	75.0

1. **Prepare a graph of ln [NH_3] versus time** (time is the x-axis). "ln" stands for natural logarithm which can be calculated easily on your calculator (for example, the value of $8.00 * 10^{-7}$ is -14.039.) Perform a linear regression analysis on the ln [NH_3] versus time data and find the equation and the correlation coefficient. **Use at least 4 sig figs for your absorbance values**.
2. **Prepare a graph of 1 / [NH_3] versus time** (for example, $1 / 8.00 * 10^{-7}$ is $1.25 * 10^{6}$. ***Note:*** You may have to enter the data as "1.25E6" to make the program understand your values.) Time will be the x-axis. Perform a linear regression analysis and find the correlation coefficient and the values for the slope and the y-intercept.
3. **Which graph gives a *better* linear regression?** Why? *Hint:* look for the better correlation coefficient.
4. Plots of ln [NH_3] versus time that are linear are called *first order reactions* while graphs of 1 / [NH_3] versus time that are linear are called *second order reactions*. **What order of reaction** does the decomposition of NH_3 follow? *Hint:* the better linear regression will determine the order of the reaction!

Procedural Notes for the Crystal Structures of Solids Lab: Complete the handouts found at the end of this lab and turn it in (with all relevant work displayed on adjacent pages) to your instructor. Help on completing the "Solids" questions can be acquired in the Learning Success Center / AVID Center at MHCC.

Many crystals are unbelievably complex, and we will limit ourselves to the simplest crystals that have cubic structures. Cubic structures imply 90° angles and sides of equal length (hence, a cube.) We will also limit ourselves to the study of only one kind of system (namely metal elements), yet they will exhibit many of the interesting properties of more complicated structures.

The Simple Cubic (SC) Crystal

The simple cubic unit cell is a cube with an **edge length**, $\mathbf{d_0}$, equal to the distance from the center of one atom to the center of the next (see Figure One). The volume of the cube is equal to $(d_0)^3$, expressed as

$$\mathbf{V = (d_0)^3}$$

and is very small since d0 is on the order of 0.5 nm. Using x-ray diffraction we can measure the value of d0 easily to four significant figures. The number of atoms in a simple cubic unit cell is equal to one, for only 1/8 of each corner atom is actually inside the cell.

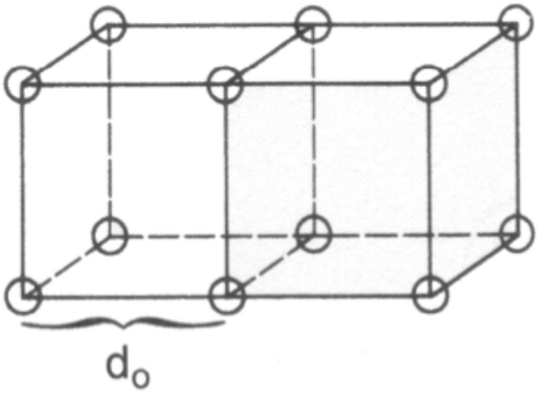


Figure One: *The Simple Cubic Crystal*

Each atom in the simple cubic unit cell is actually connected to six other atoms in the cubic lattice; hence, we say that the **coordination number** of the atoms in this structure is equal to six.

Many diagrams displaying the simple cubic unit cell show a gap between adjoining atoms. **In an actual crystal, we consider that the atoms that are closest are touching**. It is on this assumption that we determine **atomic radii, r**. In the SC crystal, if we know d_0, we can find the radius r of the atoms, since one side contains 2 atomic radii, or

$$\mathbf{d_0 = 2r}$$

for simple cubic crystals. Knowing the radius, we can calculate d0, and then we can calculate the volume of the unit cell. Knowing that one atom occupies the simple cubic cell, we can calculate the mass of the unit cell (using the molar mass and Avogadro's number), and from this we can determine the density using the volume of the cell.

Essentially no elements crystallize in the simple cubic structure, however, due to the inefficiency of the packing. The atoms in the simple cubic crystal are farther apart then they need to be, and inspection of the SC lattice will reveal a large hole in the center of the unit cell. Only about 52% of the cell volume is occupied by atoms, and more "empty space" means less stabilization for the crystal structure.

The Body Centered Cubic (BCC) Crystal

In a **body centered cubic crystal**, the unit cell still contains the corner atoms present in the SC structure, but the center of the cell now contains an additional atom. This means that every BCC crystal structure holds **two net atoms** (eight atoms are $^1/_8$ within the cell, and one whole atom within the center of the cell for two net atoms).

The edge length, d_0, can be determined using simple geometry from the **cube diagonal** (see Figure Two). The cube diagonal reaches across the cube, from an atom in the lower left front to an atom in the upper right back, or from any other appropriate combination. Geometry dictates the following relationship between the cube body diagonal and the edge length, d_0:

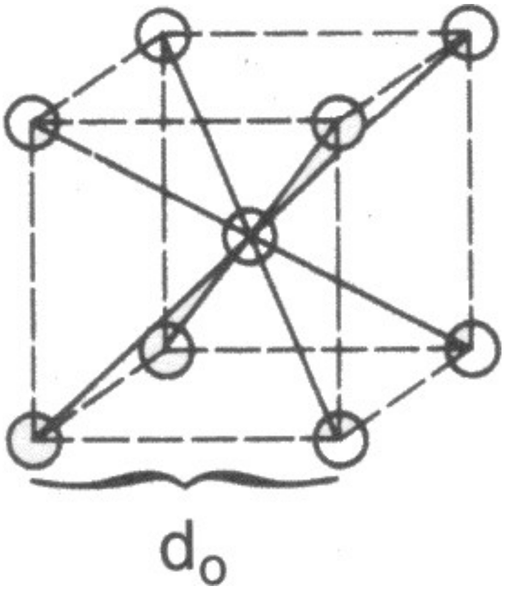


Figure Two: *Body Centered Cubic Crystal*

$$\text{cube diagonal} = \sqrt{3} \cdot d_0$$

The cube diagonal encompasses 4 radii lengths, and d_0 can be expressed in terms of the radius of the atom:

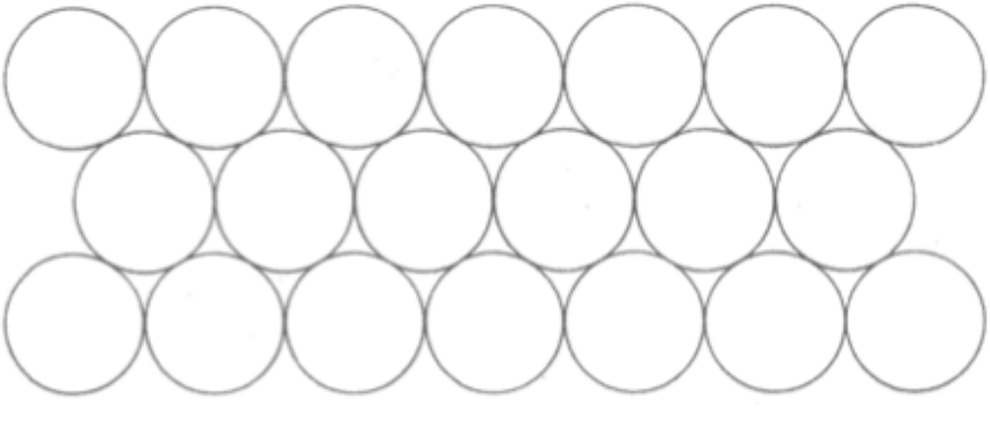
$$d_0 = \frac{4\text{r}}{\sqrt{3}}$$

The quantity d_0 can be used to find the volume of the cube; this is important for BCC cubic systems.

In a BCC lattice, each atom touches eight other atoms, and the **coordination number** is eight. The BCC lattice is much more stable than the SC structure, in part due to the higher coordination number. Many metals at room temperature display the BCC lattice, including sodium, chromium, tungsten and iron. Note that there are two atoms per unit cell in the BCC crystal. **BCC crystals are more efficient than SC crystals**, occupying approximately 68% of the total available volume.

Close Packed Structures

Although many elements prefer the BCC crystal arrangement, still more prefer structures in which the atoms are **close packed**. In close packed structures there are layers of atoms in which each atom is in contact with six others, as in the sketch below:



This is the way in which billiard balls lie in a rack or the honeycomb cells are arranged in a bees' nest. It is the most efficient way one can pack spheres, with about 74% of the volume in a close packed structure filled with atoms.

There is more than one way whereby close packed crystal structures can be stacked. One of the stacking methods is cubic and is called the **Face Centered Cubic (FCC)**. The other is called **Hexagonal Close-Packing**. We shall look at both close packed structures.

The Face Centered Cubic (FCC) Crystal

In the face centered cubic crystal unit cell there are atoms in each corner of the cell (as in the SC cell discussed earlier) and there is another atom at the center of each of the six faces. This means that FCC cubic systems consist of **four net atoms** per unit cell (eight atoms are $1/8$ within the cell, and six faces hold an atom which is $1/2$ within the cell for four net atoms). See Figure Three.

The edge length d_0 can be determined in an FCC crystal from the **face diagonal** which is defined as the distance across one face of the cube. Using geometry, we can find the edge length from the face diagonal using the following equation:

$$\text{face diagonal} = \sqrt{2} \cdot d_0$$

The face diagonal encompasses 4 radii lengths, and d_0 can be expressed in terms of the radius r:

$$d_0 = \frac{4r}{\sqrt{2}}$$

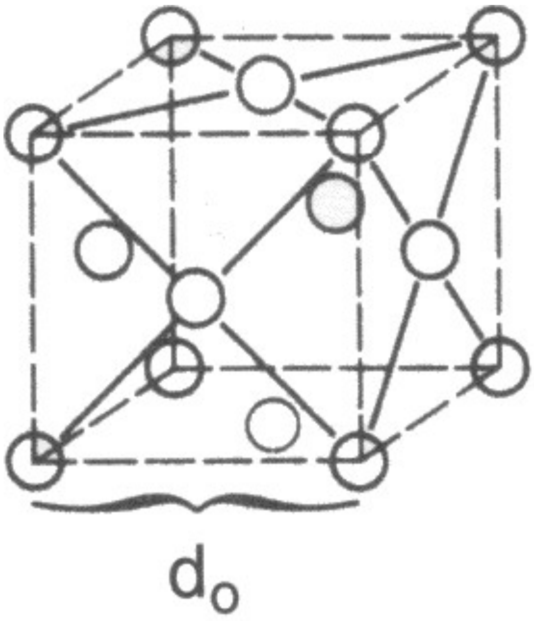


Figure Three: *Face Centered Cubic Crystal*

This expression can be used to find the volume of the cube; hence, this relationship is important for FCC cubic systems. The **coordination number** in an FCC lattice is 12, implying that FCC lattices are quite stable.

The close-packed layers of atoms in the FCC lattice are not parallel to the unit cell faces, but rather are perpendicular to the cell diagonal. If you look down the cell diagonal, you see six atoms in a close-packed triangle in the layer immediately behind the corner atom, and another layer of close-packed atoms below that, followed by another corner atom. The layers are indeed closely packed, and as one goes down the diagonal of this and succeeding cells, the layers repeat their positions in the order ABCABC…. This implies that atoms in every fourth layer lie below one another (see Figure Four (b)).

Hexagonal Close-Packing

There is another way to stack the layers as in the FCC lattice, above. The first and second layers will always be in the same relative positions, but the third layer could be below the first one if it were shifted properly. This results in a **close-packed structure** in which the order of the layers is ABABAB… (see Figure Four (a))

The crystal obtained from this arrangement of layers is not cubic but **hexagonal**. It is another common structure for metals. Cadmium, zinc and manganese have this structure. As you might expect, the stability of this structure is very similar to that of FCC crystals. We find that simply changing the temperature often converts a metal from one form to another. Calcium, for example, is FCC at room temperature, but if heated to 450 °C it converts to close-packed hexagonal.

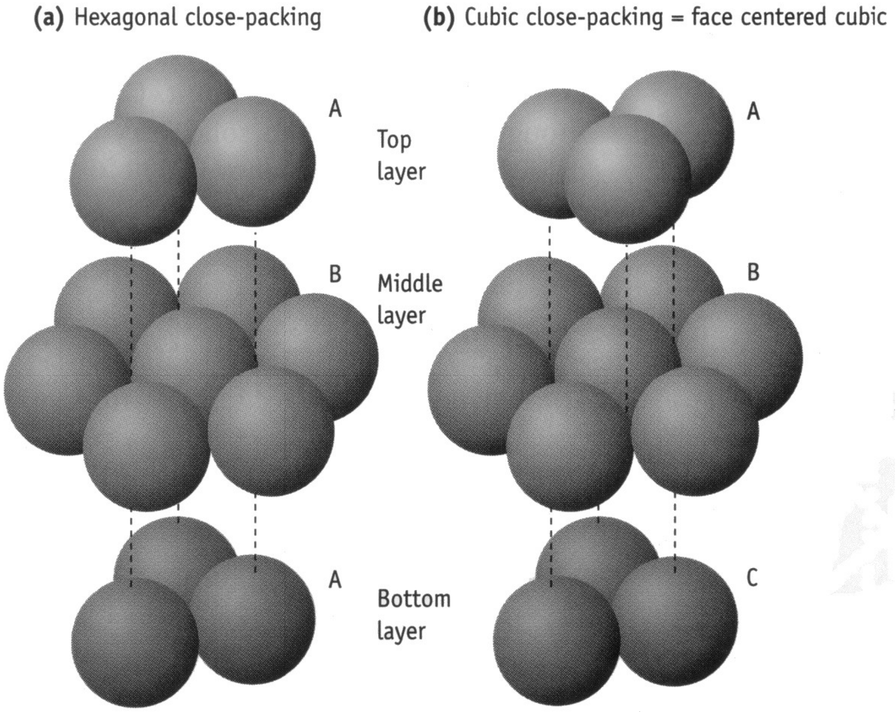
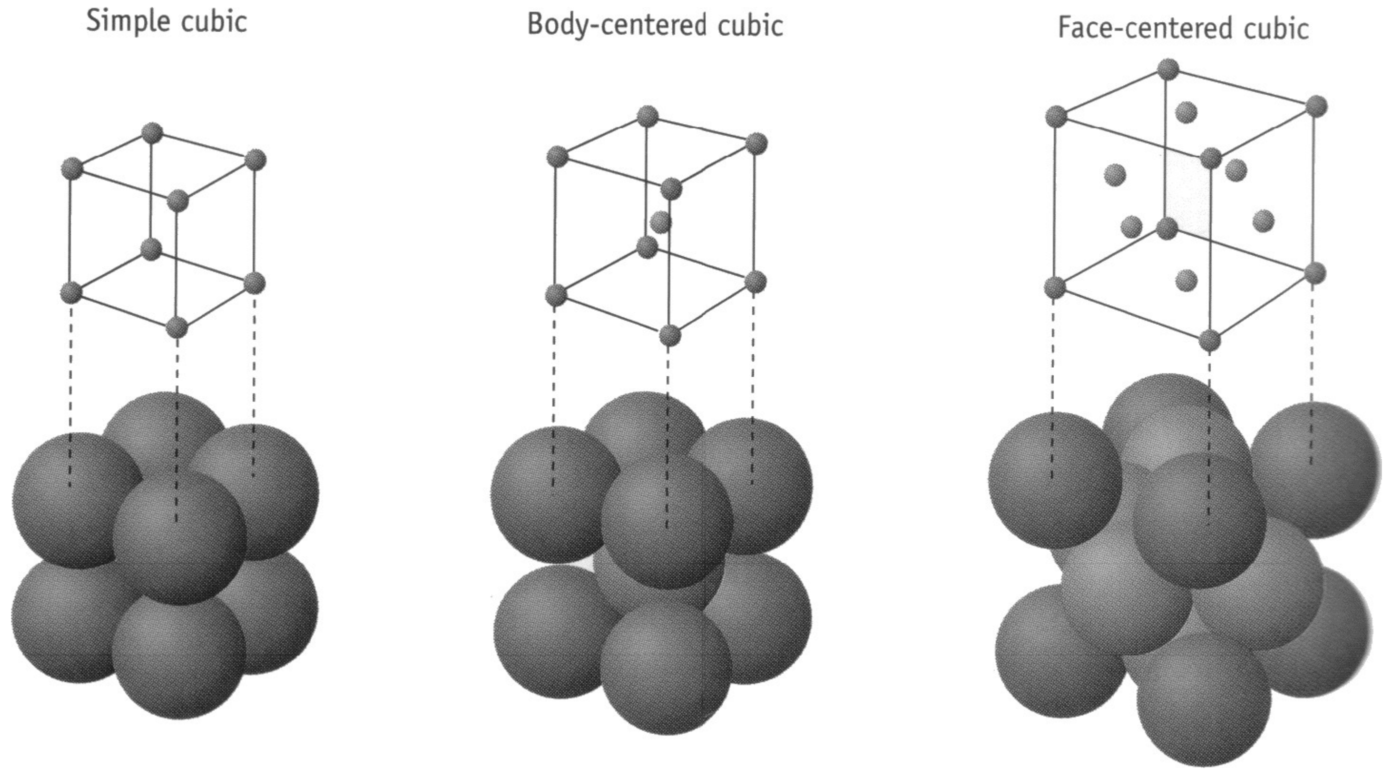


Figure Four: *Hexagonal Close Packing (left) and Cubic Close Packing (right)*

In CH 222 (and CH 223), we will consider "hexagonal close-packing" structures to be identical to FCC lattices, but technically there are many differences between the two systems.

Summary of Crystal Lattice Types

Figure Five shows the three main cubic unit crystal types that we will explore in this lab. **Figure Six** show a helpful methodology to solve problems like these in this lab.



Lattice Type	**Simple Cubic**	**Body Centered Cubic**	**Face Centered Cubic**
# net atoms per cell	1	2	4
d_0 (edge) in relation to r	$d_0 = 2r$	$d_0 = \frac{4r}{\sqrt{3}}$	$d_0 = \frac{4r}{\sqrt{2}}$

Figure Five: *Summary of the Three Cubic Unit Cell Types*

conversion *V = edge³* *density* *molar mass (g/mol)* *Avogadro (6.022 x 10^{23})*

radius ↔ edge ↔ volume ↔ mass (g) ↔ moles ↔ atoms / molecules

1 pm = 10^{-12} m / 1 Å = 10^{-10} m / 1 cm = 10^{-2} m *4 atoms = 1 fcc cell, etc.*

Figure Six: *Helpful Conversion Methodology*

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Worksheet: Linear Regression & The Crystal Structures of Solids **Name:**

All final answers must be provided on this worksheet. **Include computer generated graphs** (from Excel or a similar program) **along with any necessary calculations** needed for the lab in your submission. Values of r should be recorded to three or more significant figures. **This page *will* be placed at the front of your lab, with your first and last name included, to prevent a point penalty.**

Linear Regression:

- **<u>Problem 1</u>:** ***The Relationship Between °C and °F*** *- see page Ia-6-3 for data and questions*

Linear Regression equation: **y =** ______________________________

r = ______________ *Percent Error:* ______________________

- **<u>Problem 2</u>:** ***Solubility of Lead(II) Nitrate in Water*** *- see page* Ia-6-3 *for data and questions*

Linear Regression equation: **y =** ______________________________

r = ______________ *Solubility of lead(II) nitrate at 47.0 °C:* ______________

- **<u>Problem 3</u>:** ***Colorimetry*** *- see page* Ia-6-4 *for data and questions*

Linear Regression equation: **y =** ______________________________ **r =** ____________

Absorbance of 2.50 mg permanganate in 100 mL solution: ______________________

Linear Regression: *Continued*

- **Problem 4:** ***Kinetics*** *- see page* Ia-6-4 *for data and questions*

Linear Regression (ln [NH$_3$*] vs. time) equation:* **y =** __________________________________ **r =** __________

Linear Regression (1/[NH$_3$*] vs. time) equation:* **y =** __________________________________ **r =** __________

Which regression gives a better linear regression? Why?

Does this data behave as a first order reaction or a second order reaction?

- **Problem 5 *(Perform in the lab)*:** Experimentally determine the density of an unknown metal solid to at least three significant figures using any equipment found in your lab drawer. Explain the process (and show calculations) used to determine the density in three sentences or less **on this sheet**. *Hint:* use the displacement method. What liquid did you use?

density (g/cm^3) = __________________ **unknown letter used =** __________

Relevant calculations:

The Crystal Structures of Solids:

- **Problem 6:** What element forms a face centered cubic cell, has a density of 8.92 g/cm^3, and a radius of 128 pm?

Element = __________ *Show relevant work below*

The Crystal Structures of Solids: *Continued*

- **Problem 7:** Chromium forms a body centered cubic crystal. If the length of an edge is 2.884 angstroms, calculate the **density** (g/cm^3) and the **radius** of a chromium atom in angstroms.

density (g/cm^3) = __________________ **radius (Å) = ____________**

Show relevant work below.

The Crystal Structures of Solids: *Continued*

- **Problem 8:** Sodium (radius = 186 pm) forms a body centered cubic crystal. Calculate the **density** (g/cm^3) of sodium metal. **Propose a simple experiment to confirm your calculated density of sodium** in the lab.

density (g/cm^3) = _________________

Show relevant work below.

Proposed simple experiment:

The Crystal Structures of Solids: *Continued*

- **Problem 9:** Aluminum crystallizes in a face centered cubic unit cell. In addition, aluminum has an atomic radius of 143 pm. What is the density of aluminum?

density (g/cm^3) = _________________

Show relevant work below

CH 222 Winter 2025: **"Molar Mass Determination by Freezing Point Depression** *(in class)"* **Lab** *- Instructions*

Note: **This is the lab for section 01 and H1 of CH 222 only.**

- *If you are taking section W1 of CH 222*, please use this link:
  http://mhchem.org/r/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-7-11 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, February 17 (section 01) *or* **Wednesday, February 19 (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-7-7 through Ia-7-11 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, February 24 (section 01)** *or* **Wednesday, February 26 (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!

Molar Mass Determination by Freezing Point Depression

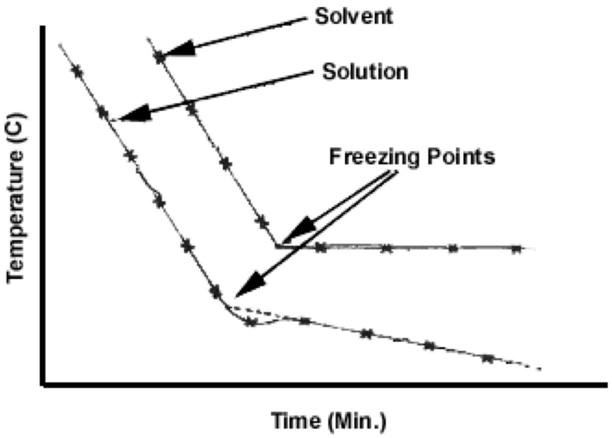
If a nonvolatile solute is added to a liquid, a number of physical properties of the pure substance change, including vapor pressure depression, freezing point depression, and boiling point elevation. These alterations are collectively known as **colligative properties** of solutions. The colligative properties of a solution change in proportion to the concentration of solute dissolved in solvent and depend only on the number of solute particles present in a given amount of solvent and not on the type of particles dissolved. Therefore, the concentration of the solute is most conveniently expressed in terms of **molality** (m) or **moles of solute / kg of solvent**.

The change in the freezing point (ΔT_f) in ˚C for a nonvolatile organic solvent can be determined using the following equation, where $\mathbf{k_f}$ is characteristic for the solvent used: $\mathbf{\Delta T_f = k_f \cdot} \boldsymbol{m}$

We can determine the **molar mass** of the solute using this equation by measuring the change in the freezing point of the solution and solving the equation for molality. The calculated molality can be used to determine the moles of solute that in turn can be used to calculate the molar mass (grams / mole) of the solute.

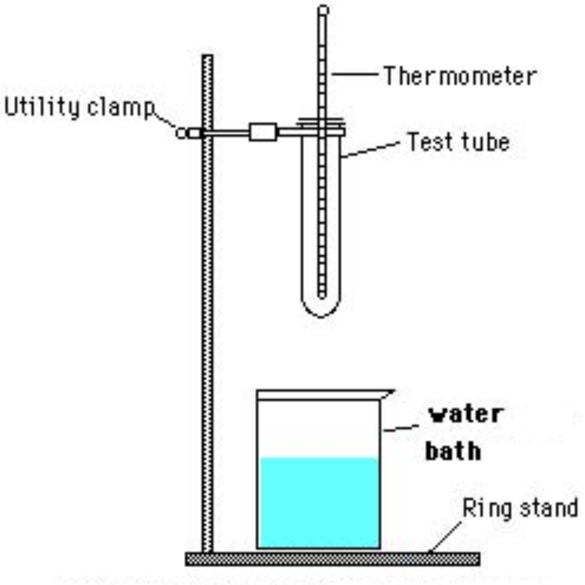
When a pure substance is heated to a liquid state and allowed to cool, initially the temperature will fall quite rapidly. As the substance approaches its **freezing point**, solid will begin to form and the temperature will begin to level. The freezing point of the pure liquid is the constant temperature observed while the liquid is solidifying.

The cooling behavior of a solution is somewhat different from that of a pure liquid. The temperature at which the solution begins to freeze is **lower** (i.e. *depressed*) than for the pure solvent. Additionally, there is a slow gradual fall in temperature as freezing proceeds. The **change in temperature, ΔT**, between the freezing point of the pure substance and the freezing point of the solution is used to calculate the molality of the solution.



In both the pure liquid and the solution, a **supercooling** effect may be seen. As the solid begins to form, the temperature may drop *below* the actual freezing point initially and then come back up to the freezing point temperature as the solid forms. Supercooling is usually not observed if adequate churning of the sample is provided. When determining the freezing point, the super-cooling effect should be ignored.

In this experiment, you will first determine the freezing point of a pure solvent, **lauric acid** ($C_{12}H_{24}O_2$). Next, you will use a known solute, **benzoic acid**, to depress the freezing point of the solvent and calculate the molar mass of the benzoic acid.



This picture should also include a hot plate under the beaker and a "swizzle stick" around the thermometer

PROCEDURE: Part A: Determining the Freezing Point for Lauric Acid

1. Set up ring stand and test tube clamp beside the hot plate

2. Fill a large beaker about ¾ full with tap water and heat on hot plate.

3. Mass 10.000-12.000 g of solid lauric acid ($C_{12}H_{24}O_2$) and record the mass. Add the lauric acid to a large test tube.

4. Place the test tube in the beaker and heat until the lauric acid is completely melted. Put a "swizzle stick" around a thermometer and place it into the liquid. Gently stir with the swizzle stick (not the thermometer!) to uniformity. Do not overheat - you will need to cool it in step 5 (perhaps read ahead?) You may need to readjust your thermometer so it remains in the solution.

5. Remove the test tube from the hot water and allow it to cool. **Once the temperature reaches 50 °C**, begin to record the temperature every 30 seconds until the temperature plateaus for 3 minutes or reaches 35 °C. (This should take 5-10 minutes). Mix ***gently*** to maintain uniformity (test tubes are glass! Careful!) The biggest source of error in this lab is the temperature so use care in reading. Make sure you record the temperature to the correct number of significant figures – if unsure, ask!

PROCEDURE: Part B: Determination of the Molar Mass for Benzoic Acid

6. Mass 1.000 - 1.200 g of benzoic acid and record the mass. Add to the test tube with the lauric acid. Be careful not to lose any sample of lauric acid that may have solidified to the thermometer or swizzle stick.

7. Place test tube in hot water and melt the mixture, stirring to uniformity.

8. Remove the test tube from the hot water and allow it to cool. Once the temperature reaches 50 °C, begin to record the temperature every 30 seconds until the temperature plateaus for 3 minutes or reaches 35 °C. (This should take 5-10 minutes). Mix ***gently*** to maintain uniformity (test tubes are glass! Careful!)

9. Repeat steps 6 through 8, adding an additional 1.000 - 1.200 grams of benzoic acid to the solution. Do ***NOT*** make a new solution!

10. Time to clean up! To clean up, re-melt your solid and pour into waste container. You may need to add a tiny bit of hot water along the sides of the tube to help melt the solid. Wipe the thermometer and swizzle stick with a wee bit of acetone and/or methanol if needed.

CALCULATIONS: Determining the Molar Mass of the Benzoic Acid

Once you have completed your experiment, you will need to create a **graph in Excel** (or a similar program; no hand drawn graphs will be accepted.) Use **time** as the x-axis and **temperature** as the y-axis. Plot the data from your pure lauric acid experiment (Part A) and also the data from each of the benzoic acid solutions (Part B) **using the same graph**. Label the axes and use different colors and/or marking symbols to distinguish the trials.

Important: your graph should take up an entire page of paper (no small graphs, print in "landscape" mode), *and* the **y-axis should *not* start at zero** (i.e. if your data points end around 25 °C, have the minimum y-axis value be 20 °C, not zero °C (which is what these programs often default to when creating graphs.)

Determine the freezing point for each trial graphically; **draw a circle** on your graph representing the freezing point in each trial. Remember, the freezing points can be found at the intersection between the "steep" cooling curve and the "gradual" cooling curve.

The k_f value for lauric acid is 3.90 °C/m. Using the data from Part A and Part B, calculate the molar mass of benzoic acid in each trial. Show how you calculated these values in your lab report.

Average the two molar mass values and calculate the **parts per thousand**.

Using the Internet or a textbook, determine the structure of benzoic acid and **include a Lewis structure for benzoic acid in your lab report**. (Hand drawn Lewis structures are ok.)

Determine the actual molar mass of benzoic acid using the Lewis structure. **Calculate the percent error** of benzoic acid using your average molar mass value and the accepted molar mass value. *Recall:* **Percent error = absolute value{(actual - experimental)/ actual}*100%.**

Molar Mass Determination by Freezing Point Depression

Record all masses to the milligram (0.001 g) and temperatures to at least 0.1 °C

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

DATA:

lauric acid mass (g): ______________________

1st benzoic acid mass (g): ______________________

2nd benzoic acid mass (g): ______________________

Record temperatures on next page.

Time (minutes)	pure lauric acid (deg C)	lauric acid + 1st benzoic acid (deg C)	lauric acid + both benzoic acids (deg C)
0.5 1.0			
1.5 2.0			
2.5 3.0			
3.5 4.0			
4.5 5.0			
5.5 6.0			
6.5 7.0			
7.5 8.0			
8.5 9.0			
9.5 10.0			
10.5 11.0			
11.5 12.0			

Note: you may not need all of the spaces in the above chart... you might also need more, depending on your solution. Feel free to include an additional page of data if needed (but most labs should probably fit in the above area.)

ANALYSIS:

Graph your freezing point data in Excel, Numbers, Sheets or a similar computer program (hand drawn graphs will not be accepted for credit.) Use **Time (minutes)** as your x-axis and **Temperature °C** as your y-axis. The **graph should take up an entire piece of paper** (about 8.5 x 11 inches, print in landscape mode) - no small graphs!

- You will have **three different data streams on the same graph.** Label the data streams as:

  a. **Lauric Acid** (pure solvent)
  b. **Solution I** (lauric acid and the first benzoic acid sample) and
  c. **Solution II** (lauric acid and both benzoic acid samples)

- The data streams will **use a common x axis**; use a color or point marking system to differentiate the different data streams.

.

- Ensure that your temperature axis reflects an appropriate range of values for the data provided (i.e. ***do not start at zero!***) Attach the graph to this lab when turning it in to the instructor.

- On your graph, determine the freezing point for the “Lauric Acid”, "Solution I” and “Solution II” curves. Look for the spot where two slopes on the same line come together; this is the official freezing point. **Mark the graph at the freezing point** (pen or pencil is ok), and enter the actual numbers below.

Freezing Point of Lauric Acid (°C): _______________

Freezing Point of "Solution I" (°C): _______________

Freezing Point of "Solution II" (°C): _______________

- Using the internet, include a Lewis structure for both lauric acid **and** benzoic acid. (hand drawn ok!)

- Using the mass of lauric acid, the mass of benzoic acid (1st sample only), and the freezing points of the pure lauric acid and “Solution I”, **determine the experimental molar mass of benzoic acid** from the “Solution I” data. (**k_{fp}** for lauric acid = **3.90 °C/m**) *Show work!*

- Using the mass of lauric acid, the mass of benzoic acid (both samples together), and the freezing points of the pure lauric acid and “Solution II”, **determine the experimental molar mass of benzoic acid** from the “Solution II” data. (**k_{fp}** for lauric acid = **3.90 °C/m**) *Show work!*

Benzoic acid molar mass value from Solution I *(g/mol):* ________________________________

Benzoic acid molar mass value from Solution II *(g/mol):* ________________________________

Average the two molar mass values and calculate the **parts per thousand**. *Show work!*

Determine the actual molar mass of benzoic acid using the Lewis structure, above.

Calculate the percent error of your benzoic acid molar mass calculation using your average molar mass value and the actual molar mass value.

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

POSTLAB QUESTIONS:

1. Determine the effect of the following on the final molar mass calculation in a freezing point depression experiment. *Briefly* give your reasoning. Indicate if the effect on molar mass will be **higher**, **lower** or **not change**.

   i. The thermometer you were using read temperatures consistently 1.2 ˚C higher than the real temperature.

   Effect on molar mass:

   Explain:

   ii. You knowingly added 3 g of unknown solute during your freezing point determination instead of 2 g.

   Effect on molar mass:

   Explain:

   iii. In adding your unknown solute to your solvent, you spilled some solute onto the table after weighing.

   Effect on molar mass:

   Explain:

2. A student performs a freezing point analysis. She determines that the freezing point of 21.00 g of stearic acid (where $k_f = 4.89°C/m$) is 68.20 ˚C. She adds 2.07 grams of an unknown compound to her sample and determines the freezing point to be 65.53 ˚C. She adds an additional 1.97 g of the unknown compound and determines the new freezing point to be 63.03 ˚C.

   a. Determine the molar mass of the unknown compound using the 2.07 g of sample.

   b. Determine the molar mass of the unknown compound using the *combined* samples (*hint:* 2.07 + 1.97 = 4.04 g total solute)

   c. Determine the average molar mass of the compound and the parts per thousand for the two trials.

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CH 222 Winter 2025: **"Kinetics I - The Iodination Of Acetone** *(in class)"* ***Lab*** *- Instructions*

Note: **This is the lab for section 01 and H1 of CH 222 only.**

- *If you are taking section W1 of CH 222*, please use this link:
  http://mhchem.org/r/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-8-15 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, February 24 (section 01) *or* **Wednesday, February 26 (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-8-5 through Ia-8-15 *only* to avoid a point penalty) **at the beginning of recitation to the instructor on Monday, March 3 (section 01)** *or* **Wednesday, March 5 (section H1.)** The graded lab will be returned to you the following week during recitation.

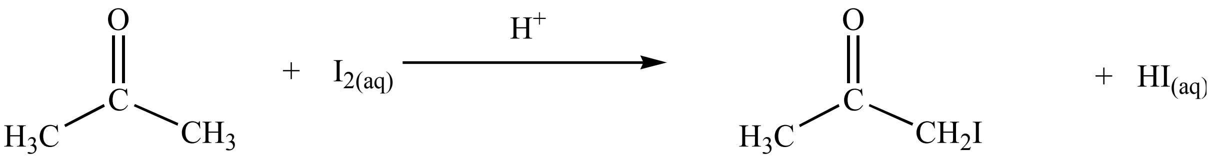
If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Kinetics I - The Iodination Of Acetone

Determining the Rate Constant for a Chemical Reaction

The **rate** of a chemical reaction depends on several factors: the nature of the reaction, the concentrations of the reactants, the temperature, and the presence of a possible catalyst. In *Part One* of this experiment we will determine the **rate law** for a reaction by changing some of the above variables and measuring the rate of the reaction. During *Part Two*, we will explore the relation between the rate constant and temperature to discover the **activation energy** for this reaction.

In this experiment, we will study the kinetics of the reaction between iodine and acetone:



The rate of this reaction is found to depend on the concentration of the hydrogen ion (acid, HCl) as well as the concentrations of the reactants (acetone and iodine). The **rate law** for this reaction is

$$\mathbf{rate = k[acetone]^m[H^+]^n[I_2]^p}$$

where **k** is the **rate constant** for the reaction and **m**, **n**, and **p** are the **orders of the reaction** with respect to acetone, hydrogen ions (acid), and iodine, respectively. Although orders of reaction can be any value, for this lab we will be looking only for **integer values** for the orders of reaction (0, 1, 2 are acceptable but not 0.5, 1.3, etc.)

The **rate** of the reaction can also be expressed as the change in the concentration of a reactant divided by the time interval:

$$rate = \frac{-\Delta[I_2]}{\Delta t}$$

The iodination of acetone is easily investigated because iodine (I_2) has a deep yellow/brown color. As the acetone is iodinated and the iodine converted to the iodide anion, this color will disappear, allowing the rate of the reaction to be easily monitored.

We can study the rate of this reaction by simply making I_2 the ***limiting reactant*** in a large excess of acetone and H^+ ion. By measuring the time required for the initial concentration of iodine (I_2) to be used up completely, the rate of the reaction can be determined by the equation

$$rate = \frac{-\Delta[\mathrm{I}_2]}{\Delta \mathrm{t}} = \frac{-([I_2]_{final} - [I_2]_{initial})}{\mathrm{t}_{final} - t_{initial}} = \frac{-(0 - [I_2]_{initial})}{\mathrm{t}_{final} - 0} = \frac{[I_2]_{initial}}{\mathrm{t}_{final}}$$

or simply as

$$rate = \frac{[I_2]}{time}$$

From the rate information, we can determine the **orders** with respect to acetone (**m**), acid (**n**) and iodine (**p**) by varying the amounts of reactants and measuring the effect on the rate. Once the orders of reaction are known, we will be able to calculate the rate constant, k. In Part One of this experiment you will determine the rates of reactions, the orders of the reactants, and finally the rate constant at room temperature.

DIRECTIONS:

You will collect data during lab and complete the worksheets at the end. No formal typed lab report is required as long as your writing is legible and easy to read. You will perform a series of experiments that will examine the relationship between the concentration of reactants and the time for the iodination of acetone reaction.

For each trial listed below: measure out the appropriate quantities of **1.0 M HCl, 4.0 M acetone** and **water** using a 10.00 mL graduated cylinder and place them in a 125 mL Erlenmeyer flask. Now measure out the appropriate amount of **0.0050 M iodine** in a 10.00 mL graduated cylinder.

Start a timer (stopwatch) as you add the iodine to the 125 mL flask with the other chemicals. Swirl the flask **until the yellow color disappears**, **then halt the timer**. It may help to place the flask on a white piece of paper to help discern when the color disappears. Record the time elapsed in seconds.

Repeat this reaction mixture until **two** trials are within 20 seconds of each other.

Repeat this process for each of the four trials listed in the table below. Waste can be placed in the drain or in a waste bottle (probably the better option!)

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Name:
Lab partner(s):

Kinetics I - The Iodination Of Acetone

Part I: Changing Concentration to Find the Rate Constant - *Complete in the Lab*

For each mixture listed below, add all of the chemicals but iodine to a 125 mL Erlenmeyer flask. Add the iodine last, starting a stopwatch and measuring how long the reaction takes to turn the solution clear. Time should be recorded in seconds. **Repeat** each reaction mixture until **two** trials are within 20 seconds of each other.... repeat the trial again if the times vary too much.

Trial #1:

HCl (mL)	Acetone (mL)	I_2 (mL)	Water (mL)	Total Volume (mL)
5	5	5	10	25.00

Time in seconds for yellow color to disappear, first time: ____________________ seconds

Time in seconds for yellow color to disappear, second time: ____________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ____________________ seconds

Trial #2:

HCl (mL)	Acetone (mL)	I_2 (mL)	Water (mL)	Total Volume (mL)
5	5	10	5	25.00

Time in seconds for yellow color to disappear, first time: ____________________ seconds

Time in seconds for yellow color to disappear, second time: ____________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ____________________ seconds

Trial #3:

HCl (mL)	Acetone (mL)	I_2 (mL)	Water (mL)	Total Volume (mL)
5	10	5	5	25.00

Time in seconds for yellow color to disappear, first time: ___________________ seconds

Time in seconds for yellow color to disappear, second time: ___________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ___________________ seconds

Trial #4:

HCl (mL)	Acetone (mL)	I_2 (mL)	Water (mL)	Total Volume (mL)
10	5	5	5	25.00

Time in seconds for yellow color to disappear, first time: ___________________ seconds

Time in seconds for yellow color to disappear, second time: ___________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ___________________ seconds

Before you leave, check your data. Trial #1 should be about half the time of trial #2, and trial #3 and trial #4 should both be about the same time elapsed. If you don't see this trend, contact the instructor before leaving lab.
.
You are now ready to complete the Kinetics I lab!

Kinetics I - The Iodination Of Acetone – *Worksheet*

These steps can be done at home and do not need to be completed in the lab. Show your work for each step

The ultimate goal of this section is to find the best value of the rate constant, k, for the iodination of acetone at room temperature. To get there, we need to first find the rates of each reaction, then the order of the reactants (acetone, HCl and iodine), and then finally, the rate constant k.

a. **Find the average time in seconds** for each Trial in Part One.

Example: the first experiment took 230 seconds, the second experiment took 250 seconds.

The *average time* would be: (230 + 250) / 2 = **240 seconds**

Record your reaction times from Part I (which were within 20 seconds of each other) and the final average times here:

	Experiment #1 (s)	Experiment #2 (s)	Average time (s)
Trial #1	________________	________________	________________
Trial #2	________________	________________	________________
Trial #3	________________	________________	________________
Trial #4	________________	________________	________________

Show an example of how you got the average values in seconds here:

b. **Find the concentration of each reactant** (acetone, HCl and iodine) before the reaction started.

Your group took bulk reactants (which were, as a reminder: **1.0 M HCl, 4.0 M acetone**, and **0.0050 M iodine**) then placed them (with water) in an Erlenmeyer flask. The final volume was always 25.00 mL. Mixing chemicals dilutes the concentrations from the "bulk" value to a smaller value.

We can determine these diluted values using: $\mathbf{M_1V_1 = M_2V_2}$

Example: Determine the concentration of iodine in trial #1.

Let $\mathbf{M_1}$ = initial (undiluted) concentration of iodine (0.0050 M), $\mathbf{V_1}$ = 5.00 mL (of undiluted iodine added to the mixture), and $\mathbf{V_2}$ = 25.00 mL (the total volume of the diluted solution once HCl, acetone and water are added). Solving for M_2, the concentration of iodine in the diluted solution, one gets:

> $\mathbf{M_2}$ = 0.0050 M * 5.00 mL / 25.00 mL = **0.0010 M**, which is the concentration of iodine used in the reaction in trial #1.

Example: Determine the concentration of iodine in trial #2.

Solution: Since $\mathbf{M_1}$ and $\mathbf{V_2}$ are the same as in the previous example, we see that only $\mathbf{V_1}$ has changed to 10.00 mL. Rearranging for $\mathbf{M_2}$ as before:

> $\mathbf{M_2}$ = 0.0050 M * 10.00 mL / 25.00 mL = **0.0020 M**, the concentration of iodine used in the reaction in trial #2.

i. **Calculate the concentration of acetone ($\mathbf{M_2}$) used in Trial 1 - Trial 4.**

The **concentration** of acetone was 4.0 M in the "bulk" solution (the "$\mathbf{M_1}$" value.)
The final volume ($\mathbf{V_2}$) is always 25.00 mL.
Trial 1, Trial 2 and Trial 4 used 5.00 mL of acetone from the bulk solution, but Trial 3 used 10.00 mL (your "$\mathbf{V_1}$" values)

	volume acetone (mL)	**Bulk Acetone (M)**	**Total Volume (mL)**	**acetone (M)**
Trial #1	5.00	4.0 M	25.00	__________
Trial #2	5.00	4.0 M	25.00	__________
Trial #3	10.00	4.0 M	25.00	__________
Trial #4	5.00	4.0 M	25.00	__________

ii. **Calculate the concentration of HCl (M_2) used in Trial 1 - Trial 4.**

The **concentration** of HCl was 1.0 M in the "bulk" solution (the "M_1" value.)
The final volume (V_2) is always 25.00 mL.
Trial 1, Trial 2 and Trial 3 used 5.00 mL of acetone from the bulk solution, but Trial 4 used 10.00 mL (your "V_1" values)

	volume HCl (mL)	**Bulk HCl (M)**	**Total Volume (mL)**	**HCl (M)**
Trial #1	5.00	1.0 M	25.00	________
Trial #2	5.00	1.0 M	25.00	________
Trial #3	5.00	1.0 M	25.00	________
Trial #4	10.00	1.0 M	25.00	________

iii. **Complete the following table** showing the diluted concentrations of all reactants used in each trial. The I_2 concentrations have been completed for you (see the examples at the beginning of part b, above.)

	acetone (M) *step i, above*	**HCl (M)** *step ii, above*	**I_2 (M)**
Trial #1	________	________	0.0010
Trial #2	________	________	0.0020
Trial #3	________	________	0.0010
Trial #4	________	________	0.0010

Use this area to show relevant calculations and at least one example as to how you found a diluted concentration.

c. **Find the rate of each trial**

In this lab, rate is best described by: **rate = $[I_2]$/(average time in seconds)** (see first page of lab)
Find the rate for each reaction by using the calculated iodine concentration then dividing by the average time for that trial (Part Three, section a, above).

Example: Trial #1 required an average of 240 seconds to go to completion. Trial #2 required an average of 496 seconds to go to completion. Find the rate of reaction for both trial #1 and trial #2.

Solution: The rate of reaction is equal to the concentration of iodine divided by the average time elapsed for the reaction. (your values are in Part Three, section a.)

In this example, the first two times are 240 s and 496 s. Using the appropriate concentrations, we can calculate rate:

rate (trial #1) = $[I_2]$/(average time in seconds) = 0.0010 M / 240 s = **$4.2 * 10^{-6}$ M s^{-1}**.

rate (trial #2) = $[I_2]$/(average time in seconds) = 0.0020 M / 496 s = **$4.0 * 10^{-6}$ M s^{-1}**.

Now, **calculate your rate values** by filling in the table below:

	I_2 (M)	**average time (s)**	**rate (M s^{-1})**
Trial #1	0.0010	__________	__________
Trial #2	0.0020	__________	__________
Trial #3	0.0010	__________	__________
Trial #4	0.0010	__________	__________

Use this space to show at least one example of how you calculated the rate of the reaction.

d. **Find the order** of the reaction with respect to each reactant (**m**, **n**, and **p**)

To find the order of each reactant, we need to localize the effect that each reactant had on the rate. This is expressed in a **rate order** which, for CH 222 and CH 223, **can only be equal to zero, one or two** (no fractions, negative numbers, etc.)

In this lab, it is safe to assume that **water does not affect the rate**, so we can ignore its contribution to any changes that the rate might exhibit.

Notice that **trial 2 has twice as much iodine as trial 1**. Any changes to the rates of trial 1 and trial 2 are a direct result of the iodine (HCl and acetone are constant), and we will use these trials to calculate **p, the reaction order for iodine.**

Also notice that **trial 3 has twice as much acetone as trial 1**, so any changes to the rates of trials 3 and 1 will be the result of acetone only (HCl and iodine are constant), and we will use these trials to calculate **m, the reaction order for acetone.**

Lastly, notice that **trial 4 has twice as much HCl as trial 1**, yet acetone and iodine concentrations remain constant; we will use these trials to calculate **n, the reaction order for HCl.**

Example: Find the order of reaction with respect to iodine (p) if these values are used:

rate (trial #1) = $[I_2]$/(average time in seconds) = 0.0010 M / 240 s = $\mathbf{4.2 * 10^{-6}\ M\ s^{-1}}$.

rate (trial #2) = $[I_2]$/(average time in seconds) = 0.0020 M / 496 s = $\mathbf{4.0 * 10^{-6}\ M\ s^{-1}}$.

Solution: Notice how in trial #2 we doubled the concentration of $[I_2]$ while leaving the other reactants' concentrations (HCl, acetone) unchanged. An effect upon the rate of the reaction will reflect the influence of the iodine only, and this will allow us to determine **p**.

In these sample calculations, doubling the concentration of iodine (to 0.0020 M from 0.0010 M) spawned a negligible change in the rate ($4.0 * 10^{-6}$ M s^{-1} versus $4.2 * 10^{-6}$ M s^{-1}). Because we are only concerned with whole integer values of rate orders, this implies a **zero order reactant**, and **p = 0**.

A more formal approach to determining p would be as follows: divide the expression for rate 2 by the expression for rate 1; this results in the following:

$$\frac{\text{rate 2}}{\text{rate 1}} = \frac{k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p}{k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p}$$

The values of k, [acetone] and [HCl] remain constant between trial 1 and trial 2 (only $[I_2]$ changes), so the expression reduces to

$$\frac{\text{rate 2}}{\text{rate 1}} = \frac{4.0 * 10^{-6}}{4.2 * 10^{-6}} = \frac{[0.0020]^p}{[0.0010]^p} = 2^p$$

$$0.95 = (2)^p$$

Taking the logarithm of both sides leads to

$$\log 0.95 = \log 2^p = p \log 2$$

and solving for p:

$$p = \frac{\log 0.95}{\log 2} = -0.074 \approx 0$$

Therefore, the order with respect to iodine equals **zero**, or **p = 0**.

Now, **find the order of each reactant** by filling in the tables below. ***Note:*** **rate** values from Part 3 section c; **concentration** (M) values from Part 3, section b, subsection iii. Show your work on the following page. ***Remember*** that m, n and p can be **0, 1, or 2 only**! Round your answers as necessary!

For I_2 (p):

	I_2 (M)	**rate (M s^{-1})**	**My value of p is:**
Trial #1	0.0010	__________	__________
Trial #2	0.0020	__________	

For acetone (m):

	acetone (M)	**rate (M s^{-1})**	**My value of m is:**
Trial #1	__________	__________	__________
Trial #3	__________	__________	

For HCl (n):

	HCl (M)	**rate (M s^{-1})**	**My value of n is:**
Trial #1	__________	__________	__________
Trial #4	__________	__________	

Use this space to show **how** you got each of the orders of reaction (m, n and p):

e. **Find the value of k, the rate constant** for the iodination of acetone reaction.

You now have all the necessary information to calculate the rate constant, k, for each trial. For this reaction,

$$\text{rate} = k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p$$

Rate values appear in Part 3 section c
[acetone], [HCl] and [I2] are the concentrations for each trial (Part 3, section b, subsection iii)
m, n and p are the orders of reaction (Part 3, section d)

Example: Find the value of k for trial #1 if the rate = 4.2×10^{-6} M s^{-1} and assuming that the order with respect to acetone (m) and HCl (n) is two and the order with respect to I_2 (p) is zero.

Solution: In trial #1, the diluted concentration of acetone is 0.80 M, the HCl is 0.20 M and I_2 is 0.0010 M. Using the given values, we can calculate k from the following equation:

$$\text{rate} = k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p$$

$$4.2 * 10^{-6} = k[0.80]^2[0.20]^2[0.0010]^0$$

$$k = 4.2 * 10^{-6} / [0.80]^2[0.20]^2$$

and solving for k we get a value of **$k = 1.6 * 10^{-4}$ M^{-1} s^{-1}**

Now **find the rate constant k for each trial** by completing the table:

My value of m = __________ *(these can be found in section d, above)*

My value of n = __________

My value of p = __________

	acetone (M)	**HCl (M)**	**I_2 (M)**	**rate (M s^{-1})**	**value of k**
Trial #1	__________	__________	0.0010	__________	__________
Trial #2	__________	__________	0.0020	__________	__________
Trial #3	__________	__________	0.0010	__________	__________
Trial #4	__________	__________	0.0010	__________	__________

Concentrations in section b, subsection iii *Rate values in section c*

Average value of k: ______________________

parts per thousand of your four k values: ______________________

Use this space to show **a sample calculation for k** and also relevant **parts per thousand calculations**:

f. **You are done!** Finish the postlab questions (which are similar to the work you just completed) and you are good to go!

Kinetics I - The Iodination Of Acetone - Postlab Questions:

1. In a reaction involving the iodination of acetone, the following reaction mixture was used: 5.00 mL 4.0 M acetone, 5.00 mL 1.0 M HCl, 5.00 mL 0.0050 M I_2, and 10.0 mL water.

    a. What was the molarity of the acetone in the reaction mixture? (Recall that $M_1V_1 = M_2V_2$)

    b. The color of the above reaction mixture disappeared in 250 seconds. What was the rate of the reaction? (Hint: First determine the initial concentration of the iodine, then use the equation for rate from the lab.)

2. A second reaction mixture was made: 10.00 mL acetone, 5.00 mL HCl, 5.00 mL I_2, and 5.00 mL of H_2O.

    a. What was the molarity of the acetone in this reaction mixture?

    b. The iodine color disappeared in 120 seconds. What was the rate of the reaction?

    c. Determine the order of the reaction (m) with respect to acetone using the information from question 1 and 2. (Round off the value of m to the nearest integer)

3. A third reaction mixture is made: 10.00 mL acetone, 5.00 mL HCl, 10.00 mL I_2. If the reaction is zero order with respect to iodine, how long will it take for the iodine color to disappear? (*Hint:* rate = $[I_2]$/time, use the rate from question 2b and the new $[I_2]$ to solve for the time elapsed.)

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CH 222 Winter 2025:
"Kinetics II - The Iodination Of Acetone *(in class)*" ***Lab*** *- Instructions*

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

- *If you are taking section W1 of CH 222*, please use this link:
  http://mhchem.org/r/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-9-11 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, March 10 (section 01) *or* **Wednesday, March 12 (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-9-5 through Ia-9-11 *only* to avoid a point penalty) **to the instructor at the time of your Lecture Final** *(tentatively Wednesday, March 19.)* **Be sure to include any necessary computer generated graphs as well.**

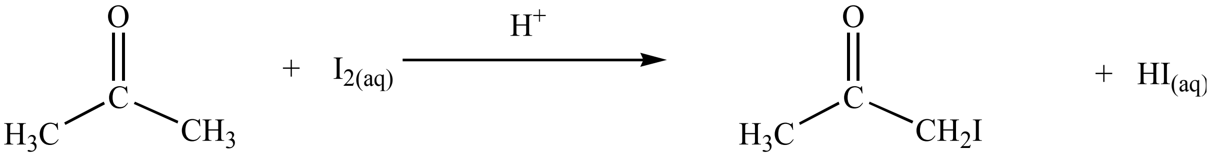
If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Kinetics II - The Iodination Of Acetone

Determining the Activation Energy for a Chemical Reaction

The **rate** of a chemical reaction depends on several factors: the nature of the reaction, the concentrations of the reactants, the temperature, and the presence of a possible catalyst. In *Part One* of this experiment we determined the **rate law** for a reaction by changing the concentrations and measuring the rate of the reaction. During *Part Two*, we will explore the relation between the rate constant and temperature to discover the **activation energy** for this reaction.

In this experiment, we continue to study the kinetics of the reaction between iodine and acetone:



In last week's lab, you discovered the **average rate constant, k**, value, as well as the **orders of reaction** (**m, n and p**) which apply to the **rate law** for this reaction:

$$\textbf{rate} = \mathbf{k[acetone]^m[H^+]^n[I_2]^p}$$

In this experiment, you will study the rate of the reaction at different temperatures to find its **activation energy**, **E_a**. The temperature at which the reaction occurs influences the rate of the reaction. An increase in temperature increases the rate.

As with concentration, there is a quantitative relation between reaction rate and temperature, but here the relation is somewhat more complicated. This relation is based on the idea that to react, the reactant species must have a certain minimum amount of energy present (and the correct geometry, if appropriate) at the time the reactants collide in the reaction step.

This amount of energy, which is typically furnished by the kinetic energy of the species present, is called the **activation energy** (**E_a,** also known as the *energy of activation*) for the reaction. The formula (called the *Arrhenius equation*) relating the rate constant k to absolute Kelvin temperature T and E_a is:

$$\ln k = \frac{-E_a}{RT} + \ln A$$

In this equation, **R** is the gas constant (**8.3145 J/mole K**), and natural logarithms (**ln**) need to be used (*do not use base 10 logs!*) The quantity **A** is referred to as the **collision frequency** and A can be used to determine the fraction of molecules present with sufficient energy and geometry to become products at a given instant in time.

By measuring k at different temperatures, we can graphically determine the activation energy for a reaction. In Part Two of this experiment you will determine the effect of temperature on rate and calculate the activation energy for the reaction.

DIRECTIONS:

You will collect data during lab and complete the worksheets at the end. No formal typed lab report is required as long as your writing is legible and easy to read. In this experiment, you will see the effect of temperature upon the reaction.

We shall measure one of the reactions from the "Kinetics I" lab at different temperatures. You do not need to repeat these experiments twice (to be within 20 seconds of each other) as in Part One.

For each entry in trial #5 listed below: measure out the appropriate quantities of **1.0 M HCl, 4.0 M acetone** and **water** using a 10.00 mL graduated cylinder and place them in a 125 mL Erlenmeyer flask. Now measure out the appropriate amount of **0.0050 M iodine** in a 10.00 mL graduated cylinder.

Using ice and/or a hot plate, get the solution to a desired **temperature** before adding the iodine. **Record the temperature**, then add the iodine to the Erlenmeyer, start a stopwatch and measure how long the reaction takes to turn the solution clear. Time should be recorded in **seconds**. The iodine does not need to be at the same temperature as the solution in the Erlenmeyer flask.

Temperatures need to be higher than 15 °C (too slow!) and lower than 60 °C (keep the acetone from boiling), and the interval between measurements needs to be at least 5 °C apart.

Waste can be placed in the drain or in a waste bottle (which is probably the better option!) Get a stamp in your lab notebook before leaving lab.

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Name:
Lab partner(s):

Kinetics II - The Iodination Of Acetone

Part Two: The Effect of Temperature on the Rate Constant - *Complete in the Lab*

As before, add all the chemicals but iodine to a 125 mL Erlenmeyer flask. Using ice and/or a hot plate, get the solution to a desired **temperature** before adding the iodine. **Record the temperature**, then add the iodine to the Erlenmeyer, start a stopwatch and measure how long the reaction takes to turn the solution clear. Time should be recorded in **seconds**. The iodine does not need to be at the same temperature as the solution in the Erlenmeyer flask.

* Record one trial at **room temperature**. *Hint:* you *may* be able to use some of your data from Part I!
* Record one trial at a **temperature lower than room temperature, but above 15 degrees Celsius**.
* Record **three trials at temperatures higher than room temperature, but under 60 degrees Celsius**.

Temperature differences should be at least 5 degrees Celsius apart (i.e. spread out your temperatures!)

<u>Trial #5</u>:

HCl (mL)	Acetone (mL)	I_2 (mL)	Water (mL)	Total Volume (mL)
5	5	5	10	25.00

Temperature (°C): ___________________ Time (seconds): _______________

Temperature (°C): ___________________ Time (seconds): _______________

Temperature (°C): ___________________ Time (seconds): _______________

Temperature (°C): ___________________ Time (seconds): _______________

Temperature (°C): ___________________ Time (seconds): _______________

Hint: As temperature increases, the reaction time should decrease. If you do not see this trend, repeat one or more experiments.

You are now ready to complete the Kinetics II lab!

Kinetics II - The Iodination Of Acetone – *Worksheet*

These steps can be done at home and do not need to be completed in the lab. Show your work for each step

Earlier you determined the time elapsed for a given set of concentrations as the temperature was altered. We shall use that information and techniques similar to that of the "Kinetics Part I" lab to determine the **energy of activation** and **collision frequency** for the iodination of acetone reaction using the Arrhenius equation.

a. **Find the inverse Kelvin temperature** for each value in Trial #5.

Convert your temperatures from °C to K, then take the *inverse* of your Kelvin temperatures.

> *Example:* Convert 37.5 °C to an inverse Kelvin temperature.
>
> *Solution:* 37.5 °C = 310.7 K. To find the inverse, calculate $(310.7\ \mathrm{K})^{-1}$ = $\mathbf{3.219 * 10^{-3}\ K^{-1}}$

Complete the table below. The first column (Temperature (°C) comes from your data collected in Part Two, Trial #5, while in lab.

	Temperature (°C)	**Temperature (K)**	**Temperature^{-1} (K^{-1})**
Temp #1	__________	__________	__________
Temp #2	__________	__________	__________
Temp #3	__________	__________	__________
Temp #4	__________	__________	__________
Temp #5	__________	__________	__________

Use this space to show **a sample calculation** for getting from Temperature (°C) to an inverse Kelvin temperature:

b. **Find the rate** for each temperature value in Trial #5.

Recall from the "Kinetics Part I" lab that, for this experiment:
rate = [I_2]/(time in seconds)

Use this equation to **find the rate of reaction** (in M s^{-1}) for each temperature. Time values come from Trial #5, above:

	I_2 (M)	**time (s)**	**rate (M s^{-1})**
Trial #1	0.0010	__________	__________
Trial #2	0.0010	__________	__________
Trial #3	0.0010	__________	__________
Trial #4	0.0010	__________	__________
Trial #5	0.0010	__________	__________

Use this space to show at least one example of how you calculated the rate of the reaction.

c. **Find the value of the rate constant, k,** for each temperature value in Trial #5.

We will use the process developed in "Kinetics Part I" lab to help us find the values of k for each temperature.

i. First, we need the **diluted concentrations**: *(these can be found in section b of the "Kinetics Part I" lab)*

Concentration (M) of **acetone** when 5.00 mL was used: _______________

Concentration (M) of **HCl** when 5.00 mL was used: _______________

Concentration (M) of **I_2** when 5.00 mL was used: _______________

ii. Next, we need the **reaction orders** for each reactant: *("Kinetics Part I", section d)*

My value of m (acetone) = __________

My value of n (HCl) = __________

My value of p (I_2) = __________

If you do not have m = 1, n = 1, p = 0, talk to the instructor!

iii. Now use the techniques from in Kinetics I", section e, to find the value of k, and then ln k.

Use the **rates from section b** *(above)* and the values for concentration and order (m, n and p) **to find k**. The only variable that will change is the rate; the orders and concentrations remain constant. Take the **natural log (ln) of each k** value as well (i.e. ln ($2.6 * 10^{-5}$) = **-10.56)** (Note: *report your ln k values to the **hundredths** place to satisfy significant figures.*)

	rate (M s^{-1})	**k**	**ln k**
Trial #1	__________	__________	__________
Trial #2	__________	__________	__________
Trial #3	__________	__________	__________
Trial #4	__________	__________	__________
Trial #5	__________	__________	__________

Show a sample calculation for these steps on the next page.

Use this space to show at least one example of how you calculated the rate constant k and ln k.

d. **Create a graph** of ln k versus inverse Kelvin temperature values

You will be creating a graph of ln k versus inverse temperature to find the energy of activation.

First, collect your inverse temperature (the x-axis) and ln k values (the y-axis) here. Inverse Kelvin temperature values come from section a in this lab. ln k values come from section c, subsection iii, in this lab.

	Temperature^{-1} (K^{-1})	**ln k**
Trial #1	__________	__________
Trial #2	__________	__________
Trial #3	__________	__________
Trial #4	__________	__________
Trial #5	__________	__________

Using Excel or a similar program, create a **graph** of your ln k values versus the inverse Kelvin temperature values. Make the graph at least as big as half a sheet of paper, and be sure to include unit labels (ln k for the y-axis and (Temperature)$^{-1}$ for the x-axis. *Note* that when using your graphing program, you may need to enter values as decimals, i.e. enter 0.00315 instead of 3.15×10^{-3}.

Staple / attach the graph to the end of this lab report packet.

e. **Find the energy of activation** for the iodination of acetone using the data in Trial #5.

The data points on the graph from the last section should correspond roughly to a straight line with a negative slope. This is the behavior predicted by the Arrhenius equation:

$$\ln k = \frac{-E_a}{RT} + \ln A$$

where **ln k** is the y-axis, **(Temperature in Kelvin)$^{-1}$** is the x-axis, **$-E_a/R$** is the slope, **$R = 8.3145$ J K^{-1} mol^{-1}** (the "energy" gas constant), **E_a** is the energy of activation, and **A** is the collision frequency.

Perform a linear regression analysis using your calculator or graphing program (inverse Kelvin temperatures will be your x-axis, ln k values will be your y-axis.) Record the values that you collected here:

Slope = ___________ y-intercept = _________ correlation coefficient (r) = ____________

The **energy of activation**, E_a, can be determined from the slope. From the value of the slope determined through linear regression, **calculate the activation energy**.

Energy of activation = **–R*slope** = _____________

Units for your Energy of activation value = _____________

The collision frequency, A, can be determined from the y-intercept. From the value of the y-intercept determined through linear regression, **calculate the collision frequency**.

Collision frequency = **e$^{\text{y-int}}$** = _____________

Note that **e** is the anti natural logarithm.

f. **You are done!** Finish the postlab question (which is similar to the work you just completed) and you are good to go!

Kinetics II - The Iodination Of Acetone - Postlab Questions:

The following reaction

$$2\ N_2O_{5(g)} \rightarrow 4\ NO_{2(g)} + O_{2(g)}$$

was studied at several temperatures, and the following values of k were obtained:

k (s^{-1})	T (˚C)
$2.0*10^{-5}$	20.0
$7.3*10^{-5}$	30.0
$2.7*10^{-4}$	40.0
$9.1*10^{-4}$	50.0
$2.9*10^{-3}$	60.0

Using linear regression and the techniques developed in this lab, calculate the **activation energy** and **collision frequency** for this reaction. **Include a computer generated graph** of ln k versus $(T)^{-1}$. *Hint:* make sure you use inverse Kelvin temperatures! Make sure the x-axis lists "0.003" numbers (and not whole integers, etc.)

Slope = ___________ y-intercept = _________ correlation coefficient (r) = ____________

Energy of activation = _____________

Units for your Energy of activation value = _____________

Collision frequency = _____________

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“Introduce Yourself” Lab for Chemistry 222 Section W1

***Create a video, sign the form below** and turn in via email to mike.russell@mhcc.edu by **9 AM, Friday, January 10**.*
Remember to turn in the video link to me as well!

Note: **This is the lab for section W1 of CH 222 only.**

- *If you are taking section 01 or section H1 of CH 222*, please use this link:
  http://mhchem.org/r/1a.htm

Welcome to Chemistry 222! I am glad you enrolled in CH 222 this quarter, and I look forward to an exciting term with you!

This class will be quite different from previous Chemistry 222 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 222? 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 "by hand".**

- 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 222 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 222 Section W1

Create a video, sign the form below** and turn in everything via email to mike.russell@mhcc.edu by **9 AM, Friday, January 10**. Also remember to turn in the v**ideo 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 222**. 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 222 Winter 2025:
“Lewis Structures *(online)*” *Lab - Instructions*

Note: **This is the lab for section W1 of CH 222 only.**

- *If you are taking section 01 or section H1 of CH 222*, please use this link: http://mhchem.org/r/2a.htm

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Step One:

Watch the lab video for the “Lewis Structures” lab, found here: **http://mhchem.org/y/2.htm**
There is no data to record at the end of the video.

Step Two:

Complete pages Ib-2-7 through Ib-2-12 using the “Lewis Structures” video and the actual lab instructions on pages Ib-2-2 through Ib-2-6. Include your name on page Ib-2-7!

Step Three:

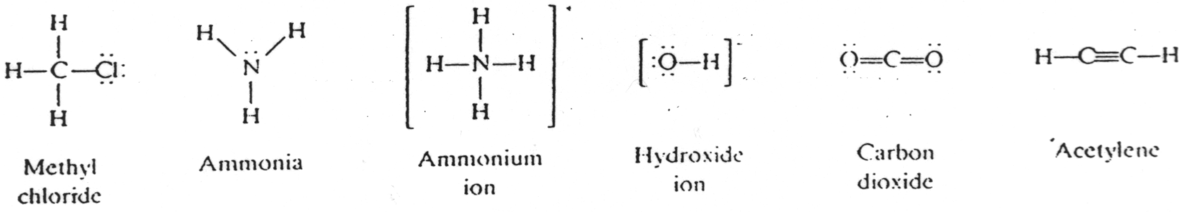
Submit your lab (pages Ib-2-7 through Ib-2-12 *only* to avoid a point penalty**) as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, January 22 by 11:59 PM.** I recommend a free program (ex: CamScanner, https://camscanner.com) or a website (ex: CombinePDF, https://combinepdf.com) 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!

Lewis Structures - The Geometry of Covalent Molecules

The formulas of many covalent compounds, especially those involving only the elements of the first few periods of the periodic table, were brought within the framework of the **octet rule**. The octet rule states that a total of eight **valence** electrons, either shared or unshared, should be in the region of each atom beyond the first period. For hydrogen the desired number is two. Electrons that are shared between two atoms are **bonding electrons** and are to be counted toward the octet of each of the bonding atoms. A single covalent bond consists of a pair of shared electrons, a double bond has two shared pairs, and a triple bond has three shared pairs. Bond distances are shorter and bond energies are greater for multiple bonds than for single bonds.

Structural formulas, such as shown in Figure 1, represent the electron distributions in covalent molecules and ions. These structures are not meant to indicate actual bond angles in threedimensional varieties; they merely show the number of bonds connecting the various atoms. In **Lewis formulas**, a single line between two atoms represents a **pair** of shared electrons and a **dot** represents an unshared electron.



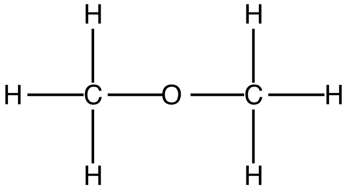
Methyl chloride | Ammonia | Ammonium ion | Hydroxide ion | Carbon dioxide | Acetylene

Figure 1

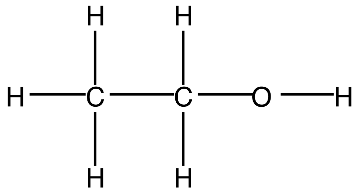
Two lines constitute a *double bond,* and *three lines* constitute a *triple bond.* The total number of electrons shown in such a molecular structure is equal to the sum of the numbers of valence (outershell) electrons in the free atoms: 1 for H, 4 for C, 5 for N, 6 for O, and 7 for Cl. For an ionic structure, one additional electron must be added to this sum for each unit of negative charge on the whole ion, as in OH^-, and one electron must be subtracted from the sum for each unit of positive charge on the ion, as in NH_4^+. The number of pairs of electrons shared by an atom is called its **covalence**.

The covalence of hydrogen is always one. The covalence of oxygen is practically always one or two. The covalence of carbon is four in almost all its stable compounds. Thus each carbon is expected to form either four single bonds, a double bond and two single bonds, two double bonds, or a single and a triple bond. Although the octet rule is not a rigid rule of chemical bonding, it is obeyed for C, N, O, and F in almost all their compounds. The octet is exceeded commonly for elements in the third and higher periods of the periodic table.

Isomerism: For some molecules with a given molecular formula, it is possible to satisfy the octet rule with different atomic arrangements. A simple example would be

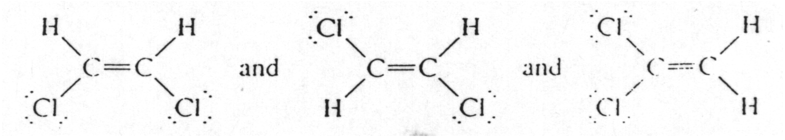


and



The two molecules are called **isomers** of each other, and the phenomenon is called **isomerism**. Although the molecular formulas of both substances are the same, C_2H_6O, their properties differ because of the different atomic arrangement. Since several different types of isomerism can exist especially in organic molecules, these are called **structural isomers**.

When **double bonds** are present, isomerism can occur in very small molecules. For example:



These isomers (called **geometric** isomers) result from the fact that there is no rotation around a double bond, as in single bonds.

Resonance: Sometimes more than one satisfactory structure can be written and there is no reason to select one over another. In such cases a single structural formula is inadequate to represent a substance correctly, and several such diagrams must be written. The true structure is then said to be a *resonance hybrid* of the several diagrams.

For example, experiment has shown that the two terminal oxygens in ozone are equivalent; that is, they are equidistant from the central oxygen. If only one of the resonance diagrams in Fig.2(a) were written, it would appear that one of the terminal oxygens is bonded more strongly to the central oxygen by a double bond than is the other (by a single bond) and that the more strongly bonded atom should be closer to the central atom. The hybrid of the two ozone structures gives equal weight to the extra bonding of the two terminal oxygen atoms. Similarly, the three resonance structures of carbonate in Fig.2(b) are needed to account for the experimental fact that all three oxygens are equidistant from the central carbon.

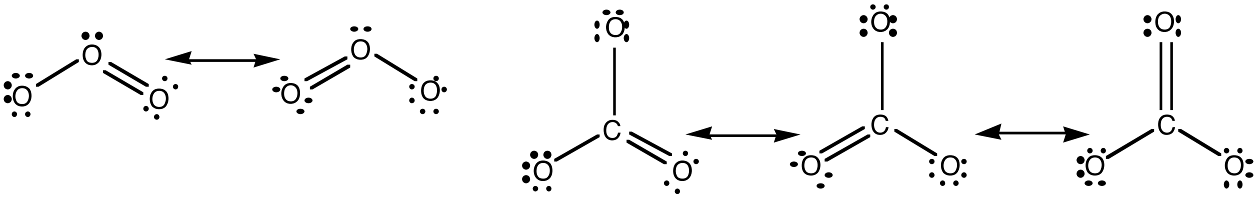


Figure 2: a) ozone *(left)* and b) carbonate *(right)*

The total bond energy of a substance for which resonance structures are written is greater than would be expected if there were only one formal Lewis structure. This additional stabilization is called **resonance energy**. It arises from the principle that is responsible for covalent bond energy, the **delocalization** of electrons about the atoms forming the bond. As a result of resonance in ozone, for example, the electrons constituting the second pair of the double bond are delocalized around the 3 oxygen atoms. The writing of two or more resonance structures is a way of overcoming the inability of a single valencebond structure to show this delocalization.

Formal Charge: Although a molecule as a whole is electrically neutral, it is a matter of much current interest to know whether there are local charges which can be identified with particular parts of a molecule, the algebraic sum of which would equal zero. In an ion, the algebraic sum would equal the charge of the ion as a whole. In one approximate method of apportioning charges within a molecule or ion, the shared electrons in a covalent bond are arbitrarily divided equally between the two atoms forming the bond. Unshared valence electrons on an atom are assigned exclusively to that atom. Each atom is then assigned a

formal charge which is equal to the number of valence electrons possessed by that atom in the neutral free state minus the number of valence electrons assigned to it in the structure. These charges may be written near the atoms on the structural diagrams.

Figure 3 shows a single resonance structure for ozone. The central oxygen is assigned just five electrons (two in the unshared pair plus half of the three pairs in the bonds); this atom, being one electron short of the complement of six in a free oxygen atom, is thus assigned a formal charge of **+1**. The terminal oxygen connected by a single bond is assigned 7 (6 in the unshared pairs plus half of one pair in the bond); having one electron more than a neutral oxygen atom, this atom is assigned a formal charge of **-1**. The other terminal oxygen has no formal charge because six electrons are assigned to it (four in the unshared pairs plus half of the two shared pairs).

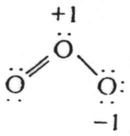


Figure 3: *Formal charges on ozone*

A rule useful in choosing one Lewis structure over another is that **structures which minimize formal charge separation are favored**. Especially to be avoided are formal charges of magnitude greater than 1 and structures in which appreciable formal charges of the same sign are located on adjacent atoms.

Dipole moments and electronegativity: There are some experimental procedures that give information about the actual distribution of charges within a molecule (as distinct from the arbitrary assignment of formal charges). One such is the measurement of **dipole moment**. An electric dipole is a neutral object that has a positive charge of magnitude q and a separately located, equal but opposite negative charge. The rotation of a dipole by an electric field is dependent upon and a measure of the dipole moment, defined as the product of q and the distance, d, separating the positive and negative charges.

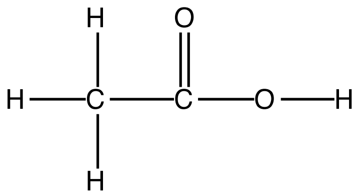
In a covalent diatomic molecule, the dipole moment would be expected to be zero if the bonding electrons were shared truly equally by the two atoms. This is indeed the case in molecules of the type XX where two identical atoms are bonded. In the more general type XY, two different kinds of atoms are bonded, and a dipole moment is usually observed. This is explained by hypothesizing that one of the atoms, say Y, has a greater attraction for the shared electrons in the bond than does X. Y is said to have a greater **electronegativity** than X. Electronegativity correlates with other atomic properties; in general, atoms with high ionization energies and/or high electron affinities tend to have high electronegativities. The most electronegative elements in order of decreasing electronegativity are **F > O > N = Cl** C is more electronegative than H. Metals are less electronegative than nonmetals.

HOW TO COMPLETE THIS LAB:

In this experiment you will be constructing models of molecules and ions containing atoms covalently bonded. If you have access to a model kit, use it! The model kit consists of plastic colored balls where each type of ball represents a certain species of atom; see the front cover of the model box to determine each atom type.

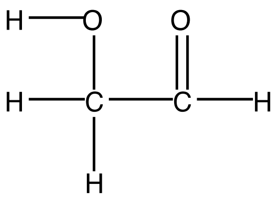
In assembling a molecular model of a molecule you should use a systematic approach keeping in mind the valence of the atoms in the molecule. Let us illustrate the recommended procedure with the following molecule, $\mathbf{C_2H_4O_2}$:

1. Draw a correct Lewis structure, remembering that carbon will *always* have four bonds and oxygen usually just two. Unless otherwise informed assume the carbon atoms form a carbon - carbon bond. Count up the total number of valence electrons (24 e^-). Those that are not bonding are nonbonding electrons. This leads to the following two isomers (other isomers exist as well):



acetic acid

and



hydroxy acetaldehyde (unstable)

2. Connect the balls and sticks to according to the Lewis structure. Use two bonds to represent the carbon-oxygen double bond.

3. Determine the electron pair geometry (EPG) and molecular geometry (MG) of the molecule from the molecular model. For acetic acid we say that the molecule is tetrahedral at one carbon, trigonal planar at the other carbon and tetrahedral / bent at the oxygen.

4. Determine whether any isomers exist by changing the location of the atoms; however you must maintain the correct valence of each atom.

5. Resonance will most often exist when the central atom forms a double bond to one atom and a single bond to another atom of the same element.

PROCEDURE:

Use the procedure outlined above to construct Lewis structures of the molecules and ions listed below using the attached worksheet. Neatness counts! Be sure to **include all lone pair electrons** on your drawings. For *each* molecule:

i) Draw the Lewis structure, including all lone pair electrons.
ii) Describe the electron pair geometry and molecular geometry around the central atom. If more than one central atom, put a star (*) by the central atom you will use in your answer.
iii) Determine if the molecule is polar or nonpolar and list any bond angles.
iv) Draw the Lewis structures of isomers and resonance forms, if any.

1. CH_4
2. CH_2Cl_2
3. CH_4O
4. H_2O
5. BF_3
6. NH_4^{+1}
7. CH_2O
8. NH_3
9. H_2O_2
10. HNO_3
11. C_2H_2
12. SO_2
13. C_2H_4
14. $C_2H_2Cl_2$
15. $C_2H_4Br_2$
16. SCN^-
17. NO_3^{-1}
18. SO_4^{2-}
19. CO_2
20. SeF_4
21. $AsCl_3Br_2$
22. PCl_5
23. SiF_6^{2-}
24. $SiF_4Cl_2^{2-}$
25. $SiF_3Cl_3^{2-}$

Also, on the last page, construct six possible isomers of $C_6H_{12}O_2$.

Name: ____

Lewis Structures Worksheet - *Hand drawn versions only, no computer generated structures*

Be sure to include all lone pair electrons for full credit! Draw all resonance forms and isomers (if any).

Molecule/Ion	Lewis Structure	Geometry and Polarity	Resonance/Isomers? (draw)
CH_4		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
CH_2Cl_2		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
CH_4O *hint: put a star next to your central atom*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
H_2O		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
BF_3		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	

The Geometrical Structure Of Covalent Molecules Worksheet - *Continued*

Molecule/Ion	Lewis Structure	Geometry and Polarity	Resonance/Isomers? (draw)
NH_4^{+1}		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
CH_2O		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
NH_3		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
H_2O_2		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
HNO_3 *hint: trigonal planar*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	

The Geometrical Structure Of Covalent Molecules Worksheet - *Continued*

Molecule/Ion	Lewis Structure	Geometry and Polarity	Resonance/Isomers? (draw)
C_2H_2 *hint: triple bond*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
SO_2		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
C_2H_4 *hint: double bond*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
$C_2H_2Cl_2$		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
$C_2H_4Br_2$		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	

The Geometrical Structure Of Covalent Molecules Worksheet - *Continued*

Molecule/Ion	Lewis Structure	Geometry and Polarity	Resonance/Isomers? (draw)
SCN^{-1} *C in middle*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
NO_3^{-1}		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
SO_4^{-2}		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
CO_2		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
SeF_4		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	

The Geometrical Structure Of Covalent Molecules Worksheet - *Continued*

Molecule/Ion	Lewis Structure	Geometry and Polarity	Resonance/Isomers? (draw)
$AsCl_3Br_2$ *hint: trigonal bipyramid*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
PCl_5		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
SiF_6^{2-} *hint:octahedral*		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
$SiF_4Cl_2^{2-}$		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	
$SiF_3Cl_3^{2-}$		***Electron Pair Geometry:*** ***Molecular Geometry:*** ***Polar or Nonpolar:*** ***Bond Angle(s):***	

Finally, **construct six possible isomers of $C_6H_{12}O_2$ in the space provided below**. For this question, carbon will always have four bonds and no lone pairs, and oxygen should always have two bonds (two single or one double) and two lone pairs.

CH 222 Winter 2025:
“Valence Bond and Molecular Orbitals *(online)*”
Lab *- Instructions*

Note: **This is the lab for section 01 *and* section W1 of CH 222 only.**

- *If you are taking section 01 or section H1 of CH 222*, please use this link:
  http://mhchem.org/r/3a.htm

Step One:

Watch the lab video for the “VB and MO” lab, found here: **http://mhchem.org/y/3.htm** There is no data to record at the end of the video.

Step Two:

Complete pages Ib-3-5 through Ib-3-12 using the “VB and MO” video and the actual lab instructions on pages Ib-3-2 through Ib-3-3. Include your name on page Ib-3-5!

Step Three:

Section 01: Submit your lab during recitation in **AC 2501 at 1:10 PM on Monday, January 27**. A **printed copy is required** for full credit; emailed labs will invoke a point penalty.

Section W1: **Submit your lab** (pages Ib-2-7 through Ib-2-12 *only* to avoid a point penalty) **as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, January 29 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!

VALENCE BOND (VB) THEORY

and

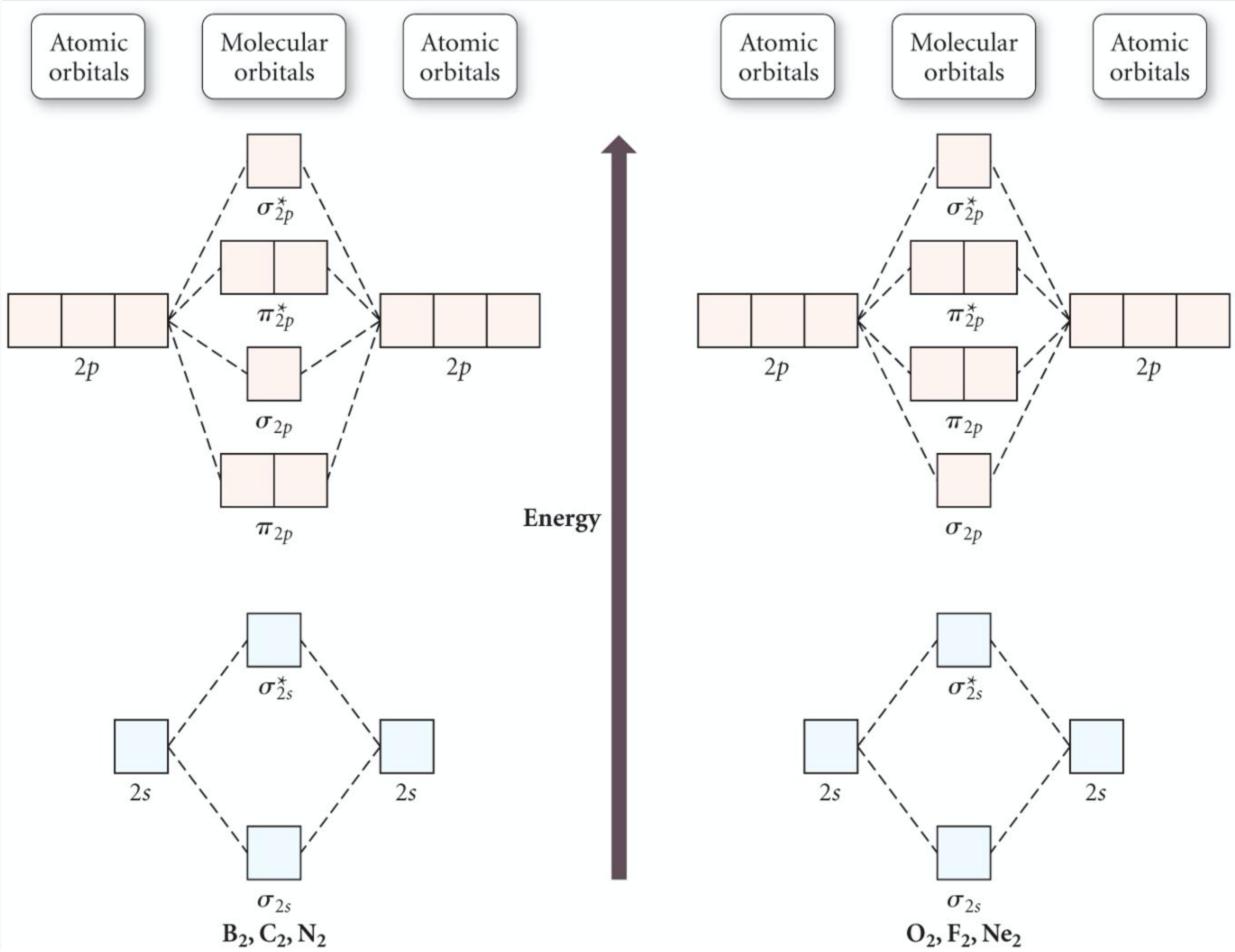
MOLECULAR ORBITAL (MO) THEORY

LAB

This lab contains a series of handouts that you should complete. Neatness counts!

Useful handouts from the Chemistry 222 website (http://mhchem.org/222):

- Geometry and Polarity Guide (http://mhchem.org/geopo/)
- MO Diagram - B_2 through N_2 (http://mhchem.org/MO)
- MO Diagram - O_2 through Ne_2 (http://mhchem.org/MO)



Molecular orbital diagrams for B, C, N (left) and O, F and Ne (right).
Notice that the 1s interactions are ***not*** *included in these diagrams.*
Use [core electrons] *if not showing the 1s interactions in your molecular orbital diagram.*

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VALENCE BOND THEORY and MOLECULAR ORBITAL THEORY LAB - ***Worksheet*** **Name:**

Hand drawn versions only, no computer generated structures

Complete the sections below by providing the appropriate information in the spaces provided. *Neatness counts!*

Part One: Valence Bond Theory *Complete the following sections using Valence Bond Theory.*

Molecule/Ion	Lewis Structure (& Isomers, if any)		
SeI_2		***Electron Pair Geometry:***	***Hybridization:***
		Molecular Geometry:	***Bond Order:***
		Polar or Nonpolar:	***Resonance Forms? (Y/N)***

Molecule/Ion	Lewis Structure (& Isomers, if any)		
$AsCl_5$		***Electron Pair Geometry:***	***Hybridization:***
		Molecular Geometry:	***Bond Order:***
		Polar or Nonpolar:	***Resonance Forms? (Y/N)***

Molecule/Ion	Lewis Structure (& Isomers, if any)		
CO_3^{2-}		***Electron Pair Geometry:***	***Hybridization:***
		Molecular Geometry:	***Bond Order:***
		Polar or Nonpolar:	***Resonance Forms? (Y/N)***

Molecule/Ion	Lewis Structure (& Isomers, if any)		
ClO_3^{-1}		***Electron Pair Geometry:***	***Hybridization:***
		Molecular Geometry:	***Bond Order:***
		Polar or Nonpolar:	***Resonance Forms? (Y/N)***

Molecule/Ion	Lewis Structure (& Isomers, if any)		
ClO_4^{-1}		*Electron Pair Geometry:* *Molecular Geometry:* *Polar or Nonpolar:*	*Hybridization:* *Bond Order:* *Resonance Forms? (Y/N)*

Molecule/Ion	Lewis Structure (& Isomers, if any)		
$XeOF_4$		*Electron Pair Geometry:* *Molecular Geometry:* *Polar or Nonpolar:*	*Hybridization:* *Bond Order:* *Resonance Forms? (Y/N)*

Part Two: Molecular Orbital Theory *Complete the following sections using Molecular Orbital Theory. Draw a complete Molecular Orbital diagram to answer these questions (include all 1s and 2s interactions, no short hand notation) and provide the missing information.*

Molecule / Ion: Li_2

Molecular Orbital Diagram:

Bond Order: ______ Number of sigma bonds: ______ Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic** Should this molecule exist? *(Circle)* **Yes** *or* **No**

Molecule / Ion: **Be_2**

Molecular Orbital Diagram:

Bond Order: ______ Number of sigma bonds: ______ Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic** Should this molecule exist? *(Circle)* **Yes** *or* **No**

Molecule / Ion: **B_2**

Molecular Orbital Diagram:

Bond Order: ______ Number of sigma bonds: ______ Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic** Should this molecule exist? *(Circle)* **Yes** *or* **No**

Molecule / Ion: **N_2**

Molecular Orbital Diagram:

Bond Order: ______ Number of sigma bonds: ______ Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic** Should this molecule exist? *(Circle)* **Yes** *or* **No**

Molecule / Ion: **F_2**

Molecular Orbital Diagram:

Bond Order: ______ Number of sigma bonds: ______ Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic** Should this molecule exist? *(Circle)* **Yes** *or* **No**

Molecule / Ion: **Ne_2**

Molecular Orbital Diagram:

Bond Order: ______ Number of sigma bonds: ______ Number of pi bonds: ______

(Circle) **Paramagnetic** *or* **Diamagnetic** Should this molecule exist? *(Circle)* **Yes** *or* **No**

Part Three: Theory Comparison

Complete the following sections using both Valence Bond (VB) Theory and Molecular Orbital (MO) Theory. ***Shorthand notation for MO Diagrams is recommended****, include 1s and 2s interactions.*

Molecule / Ion: **CN^{-1}** *(Use the MO Diagram for B, C and N on this problem)*

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes** **No**

(MO) Paramagnetic? *(circle)* **Yes** **No**

Molecule / Ion: **NO^{+1}** *(Use the MO Diagram for O, F and Ne on this problem)*

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes** **No**

(MO) Paramagnetic? *(circle)* **Yes** **No**

Molecule / Ion: **NO** *(Use the MO Diagram for O, F and Ne on this problem)*

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes** **No**

(MO) Paramagnetic? *(circle)* **Yes** **No**

Molecule / Ion: **O_2**

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes** **No**

(MO) Paramagnetic? *(circle)* **Yes** **No**

Molecule / Ion: **OF^{-1}**

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes** **No**

(MO) Paramagnetic? *(circle)* **Yes** **No**

Molecule / Ion: **Ne_2^{+2}**

Valence Bond Lewis Structure:

Molecular Orbital Diagram:

Bond Order (VB): ______

Bond Order (MO): ______

Number of sigma bonds (VB): ______

Number of sigma bonds (MO): ______

Number of pi bonds (VB): ______

Number of pi bonds (MO): ______

(VB) Paramagnetic? *(circle)* **Yes** **No**

(MO) Paramagnetic? *(circle)* **Yes** **No**

CH 222 Winter 2025:
“Organic Chemistry”
(online) ***Lab*** *- Instructions*

Note: **This is the lab for section W1 of CH 222 only.**

- *If you are taking section 01 or section H1 of CH 222*, please use this link: http://mhchem.org/r/4a.htm

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Step One:

Watch the lab video for the “Organic Chemistry” lab, found here: **http://mhchem.org/y/4.htm**
There is no data to record at the end of the video.

Step Two:

Complete pages Ib-4-9 through Ib-4-13 using the “Organic Chemistry” video and the actual lab instructions on pages Ib-4-2 through Ib-4-8. Include your name on page Ib-4-9!

Step Three:

Submit your lab (pages Ib-4-9 through Ib-4-13 *only* to avoid a point penalty) **as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, February 5 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!

Organic Chemistry

The classification of chemical compounds in to the general areas of organic and inorganic derives from the use of the "mineral, vegetable and animal" designation by the early workers in chemistry. Those compounds derived from living systems were termed **organic** (about 1777) whereas those derived from mineral sources were termed **inorganic**. In modern times, organic compounds are classified as compounds of carbon containing either carbon-carbon or carbon-hydrogen bonds or both. Originally, organic compounds were thought to be imbued with a "vital essence" attainable only from God. Thus, it was believed that organic compounds could be prepared from sources that had once lived, as this would be the only way that this vital essence could be obtained by man. In 1828, Freidrich Wöhler prepared the organic compound urea (found in human urine) from entirely nonliving sources, thereby destroying the theory of organic vitalism. Since Wöhler's time, approximately 5 million organic compounds have been synthesized and characterized, many of which are not found in nature.

Why are there so many organic compounds? The reason is that carbon atoms have the ability to link to other carbon atoms (concatenate) to produce chains or rings of almost infinite size. Other elements do not concatenate nearly as well due to such factors as poor orbital overlap and lone pair-lone pair electronic repulsions. Other elements can also combine with carbon to form hetero-species, including hydrogen, oxygen, nitrogen, sulfur and the halogens.

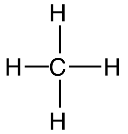
The distinction between the organic and inorganic disciplines is not very sharp. The bonding of metals to carbon has resulted in the large, important and fast growing area of **organometallic** chemistry. Organometallic compounds containing metals and metalloids such as lithium, magnesium, copper, iron, boron, silicon, and other elements play major roles as synthetic reagents.

The purpose of this experiment is to prepare models of the more common organic compound types to enhance their three dimensional nature. Bond angles between atom groupings will become apparent as you build the models.

Saturated Hydrocarbons: The Alkanes

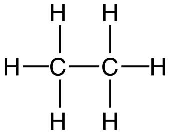
Alkanes are **hydrocarbons** (compounds with hydrogen and carbon atoms) linked with single bonds. All the carbon atoms are sp^3 hybridized and are tetrahedrally bonded to four other carbon or hydrogen atoms. Members of this class have the general formula $\mathbf{C_nH_{2n+2}}$, where **n** is an integer. Examples of this class of organic compound include methane (CH_4, n=1), ethane (C_2H_6, n=2) and propane (C_3H_8, n=3). Structures of the first four straight chain compounds are shown below.

methane



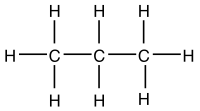
CH_4

ethane



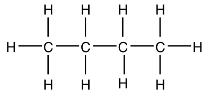
CH_3CH_3

propane



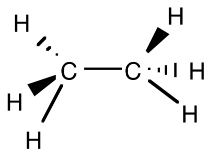
$CH_3CH_2CH_3$

butane

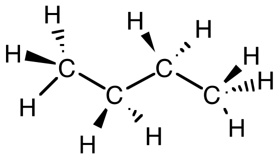


$CH_3CH_2CH_2CH_3$

These representations do not show the actual geometrical structure of these compounds. Due to the sp^3 hybridization of the carbon atoms, all the bond angles in the molecules are close to 109.5 ° (tetrahedral) and thus, the carbon chain is nonlinear. The **wedge-dash** notation can be used to represent the nonlinear nature of these molecules. Examples of ethane and butane are given below using the wedge-dash notation.



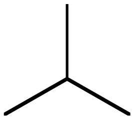
ethane



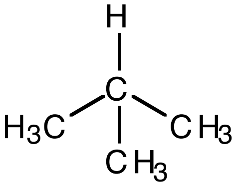
butane

As you prepare models of the alkanes, note that each carbon atom can rotate about its respective carbon-carbon bond. Some conformations (arrangements of the bonds and groups relative to each other) are more stable than others, since in these arrangements there is less interference (steric repulsion) between the hydrogen atoms attached to nearby carbons.

An interesting aspect of the alkanes is that **structural isomers** can exist in compounds having more than three carbon atoms. Structural isomers are species that have the same molecular formula but possess different physical properties due to different arrangements of the carbon backbone. The two compounds below are an example of an isomeric pair of hydrocarbons. Both are butanes and have the formula C_4H_{10}, but their structures are different.



$CH_3CH_2CH_2CH_3$

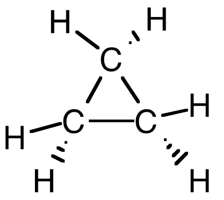


n-butane

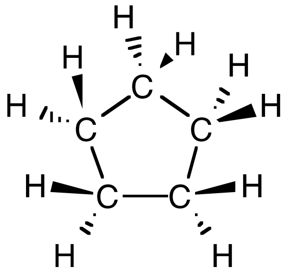
methylpropane (isobutane)

This figure relates several new points. First is the use of the **skeleton formula**. These structures represent the carbon backbone without the hydrogen atoms being shown and are often used as a shorthand method of representing the structures. The second point is the use of substituents in naming compounds. We can see that in methylpropane, one of the CH_3 groups is a "twig" off of the main "branch" of the compounds. Such twigs are called **alkyl groups**. In the figure, the **methyl group** (CH_3) name is derived from the hydrocarbon methane (CH_4) having lost a hydrogen. The number of possible isomers increases rapidly as the number of carbon atoms increases in a compound. For example, the pentane system (five carbons) has three isomers, the heptane system (seven carbons) has nine isomers, and the decane system (ten carbons) has seventy-five isomers!

Alkanes can also exist as **cyclic** hydrocarbons where the carbon atoms are arranged in rings. The general molecular formula is $\mathbf{C_nH_{2n}}$, where **n** is an integer. These carbons are named identically to the alkanes except for the additional prefix *cyclo*. The structures for two cyclic hydrocarbons are given below.



cyclopropane



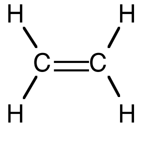
cyclopentane

In the smaller rings such as cyclopropane and cyclobutane, smaller bond angles (60° and 90°, respectively) are evident. Such angles are seen as being strained from their normal tetrahedral angle, and these compounds have less stability than their larger counterparts. The most stable cyclic compounds contain rings of five and six carbons (cyclopentane and cyclohexane, respectively).

Unsaturated Hydrocarbons: The Alkenes

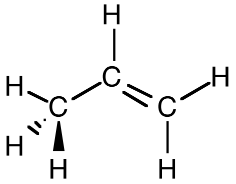
Alkenes are hydrocarbons in which there are one or more carbon-carbon double bonds, C=C. The carbon atoms attached to the double bond are sp^2 hybridized. One of the double bonds is a **sigma** bond (oriented along the internuclear axis) and the other is a **pi** bond (oriented perpendicular to the internuclear axis). Members of this class have the general formula $\mathbf{C_nH_{2n}}$, where **n** is an integer (note that this is the same formula as for the cycloalkanes.) Examples of alkenes include ethene (C_2H_4, n=2, also known as ethylene), propene (C_3H_6, n=3, also known as propylene) and butene (C_4H_8, n=4). Structures of several of the alkenes are shown below. Alkenes are named in the same fashion as the alkanes except that the ending *-ene* replaces the ending *-ane.* In cases where more than one structural isomer exists (such as butane, below), it is necessary to indicate the location of the double bond by numbering the carbon atoms in the longest chain containing the double bond, and then giving it the lowest possible number.

ethene (ethylene)



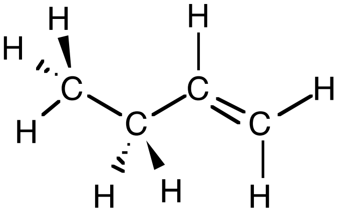
$CH_2{=}CH_2$

propene (propylene)



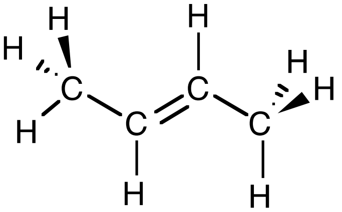
$CH_3CH=CH_2$

1-butene
or but-1-ene



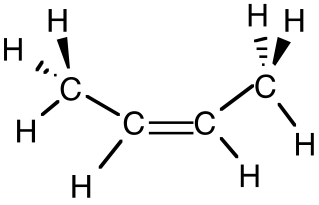
$CH_3CH_2CH=CH_2$

trans-2-butene
or trans-but-2-ene



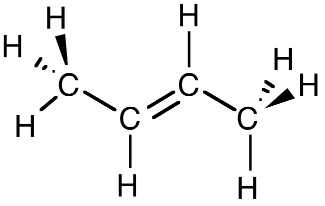
$CH_3CH=CHCH_3$

Many kinds of isomerism exist in organic compounds. Since carbon atoms are free to rotate around the C-C single bonds but <u>not</u> around a C=C double bond (this disrupts the overlap of the pi part of the double bond), it is possible to have two separate geometrical isomers of 2-butene. These are known as the *cis*- and *trans*- isomers and are shown below:



cis-2-butene

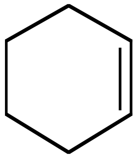
cis-but-2-ene



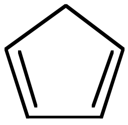
trans-2-butene

trans-but-2-ene

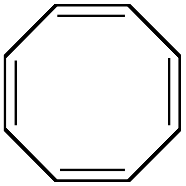
Cyclic structures containing C=C bonds are also possible, and three examples follow.



cyclohexene



1,3-cyclopentadiene



1,3,5,7-cyclooctatetraene

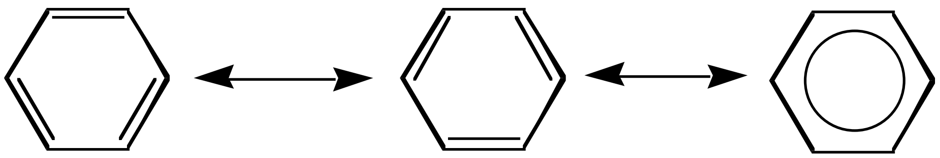
Unsaturated Hydrocarbons: The Alkynes

Hydrocarbons containing a carbon-carbon triple bond are named **alkynes**. The two carbon atoms forming the triple bond are joined by one sigma and two pi bonds and are sp hybridized. The alkynes have the general molecular formula $\mathbf{C_nH_{2n-2}}$. Several representatives of this class are shown below. Alkynes are names in the same manner as alkenes except that the ending *-yne* replaces the ending *-ene*.

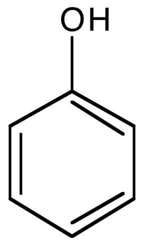
$H\text{-}C\equiv C\text{-}H$	$CH_3\text{-}C\equiv C\text{-}H$	$CH_3\text{-}C\equiv C\text{-}CH_3$
ethyne (acetylene)	propyne	2-butyne

Aromatic Hydrocarbons

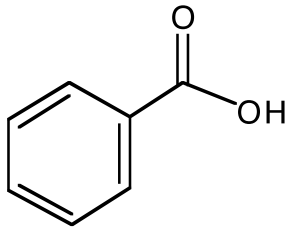
Benzene, C_6H_6, is the most crucial member of this important class of hydrocarbons. The molecule contains a ring of six sp^2 hybridized carbon atoms with the unhybridized p orbitals perpendicular to the ring system. The six electrons in the pi parts of the bonds are **delocalized** (spread evenly over the six carbon nuclei). Such molecules are often unusually stable. There are two equally valid ways of representing the structure of benzene. These two ways are called resonance forms, and the molecule is a **resonance hybrid** with the "true structure" of benzene lying midway between the two resonance forms. As an example, a mule is a genetic hybrid descendant of a male donkey and a female horse. The mule does not change back and forth, being a donkey half the time and a horse the other half. Thus, the properties of a resonance hybrid (such as benzene) are fixed.



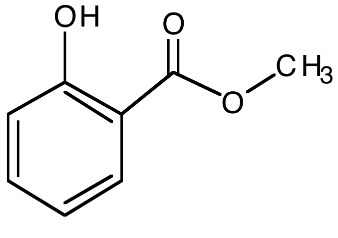
There are many familiar hydrocarbon species which are based on the benzene structure that contain various functional groups. A few representative examples are given below. Note that the delocalized pi electrons can be represented by a circle in the ring. This is an alternate representation of the ring often used by chemists.



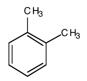
phenol

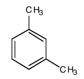


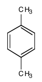
benzoic acid



oil of wintergreen



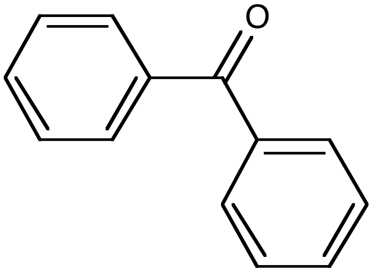




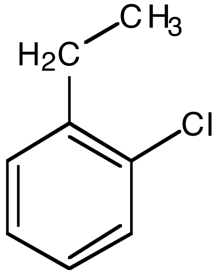
1,2-dimethylbenzene
(ortho-xylene)

1,3-dimethylbenzene
(meta-xylene)

1,4-dimethylbenzene
(para-xylene)

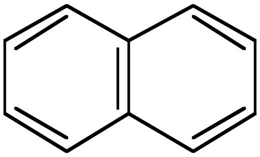


benzophenone

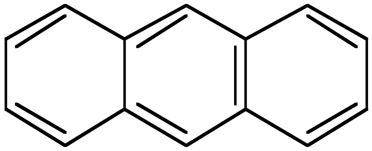


2-chloroethylbenzene

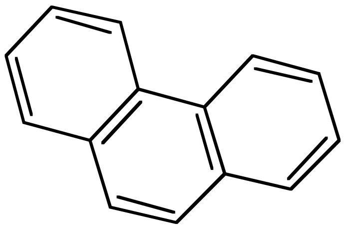
There are also many important aromatic hydrocarbons that contain more than one fused benzene ring. Several are illustrated here.



naphthalene



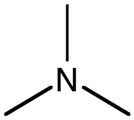
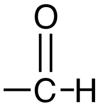
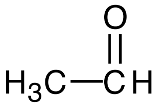
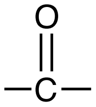
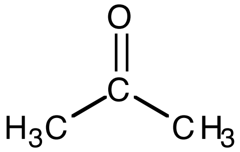
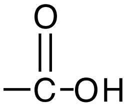
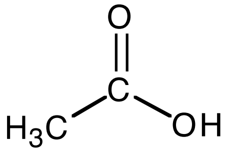
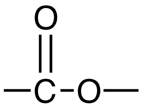
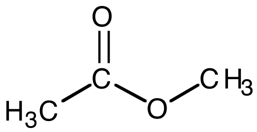
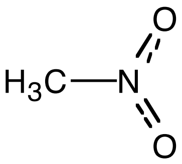
anthracene



phenanthrene

Hydrocarbons Containing Functional Groups

The basic types of hydrocarbon compounds outlined above may have one or more of their hydrogen atoms replaced by a **functional group**. The substituted benzenes earlier illustrated a number of functionalities ($-CH_3$, -OH, etc.) attached to the aromatic ring. Additional examples are shown in the table below.

Functional Group	Class of Compound	Example	Name
-OH	alcohol	H_3C-CH_2-OH	ethanol (ethyl alcohol)
-O-	ether	$H_3C-O-CH_3$	dimethyl ether
	amine	H_2N-CH_3	methylamine
 $-\overset{O}{\overset{\|}{C}}-H$	aldehyde	 $H_3C-\overset{O}{\overset{\|}{C}}H$	ethanal (acetaldehyde)
 $-\overset{O}{\overset{\|}{C}}-$	ketone		propanone (acetone)
 $-\overset{O}{\overset{\|}{C}}-OH$	carboxylic acid		ethanoic acid (acetic acid)
 $-\overset{O}{\overset{\|}{C}}-O-$	ester		methyl acetate
$-NO_2$	nitro		nitromethane
-X (X = F, Cl, Br, I)	haloalkane	H_3C-CH_2-Cl	chloroethane (ethyl chloride)

Name: ____

Organic Chemistry Worksheet - *Hand drawn versions only, no computer generated structures*

Answers to this lab should be provided in this worksheet. Neatness and attention to detail will count.

---

1. **Alkanes and Cycloalkanes**
   a. Write the structural formula and name for each straight-chain alkane with between one and eight carbons.

   b. What are the geometries, bond angles and hybridizations of the carbon atoms in the alkanes? What pattern do you notice for the relationship of carbon to hydrogen? (i.e. if an alkane has *n* carbons, how many hydrogen atoms are on the molecule?)

   c. Draw and name the five structural isomers of the hexane molecule. Give a suitable name to each of the hexane isomers. For example, one of the isomers is named 3-methylpentane. (Remember that one of the isomers is *n*-hexane!)

   d. Draw the structure for cyclohexane. What is the molecular formula of cyclohexane? What is the relationship between the number of carbon atoms and hydrogen atoms? Is cyclohexane an isomer of hexane? Explain.

2. **Alkenes and Alkynes**
   a. Draw the structure and molecular formula for ethene. What is the geometry and bond angle around the double bond? What is the hybridization of the carbon atoms? What is the relationship between the number of carbon atoms and hydrogen atoms?

   b. Construct four alkene isomers with four carbons each. Draw and name the structural formulas. Be sure to include *cis* and *trans* labels when appropriate.

   c. Draw the structure and molecular formula for ethyne. What is the geometry and bond angle around the triple bond? What is the hybridization of the carbon atoms? What is the relationship between the number of carbon atoms and hydrogen atoms? How many sigma and pi bonds are in ethyne?

   d. Construct models of the three five carbon alkyne isomers. Draw and name the isomers.

   e. Draw cyclobutyne. Using your knowledge of bond angles and hybridization, would you expect cyclobutyne to be a stable compound? Explain.

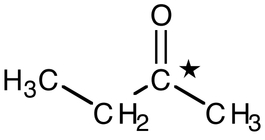
3. **Aromatic Compounds**
   a. Sketch the line structure of the aromatic compound benzene, C_6H_6. Comment on the fact that measurement of the C-C bond distances in benzene shows that all are the same length (1.397 Angstroms) and that the C-C bond angles are all 120°. Is benzene a flat molecule?

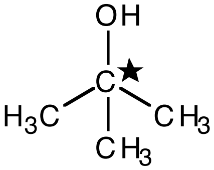
   b. Benzene was at one time used extensively as a solvent. However, since it has been found to be carcinogenic, it has largely been replaced by toluene. Toluene, which is formally known as methylbenzene, contains an aromatic ring substituted with a methyl group and has a molecular formula of C_7H_8. Sketch a possible structure for toluene. Do you think toluene would have a higher or lower boiling point than benzene? Explain.

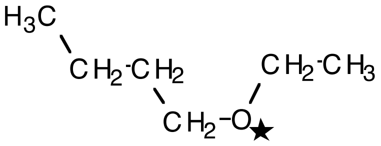
   c. Draw two resonance structures for naphthalene. (*Hint*: look in this lab for the naphthalene structure!)

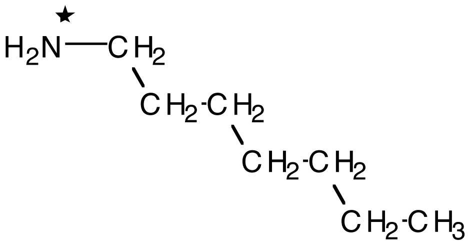
   d. There are three dichlorobenzene isomers. Draw and name them. Which one would you expect not to have a dipole moment? Explain.

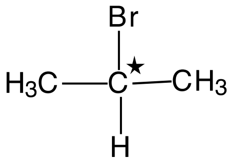
4. Identify the functional group and/or class of compound in each of the following molecules. List the hybridization around the marked atom in each compound. Finally, name the compound.



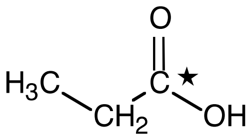








$$H_3C-\overset{\overset{Br}{|}}{\underset{\underset{H}{|}}{\overset{\star}{C}}}-CH_3$$



$$H_3C-CH_2-C\equiv\overset{\star}{C}-CH_3$$

5. Draw structures for acetone, formaldehyde and acetic acid. What are the bond angles around the C=O carbon atom in each of these compounds?

---

6. Draw the structures for the following compounds:

a. 2,3-difluorohexane

b. dimethyl ether

c. butanone

d. dimethylamine

e. pentan-2-ol

f. 2,3-diethyltoluene

g. *cis,trans*-octa-2,6-diene

h. 2-hexyne

i. 3-nitrotoluene

j. 4-bromobenzoic acid

k. 2,4,6-trinitrotoluene

l. 2-chlorophenol

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CH 222 Winter 2025:
"Molar Mass of a Volatile Liquid *(online)*" ***Lab*** - *Instructions*

Note: **This is the lab for section W1 of CH 222 only.**

- *If you are taking section 01 or section H1 of CH 222*, please use this link: http://mhchem.org/r/5a.htm

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Step One:

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

Step Two:

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

Step Three:

Submit your lab (pages Ib-5-5 through Ib-5-9 *only* to avoid a point penalty**) as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, February 12 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!

Molar Mass of a Volatile Liquid

The experimental determination of the molar mass of gases and vapors is one important application of the Ideal Gas Law. To determine the molar mass of a gas or vapor, we need to determine the mass (g, below) of the gas sample under known conditions of temperature (**T**), Volume (**V**) and pressure (**P**). Assuming the gas obeys the Ideal Gas Law:

$$\mathbf{PV = nRT}$$

and if pressure, temperature and volume are determined experimentally and **R** is a constant (**0.082057 L atm mol^{-1} K^{-1}**), we can solve the equation for moles (**n**) of gas. The molar mass (grams/mole, or **MM**) of the gas or vapor can be determined based on the mass of the liquid remaining (**g**) in the flask and the number of moles calculated from the ideal gas law; or you can use:

$$\mathbf{MM = gRT/PV}$$

In this experiment, the molar mass of a volatile liquid will be determined using the ideal gas law. A small amount of the liquid is introduced into a weighed flask. The flask is then placed in boiling water where the liquid will completely vaporize, driving out any air and filling the flask with vapor at barometric pressure and the temperature of the boiling water. The mass of the vapor can be determined by cooling the flask to condense the vapor. (Note: ***the thermometer does not need a separate clamp*** (as shown in Figure One, below); you can place the thermometer directly in the boiling water when a reading is required.)

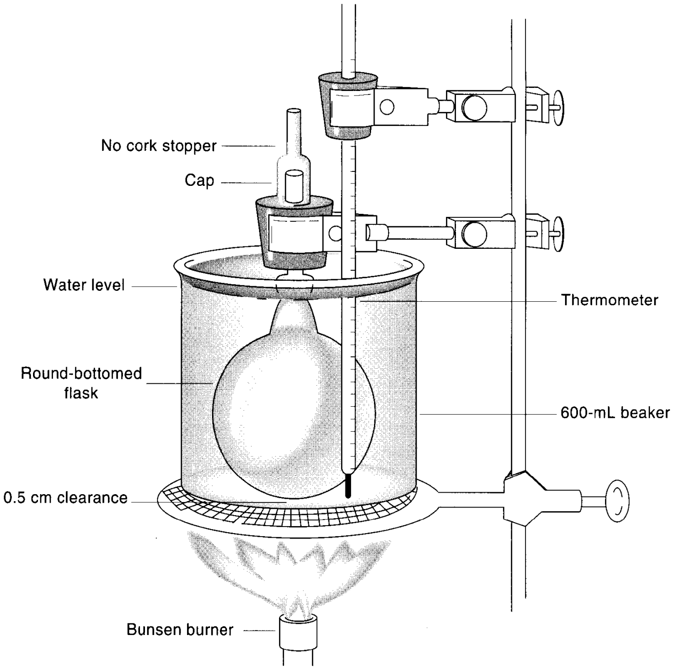

Figure One

PROCEDURE:
Obtain a round bottom flask fitted with a glass cap with a capillary opening, and a cork ring should support the entire apparatus at all times. If necessary, wash the flask with a touch of acetone, then dry the flask and cap and record the mass of the empty, dry flask and cap to the nearest 0.001g using the cork ring for support.

Obtain an unknown liquid and record the identification number. Pour approximately 5 mL of your unknown sample into the flask. (Note: the exact volume is not important; the mass of the final condensed vapor will be determined.) Assemble the apparatus as shown in Figure 1. It is important to have the round bottom flask immersed as deeply as possible in the 1000 mL water bath for uniform heating. Add a few boiling chips to the water in the 1000 mL beaker and heat the water to the boiling point. Watch the liquid level in your flask; the level should gradually disappear as vapor expands and pushes first air from the flask then excess vapor. Once all the liquid has disappeared, continue to heat the flask for 5 minutes to allow the vapor to reach the temperature of the boiling water. Measure and record the temperature of the boiling water. Also record the barometric pressure.

Turn off the burner and wait for the water to stop boiling (about 30 seconds). Remove the flask from the beaker of water and allow the flask to cool to room temperature and the vapor to condense. (You can speed up this process by running tap water over the stoppered flask.) Dry the outside of the flask and weigh the flask with the condensed vapor.

Repeat the procedure using another 5 mL sample of your liquid unknown. **Do not empty or clean the flask between trials**; just pour another 5 mL into the flask

Once the two trials are complete (**and not before**!), find the volume of the flask. Fill the flask and cap with water. Determine the mass of the flask filled with water, then determine the temperature of the water used to fill the flask. Using the *Handbook of Chemistry and Physics*, determine the density of the water to at least six sig figs. If the *Handbook* is difficult to read, use this link as an alternative: http://mhchem.org/den

Rinse the flask with a touch of acetone, then blow dry. Return the flask to the cart.

Using the density of water and the mass of water in the flask, calculate the volume of the flask. Be sure to subtract the empty mass of the flask first! When converting to atmospheres, use **1013 mbar = 1 atm** (other conversions will not receive full credit), use **273.15** to convert to Kelvin temperatures, and **use the complete version of R** (listed on the previous page.)

Determine the **average molar mass** of your liquid sample and **parts per thousand**.

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Molar Mass of a Volatile Liquid

YOUR NAME: ____

DATA: *Watch the video (***http://mhchem.org/y/5.htm***) to get these values using the data at the very end:*

empty flask (g): ____________________ Density of water (g/mL): ____________________

flask filled with water (g): ____________________

Trial I		**Trial II**	
Boiling water temperature (°C):	____________________	Boiling water temperature (°C):	____________________
flask + condensed vapor (g):	____________________	flask + condensed vapor (g):	____________________
barometric pressure (mbar):	____________________	barometric pressure (mbar):	____________________

Part A Calculations: *Molar Mass Determination of a Volatile Liquid*

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

1. Using the data from the video, determine the temperature of the gas in Kelvin for both Trials.

Trial I Temperature (K): ______________

Trial II Temperature (K): ______________

2. Using the data from the video, determine the pressure of the gas in atmospheres for both Trials.

Trial I Pressure (atm): ______________

Trial II Pressure (atm): ______________

3. Using the data from the video, determine the mass of the unknown liquid remaining at the end of the experiment for both Trials.

Trial I mass of liquid (g): ______________

Trial II mass of liquid (g): ______________

4. Using the data from the video, calculate the volume (L) that the gas occupies in the flask (*hint:* use only the **mass of water** in the flask; this will be the same for both Trial I and Trial II.)

Volume of flask (L): ______________

5. What is the value of R for this lab? Report to five significant figures and include units.

Value of R with units: _______________________________

6. Using the previously recorded and calculated values, **determine the molar mass** of the gas **for Trial I and Trial II**. **Calculate** the **average molar mass** for both trials. **Determine** the **average deviation** and the **parts per thousand** using your molar mass values. *Hint:* do not average any values until you have calculated the molar mass values for the unknown liquid. *Show all work!*

molar mass (Trial I) (g/mol): ______________

molar mass (Trial II) (g/mol): ______________

average molar mass (g/mol): ______________

average deviation: ______________

Parts per thousand: ________________

Postlab Questions:

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

1. Use the data below to find the moles of unknown and the molar mass of the unknown. *Show all work!*
    - mass of an empty flask and stopper = 55.441 g
    - 5 mL of unknown added and heated; boiling water bath at 100.˚C, all unknown liquid vaporized
    - mass of the flask, stopper and condensed vapor = 56.039 g.
    - volume of the flask = 215.9 mL
    - barometric pressure = 1003 mbar

moles unknown: _____________ **Molar mass of unknown (g/mol)**: __________

2. Determine if each of the following procedural errors would **increase**, **decrease**, or have **no effect** on the molar mass calculations in this experiment. **Explain** your reasoning.

   i. The flask was not dried before the final weighing with the condensed vapor inside.

      Effect on molar mass:

      Explain:

   ii. The flask was removed from the boiling water before the vapor had reached the temperature of the boiling water. All the liquid had vaporized.

      Effect on molar mass:

      Explain:

   iii. The volume of the cap was not measured with the volume of the flask.

      Effect on molar mass:

      Explain:

   iv. The student uses 7.5 mL of liquid sample.

      Effect on molar mass:

      Explain:

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CH 222 Winter 2025: **"Linear Regression & The Crystal Structures of Solids** *(online)"* ***Lab*** *- Instructions*

Note: **This is the lab for section W1 of CH 222 only.**

- *If you are taking section 01 or section H1 of CH 222*, please use this link:
  http://mhchem.org/r/6a.htm

Step One:

Watch the lab video for the "Linear Regression / Crystals" lab, found here: **http://mhchem.org/y/6.htm**
Record the data found at the ***end*** of the lab video on page Ib-6-12.

Step Two:

Complete pages Ib-6-11 through Ib-6-16 using the "Linear Regression / Crystals" video and the actual lab instructions on pages Ib-6-2 through Ib-6-9.

Step Three:

Submit your lab (pages Ib-6-11 through Ib-6-16 *only* with computer generated graphs to avoid a point penalty) **as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, February 19 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. Do ***not*** include graphs as separate file(s).... have all documents in one file to avoid a point penalty.

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

Linear Regression & The Crystal Structures of Solids

This lab will demonstrate the power of a linear regression analysis while graphing linear data. We will also explore the techniques used to analyze several common types of metals in crystallography.

An **equation** is a mathematical model used to describe the relationship between variables. We will focus on **linear** equations in this lab which use a horizontal (X) axis (the **independent** variable, the variable we input when we make a measurement) and a vertical (Y) axis (the **dependent** variable, the number we measure after we set the X value.) If the plotted data points form a straight line, this means we have a **linear equation**, and we can use: $\mathbf{y = mx + b}$ where **y** is the vertical axis value, **x** is the horizontal axis value, **m** is the slope of the line, and **b** is the y-intercept.

Computer programs and calculators can perform a **linear regression** analysis by plotting the "best fit" line through the data and then writing the slope-intercept equation. The **correlation coefficient** (with the symbol "**r**") is a measure of how well the regression line fits with the observed data. A ***perfect*** fit produces a correlation coefficient of either +1.000 (positive slope) or -1.000 (negative slope), depending on if the line slopes up (a positive slope) or down (a negative slope.) The closer the correlation coefficient is to +/-1.000, the better the regression line expresses the data (the better fit for the data.) Note that R^2 is slightly different than r!

Observing the **crystals** of an ordinary substance (such as table salt) using a magnifying glass, one sees many planes at right angles within the solid. This occurs in many common solids, and the regularity we see implies a deeper regularity in the arrangement of atoms or ions in the solid. The atomic nuclei are present in remarkably symmetrical arrays that continue for millions of units in three dimensions. Substances having a regular arrangement of atom-size particles in the solid are called **crystalline**, and the solid material consists of **crystals**. This lab deals with some of the simpler arrays in which atoms or ions occur in crystals and what these arrays can tell us about such properties as atomic sizes, densities of solids, and the efficiency of packing of particles.

Procedural Notes for the Linear Regression Lab: *Complete the problems using the worksheets at the end*

Each linear regression problem will require one or more **computer generated graphs** that will be stapled to the back of the worksheets. Recommended programs to graph your date include **Microsoft Excel** (**free** for MHCC students; see https://www.mhcc.edu/OfficeInstall/), **Apple Numbers** (free with a Mac computer, https://www.apple.com/numbers/), or **Google Sheets** (https://www.google.com/sheets). ***Note*** that Excel Online (the online version of Excel) and iPad/iPhone/Droid versions of these programs will generally not perform linear regressions, so try to use the "full" computer version instead. Calculators will perform linear regressions, but printing from a calculator might be difficult.

Use a **X-Y scatter plot** when graphing these data sets. **The computer program will analyze the data and perform the linear regression analysis** for you. Each program is different, but generally the user selects the actual data points on the X-Y Scatter plot and either right-clicks or control-clicks on the data to see a new menu…. you wish to "Add a Trendline" and "Display the R^2 value". If an equation appears with an R^2 value, you have performed your linear regression. ***Help*** on performing the linear regression lab can be acquired in the Learning Success Center / AVID Center at MHCC. You can also search YouTube for videos (i.e. search "linear regression Excel 2019" and almost inevitably a helpful video appears.)

Converting R^2 to r is not difficult. Take the square root of R^2 to get r. If the value of the slope is a negative number, then the value of r will also be negative…. watch for this in this lab! Calculators will often give both R^2 and r values when linear regression techniques are applied. Information on making an acceptable graph in this class can be found here: https://mhchem.org/lab

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The Linear Regression Problems:

<u>Problem 1</u>: The Relationship Between Celsius and Fahrenheit

In 1724, the German scientist Gabriel Fahrenheit developed a temperature scale based on phenomenon he thought could be easily repeated in laboratories around the world. For his zero degree point, Fahrenheit chose the coldest mixture of ice, water, and salt that he could produce in his laboratory. For ninety-six degrees, he chose what he believed to be normal body temperature. Fahrenheit wanted a temperature scale that could be divided into twelfths. On this scale, pure water freezes at 32 degrees, and pure water boils at sea level at 212 degrees.

A few years later, in 1742, the Swedish scientist Anders Celsius developed a different temperature scale. This scale used pure water as its standard. Zero degrees was the temperature where pure water froze, and one hundred degrees was the temperature where pure water boiled at sea level. Because Celsius had one hundred degrees between the two reference points on his temperature scale, it was called the *centigrade* scale. Recently this was renamed the Celsius scale in honor of Anders Celsius.

A student measures the following data points in the laboratory using two thermometers:

Temperature (°C)	20.0	40.0	60.0	80.0	100.0
Temperature (°F)	67.6	104.8	141.1	175.0	211.1

1. Construct and print a graph of degrees Fahrenheit (y) as a function of temperature in degrees Celsius (x).
2. Using your calculator, determine the mathematical equation of °F as a function of °C as well as the correlation coefficient, r. Record r to at least four significant figures.
3. Using the actual equation: °F = 1.8°C + 32 *and* your experimental equation, convert 29.0 °C to °F. Calculate **percent error** = (difference / actual value) x 100% Comment on discrepancies.

<u>Problem 2</u>: Solubility of Lead(II) Nitrate in Water

The solubility of lead(II) nitrate in water was measured as a function of temperature. The solubility is given in units of grams of lead(II) nitrate per 100 grams of water.

Temperature (°C)	20.0	40.0	60.0	80.0	100.0
Solubility (g / 100 g water)	56.9	74.5	93.4	114.1	131.1

1. Graph and print the data; temperature will be the independent (x) variable.
2. Determine the equation of the best-fit line. Record the equation and correlation coefficient.
3. What is the solubility of lead(II) nitrate at 47.0 °C?

Problem 3: Colorimetry

The colors in the visible spectrum of light are shown by a rainbow. Colored substances absorb segments of the visible spectrum of light. Pink solutions, for example, are pink because they absorb green light and transmit all other colors of the visible spectrum. If light of the particular color absorbed is passed through a sample, the amount of light absorbed will be related to the number of absorbing molecules in the light beam. Dilute solutions absorb little light, concentrated solutions absorb more. Typically the amount of light transmitted through the solution is measured; *transmittance* is inversely proportional to *absorbance*. The following data was obtained for the transmittance of 525 nm light by solutions containing different concentrations of permanganate ion.

Concentration (mg/100 mL)	1.00	2.00	3.00	4.00
Transmittance *(unitless)*	0.418	0.149	0.058	0.0260

1. **Convert the Transmittance values to Absorbance** using the following equation: **A = log (1/T)**, where A = Absorbance and T = Transmittance. **Use 3 sig figs for your absorbance values**.
2. Graph and print the Absorbance (y) versus Concentration (x) data. Perform a linear regression analysis. Record the equation and the correlation coefficient.
3. Predict the absorbance of 2.50 mg permanganate ion / 100 mL solution.

Problem 4: Kinetics

The branch of chemistry that studies the rate or speed of reactions is called *kinetics*. One must often plot concentration versus time data in a variety of mathematical formats to find a linear relationship; this assists in finding the *order of reaction*. We shall explore this topic more in CH 222. The following data was collected at 25.6 °C while measuring the disappearance of NH_3:

Concentration [NH_3] (mol/L)	$8.00 * 10^{-7}$	$6.75 * 10^{-7}$	$5.84 * 10^{-7}$	$5.15 * 10^{-7}$
Time (h)	0	25.0	50.0	75.0

1. **Prepare a graph of ln [NH_3] versus time** (time is the x-axis). "ln" stands for natural logarithm which can be calculated easily on your calculator (for example, the value of $8.00 * 10^{-7}$ is -14.039.) Perform a linear regression analysis on the ln [NH_3] versus time data and find the equation and the correlation coefficient. **Use at least 4 sig figs for your absorbance values**.
2. **Prepare a graph of 1 / [NH_3] versus time** (for example, $1 / 8.00 * 10^{-7}$ is $1.25 * 10^{6}$. ***Note:*** You may have to enter the data as "1.25E6" to make the program understand your values.) Time will be the x-axis. Perform a linear regression analysis and find the correlation coefficient and the values for the slope and the y-intercept.
3. **Which graph gives a *better* linear regression?** Why? *Hint:* look for the better correlation coefficient.
4. Plots of ln [NH_3] versus time that are linear are called *first order reactions* while graphs of 1 / [NH_3] versus time that are linear are called *second order reactions*. **What order of reaction** does the decomposition of NH_3 follow? *Hint:* the better linear regression will determine the order of the reaction!

Procedural Notes for the Crystal Structures of Solids Lab: Complete the handouts found at the end of this lab and turn it in (with all relevant work displayed on adjacent pages) to your instructor. Help on completing the "Solids" questions can be acquired in the Learning Success Center / AVID Center at MHCC.

Many crystals are unbelievably complex, and we will limit ourselves to the simplest crystals that have cubic structures. Cubic structures imply 90° angles and sides of equal length (hence, a cube.) We will also limit ourselves to the study of only one kind of system (namely metal elements), yet they will exhibit many of the interesting properties of more complicated structures.

The Simple Cubic (SC) Crystal

The simple cubic unit cell is a cube with an **edge length**, **d_0**, equal to the distance from the center of one atom to the center of the next (see Figure One). The volume of the cube is equal to $(d_0)^3$, expressed as

$$\mathbf{V = (d_0)^3}$$

and is very small since d0 is on the order of 0.5 nm. Using x-ray diffraction we can measure the value of d0 easily to four significant figures. The number of atoms in a simple cubic unit cell is equal to one, for only 1/8 of each corner atom is actually inside the cell.

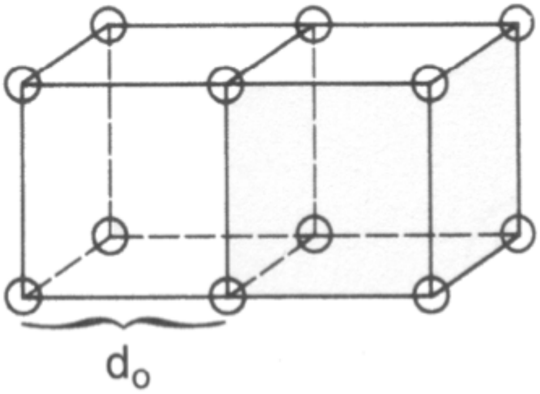


Figure One: *The Simple Cubic Crystal*

Each atom in the simple cubic unit cell is actually connected to six other atoms in the cubic lattice; hence, we say that the **coordination number** of the atoms in this structure is equal to six.

Many diagrams displaying the simple cubic unit cell show a gap between adjoining atoms. **In an actual crystal, we consider that the atoms that are closest are touching**. It is on this assumption that we determine **atomic radii, r**. In the SC crystal, if we know d_0, we can find the radius r of the atoms, since one side contains 2 atomic radii, or

$$\mathbf{d_0 = 2r}$$

for simple cubic crystals. Knowing the radius, we can calculate d0, and then we can calculate the volume of the unit cell. Knowing that one atom occupies the simple cubic cell, we can calculate the mass of the unit cell (using the molar mass and Avogadro's number), and from this we can determine the density using the volume of the cell.

Essentially no elements crystallize in the simple cubic structure, however, due to the inefficiency of the packing. The atoms in the simple cubic crystal are farther apart then they need to be, and inspection of the SC lattice will reveal a large hole in the center of the unit cell. Only about 52% of the cell volume is occupied by atoms, and more "empty space" means less stabilization for the crystal structure.

The Body Centered Cubic (BCC) Crystal

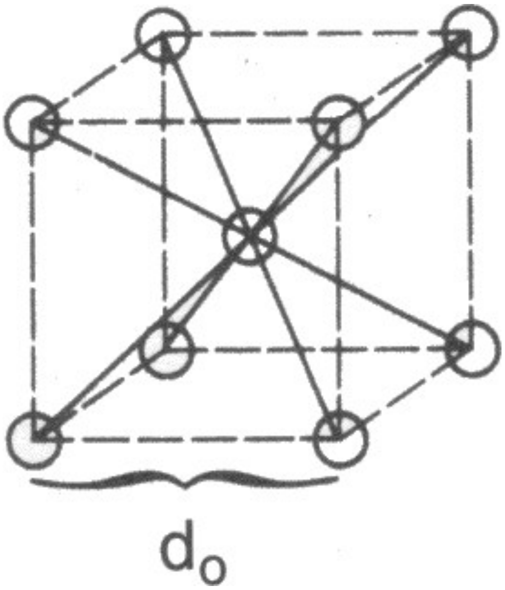


Figure Two: *Body Centered Cubic Crystal*

In a **body centered cubic crystal**, the unit cell still contains the corner atoms present in the SC structure, but the center of the cell now contains an additional atom. This means that every BCC crystal structure holds **two net atoms** (eight atoms are $^1/_8$ within the cell, and one whole atom within the center of the cell for two net atoms).

The edge length, d_0, can be determined using simple geometry from the **cube diagonal** (see Figure Two). The cube diagonal reaches across the cube, from an atom in the lower left front to an atom in the upper right back, or from any other appropriate combination. Geometry dictates the following relationship between the cube body diagonal and the edge length, d_0:

$$\text{cube diagonal} = \sqrt{3} \cdot d_0$$

The cube diagonal encompasses 4 radii lengths, and d_0 can be expressed in terms of the radius of the atom:

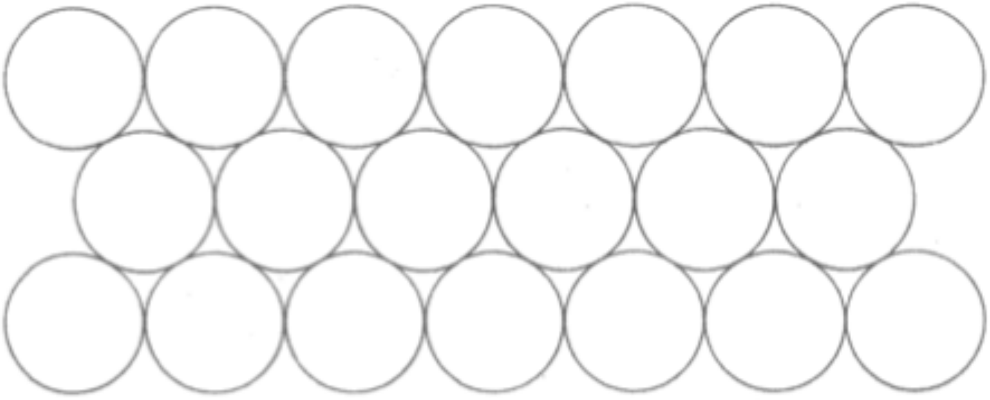
$$d_0 = \frac{4\text{r}}{\sqrt{3}}$$

The quantity d_0 can be used to find the volume of the cube; this is important for BCC cubic systems.

In a BCC lattice, each atom touches eight other atoms, and the **coordination number** is eight. The BCC lattice is much more stable than the SC structure, in part due to the higher coordination number. Many metals at room temperature display the BCC lattice, including sodium, chromium, tungsten and iron. Note that there are two atoms per unit cell in the BCC crystal. **BCC crystals are more efficient than SC crystals**, occupying approximately 68% of the total available volume.

Close Packed Structures

Although many elements prefer the BCC crystal arrangement, still more prefer structures in which the atoms are **close packed**. In close packed structures there are layers of atoms in which each atom is in contact with six others, as in the sketch below:



This is the way in which billiard balls lie in a rack or the honeycomb cells are arranged in a bees' nest. It is the most efficient way one can pack spheres, with about 74% of the volume in a close packed structure filled with atoms.

There is more than one way whereby close packed crystal structures can be stacked. One of the stacking methods is cubic and is called the **Face Centered Cubic (FCC)**. The other is called **Hexagonal Close-Packing**. We shall look at both close packed structures.

The Face Centered Cubic (FCC) Crystal

In the face centered cubic crystal unit cell there are atoms in each corner of the cell (as in the SC cell discussed earlier) and there is another atom at the center of each of the six faces. This means that FCC cubic systems consist of **four net atoms** per unit cell (eight atoms are $^1/_8$ within the cell, and six faces hold an atom which is $^1/_2$ within the cell for four net atoms). See Figure Three.

The edge length d_0 can be determined in an FCC crystal from the **face diagonal** which is defined as the distance across one face of the cube. Using geometry, we can find the edge length from the face diagonal using the following equation:

$$\text{face diagonal} = \sqrt{2} \cdot d_0$$

The face diagonal encompasses 4 radii lengths, and d_0 can be expressed in terms of the radius r:

$$d_0 = \frac{4\text{r}}{\sqrt{2}}$$

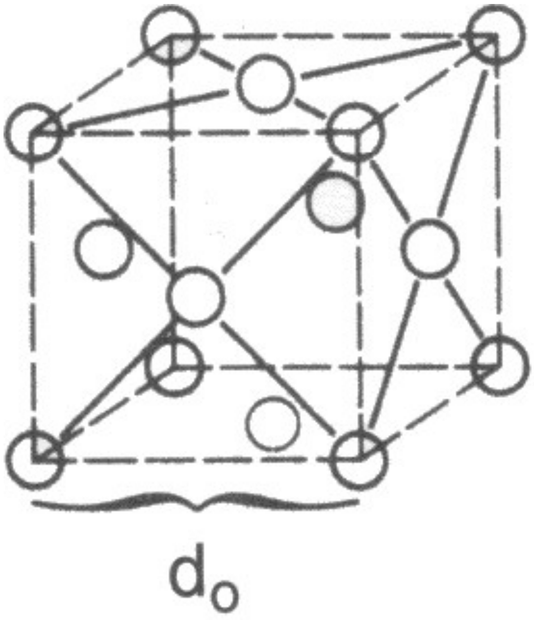


Figure Three: *Face Centered Cubic Crystal*

This expression can be used to find the volume of the cube; hence, this relationship is important for FCC cubic systems. The **coordination number** in an FCC lattice is 12, implying that FCC lattices are quite stable.

The close-packed layers of atoms in the FCC lattice are not parallel to the unit cell faces, but rather are perpendicular to the cell diagonal. If you look down the cell diagonal, you see six atoms in a close-packed triangle in the layer immediately behind the corner atom, and another layer of close-packed atoms below that, followed by another corner atom. The layers are indeed closely packed, and as one goes down the diagonal of this and succeeding cells, the layers repeat their positions in the order ABCABC…. This implies that atoms in every fourth layer lie below one another (see Figure Four (b)).

Hexagonal Close-Packing

There is another way to stack the layers as in the FCC lattice, above. The first and second layers will always be in the same relative positions, but the third layer could be below the first one if it were shifted properly. This results in a **close-packed structure** in which the order of the layers is ABABAB… (see Figure Four (a))

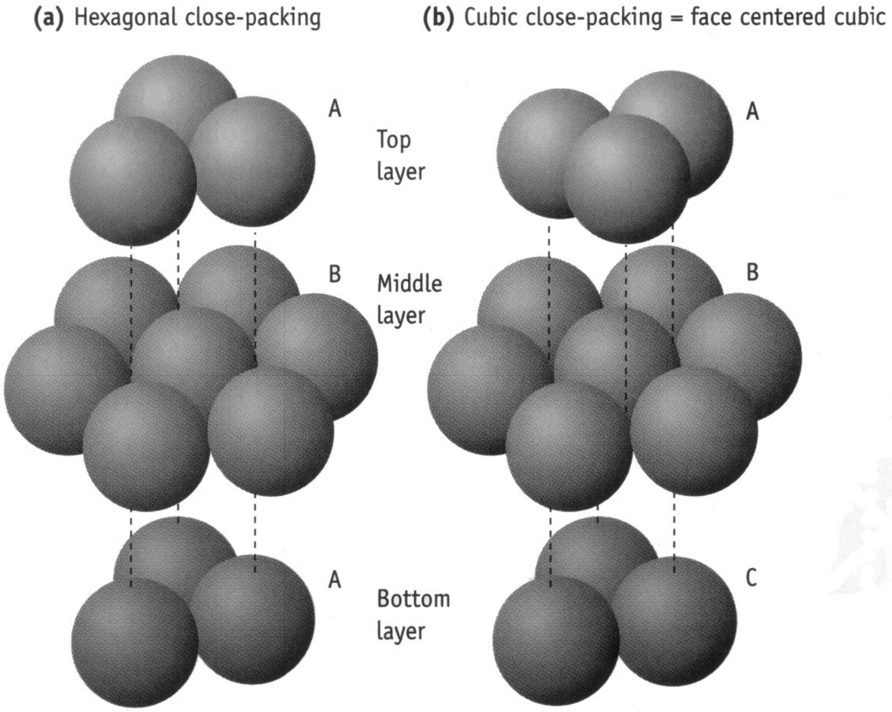


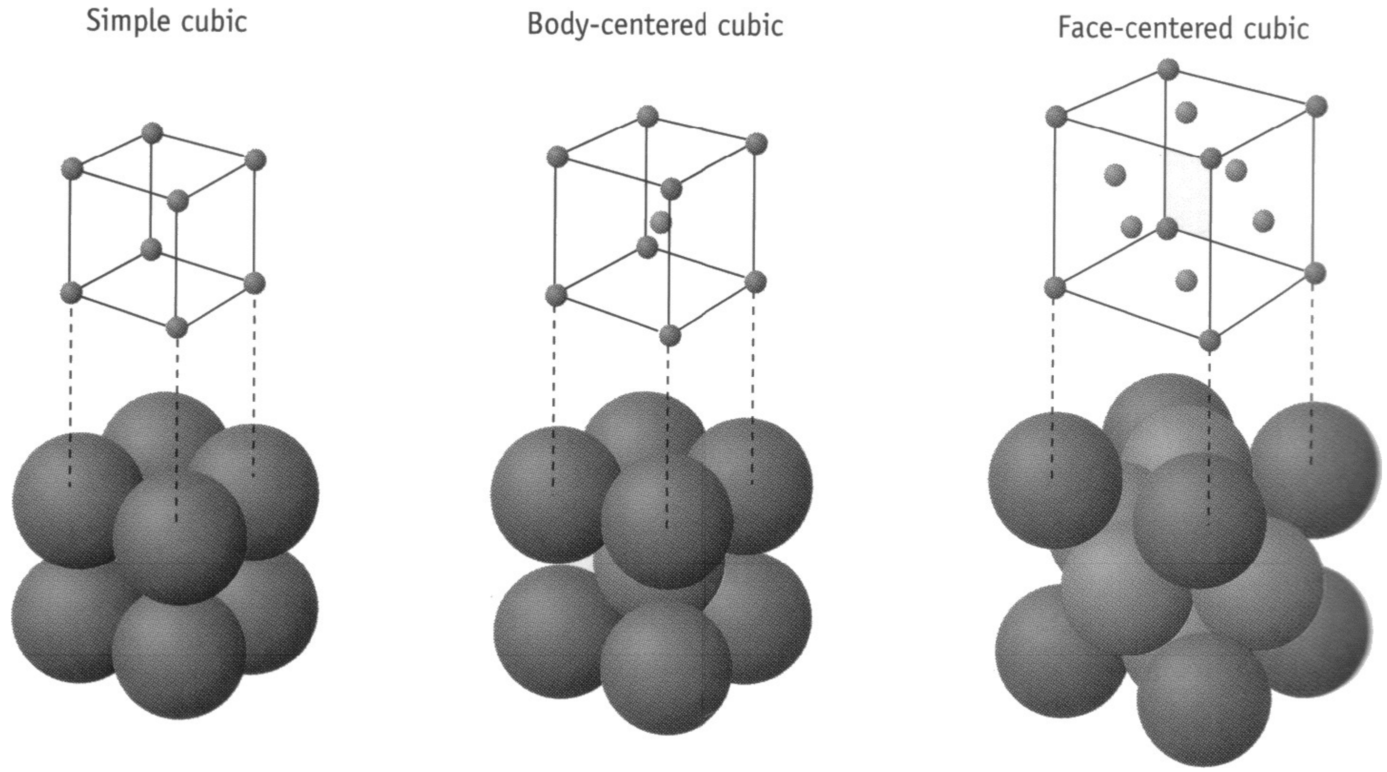
Figure Four: *Hexagonal Close Packing (left) and Cubic Close Packing (right)*

The crystal obtained from this arrangement of layers is not cubic but **hexagonal**. It is another common structure for metals. Cadmium, zinc and manganese have this structure. As you might expect, the stability of this structure is very similar to that of FCC crystals. We find that simply changing the temperature often converts a metal from one form to another. Calcium, for example, is FCC at room temperature, but if heated to 450 °C it converts to close-packed hexagonal.

In CH 222 (and CH 223), we will consider "hexagonal close-packing" structures to be identical to FCC lattices, but technically there are many differences between the two systems.

Summary of Crystal Lattice Types

Figure Five shows the three main cubic unit crystal types that we will explore in this lab. **Figure Six** show a helpful methodology to solve problems like these in this lab.



Lattice Type	**Simple Cubic**	**Body Centered Cubic**	**Face Centered Cubic**
# net atoms per cell	1	2	4
d_0 (edge) in relation to r	$d_0 = 2r$	$d_0 = \frac{4r}{\sqrt{3}}$	$d_0 = \frac{4r}{\sqrt{2}}$

Figure Five: *Summary of the Three Cubic Unit Cell Types*

conversion *V = edge³* *density* *molar mass (g/mol)* *Avogadro (6.022×10^{23})*

$$\text{radius} \leftrightarrow \text{edge} \leftrightarrow \text{volume} \leftrightarrow \text{mass (g)} \leftrightarrow \text{moles} \leftrightarrow \text{atoms / molecules}$$

1 pm = 10^{-12} m / 1 Å = 10^{-10} m / 1 cm = 10^{-2} m *4 atoms = 1 fcc cell, etc.*

Figure Six: *Helpful Conversion Methodology*

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Worksheet: Linear Regression & The Crystal Structures of Solids *Name:*

All final answers must be provided on this worksheet. **Include computer generated graphs** (from Excel or a similar program) **along with any necessary calculations** needed for the lab in your submission. Values of r should be recorded to three or more significant figures. **This page *will* be placed at the front of your lab, with your first and last name included, to prevent a point penalty.**

Linear Regression:

- **Problem 1**: ***The Relationship Between °C and °F*** *- see page Ib-6-3 for data and questions*

Linear Regression equation: **y =** ____

r = ______________ *Percent Error:* ______________________________

- **Problem 2**: ***Solubility of Lead(II) Nitrate in Water*** *- see page* Ib-6-3 *for data and questions*

Linear Regression equation: **y =** ____

r = ______________ *Solubility of lead(II) nitrate at 47.0 °C:* __________________

- **Problem 3**: ***Colorimetry*** *- see page* Ib-6-4 *for data and questions*

Linear Regression equation: **y =** ____ **r =** ____________

Absorbance of 2.50 mg permanganate in 100 mL solution: ______________________________

Linear Regression: *Continued*

- **Problem 4:** ***Kinetics*** *- see page* Ib-6-4 *for data and questions*

Linear Regression (ln [NH_3] vs. time) equation: **y =** ________________________________ **r =** __________

Linear Regression (1/[NH_3] vs. time) equation: **y =** ________________________________ **r =** __________

Which regression gives a better linear regression? Why?

Does this data behave as a first order reaction or a second order reaction?

- **Problem 5 *(Use the data from the video)*:** Experimentally determine the density of an unknown metal solid to at least three significant figures using the displacement method. Report and use the data from the lab video.

Relevant calculations and data:

The Crystal Structures of Solids:

- **Problem 6:** What element forms a face centered cubic cell, has a density of 8.92 g/cm^3, and a radius of 128 pm?

Element = __________ *Show relevant work below*

The Crystal Structures of Solids: *Continued*

- **Problem 7:** Chromium forms a body centered cubic crystal. If the length of an edge is 2.884 angstroms, calculate the **density** (g/cm^3) and the **radius** of a chromium atom in angstroms.

density (g/cm^3) = __________________ **radius (Å) = ___________________**

Show relevant work below.

The Crystal Structures of Solids: *Continued*

- **Problem 8:** Sodium (radius = 186 pm) forms a body centered cubic crystal. Calculate the **density** (g/cm^3) of sodium metal. **Propose a simple experiment to confirm your calculated density of sodium** in the lab.

density (g/cm^3) = _________________

Show relevant work below.

Proposed simple experiment:

The Crystal Structures of Solids: *Continued*

- **Problem 9:** Aluminum crystallizes in a face centered cubic unit cell. In addition, aluminum has an atomic radius of 143 pm. What is the density of aluminum?

density (g/cm^3) = __________________

Show relevant work below.

CH 222 Winter 2025: **“Molar Mass Determination by Freezing Point Depression** *(online)” **Lab** - Instructions*

Note: **This is the lab for section W1 of CH 222 only.**

- *If you are taking section 01 or section H1 of CH 222*, please use this link:
  http://mhchem.org/r/7a.htm

Step One:

Watch the lab video for the “Freezing Point” lab, found here:
http://mhchem.org/y/7.htm

Step Two:

Complete pages Ib-7-7 through Ib-7-10 using the “Freezing Point” video and the actual lab instructions on pages Ib-7-2 through Ib-7-5. Include your name on page Ib-7-7!

Step Three:

Submit your lab (pages Ib-7-7 through Ib-7-10 *only* to avoid a point penalty) **with a computer generated graph as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, February 26 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. Do ***not*** include graphs as separate file(s).... have all documents in one file to avoid a point penalty.

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

Molar Mass Determination by Freezing Point Depression

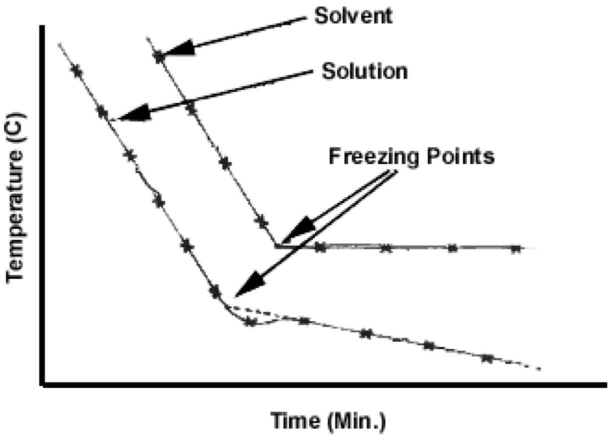
If a nonvolatile solute is added to a liquid, a number of physical properties of the pure substance change, including vapor pressure depression, freezing point depression, and boiling point elevation. These alterations are collectively known as **colligative properties** of solutions. The colligative properties of a solution change in proportion to the concentration of solute dissolved in solvent and depend only on the number of solute particles present in a given amount of solvent and not on the type of particles dissolved. Therefore, the concentration of the solute is most conveniently expressed in terms of **molality** (*m*) or **moles of solute / kg of solvent**.

The change in the freezing point (ΔT_f) in ˚C for a nonvolatile organic solvent can be determined using the following equation, where $\mathbf{k_f}$ is characteristic for the solvent used: $\mathbf{\Delta T_f = k_f \cdot} m$

We can determine the **molar mass** of the solute using this equation by measuring the change in the freezing point of the solution and solving the equation for molality. The calculated molality can be used to determine the moles of solute that in turn can be used to calculate the molar mass (grams / mole) of the solute.

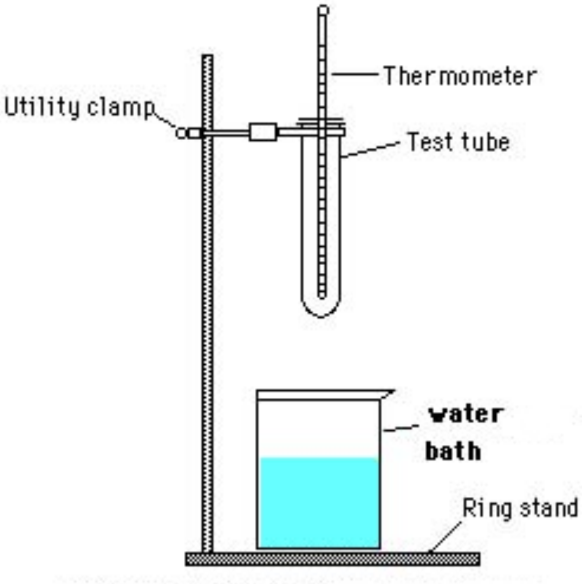
When a pure substance is heated to a liquid state and allowed to cool, initially the temperature will fall quite rapidly. As the substance approaches its **freezing point**, solid will begin to form and the temperature will begin to level. The freezing point of the pure liquid is the constant temperature observed while the liquid is solidifying.

The cooling behavior of a solution is somewhat different from that of a pure liquid. The temperature at which the solution begins to freeze is **lower** (i.e. *depressed*) than for the pure solvent. Additionally, there is a slow gradual fall in temperature as freezing proceeds. The **change in temperature, ΔT**, between the freezing point of the pure substance and the freezing point of the solution is used to calculate the molality of the solution.



In both the pure liquid and the solution, a **supercooling** effect may be seen. As the solid begins to form, the temperature may drop *below* the actual freezing point initially and then come back up to the freezing point temperature as the solid forms. Supercooling is usually not observed if adequate churning of the sample is provided. When determining the freezing point, the super-cooling effect should be ignored.

In this experiment, you will first determine the freezing point of a pure solvent, **lauric acid** ($C_{12}H_{24}O_2$). Next, you will use a known solute, **benzoic acid**, to depress the freezing point of the solvent and calculate the molar mass of the benzoic acid.



This picture should also include a hot plate under the beaker and a "swizzle stick" around the thermometer

PROCEDURE: Part A: Determining the Freezing Point for Lauric Acid

1. Set up ring stand and test tube clamp beside the hot plate
2. Fill a large beaker about ¾ full with tap water and heat on hot plate.
3. Mass 10.000-12.000 g of solid lauric acid ($C_{12}H_{24}O_2$) and record the mass. Add the lauric acid to a large test tube.
4. Place the test tube in the beaker and heat until the lauric acid is completely melted. Put a "swizzle stick" around a thermometer and place it into the liquid. Gently stir with the swizzle stick (not the thermometer!) to uniformity. Do not overheat - you will need to cool it in step 5 (perhaps read ahead?) You may need to readjust your thermometer so it remains in the solution.
5. Remove the test tube from the hot water and allow it to cool. Once the temperature reaches 50 °C, begin to record the temperature every 30 seconds until the temperature plateaus for 3 minutes or reaches 35 °C. (This should take 5-10 minutes). Mix ***gently*** to maintain uniformity (test tubes are glass! Careful!) The biggest source of error in this lab is the temperature so use care in reading. Make sure you record the temperature to the correct number of significant figures – if unsure, ask!

Note: Be sure that data is recorded in the lab data books for all lab partners (perhaps the note keeper could record in both note books as the other student mixes?)

PROCEDURE: Part B: Determination of the Molar Mass for Benzoic Acid

6. Mass 1.000 - 1.200 g of benzoic acid and record the mass. Add to the test tube with the lauric acid. Be careful not to lose any sample of lauric acid that may have solidified to the thermometer or swizzle stick.

7. Place test tube in hot water and melt the mixture, stirring to uniformity.

8. Remove the test tube from the hot water and allow it to cool. Once the temperature reaches 50 °C, begin to record the temperature every 30 seconds until the temperature plateaus for 3 minutes or reaches 35 °C. (This should take 5-10 minutes). Mix ***gently*** to maintain uniformity (test tubes are glass! Careful!)

9. Repeat steps 6 through 8, adding an additional 1.000 - 1.200 grams of benzoic acid to the solution. Do ***NOT*** make a new solution!

10. Time to clean up! To clean up, re-melt your solid and pour into waste container. You may need to add a tiny bit of hot water along the sides of the tube to help melt the solid. Wipe the thermometer and swizzle stick with a wee bit of acetone and/or methanol if needed.

CALCULATIONS: Determining the Molar Mass of the Benzoic Acid

Once you have completed your experiment, you will need to create a **graph in Excel** (or a similar program; no hand drawn graphs will be accepted.) Use **time** as the x-axis and **temperature** as the y-axis. Plot the data from your pure lauric acid experiment (Part A) and also the data from each of the benzoic acid solutions (Part B) **using the same graph**. Label the axes and use different colors and/or marking symbols to distinguish the trials.

Important: your graph should take up an entire page of paper (no small graphs, print in "landscape" mode), *and* the **y-axis should *not* start at zero** (i.e. if your data points end around 25 °C, have the minimum y-axis value be 20 °C, not zero °C (which is what these programs often default to when creating graphs.)

Determine the freezing point for each trial graphically; **draw a circle** on your graph representing the freezing point in each trial. Remember, the freezing points can be found at the intersection between the "steep" cooling curve and the "gradual" cooling curve.

The k_f value for lauric acid is 3.90 °C/m. Using the data from Part A and Part B, calculate the molar mass of benzoic acid in each trial. Show how you calculated these values in your lab report.

Average the two molar mass values and calculate the **parts per thousand**.

Using the Internet or a textbook, determine the structure of benzoic acid and **include a Lewis structure for benzoic acid in your lab report**. (Hand drawn Lewis structures are ok.)

Determine the actual molar mass of benzoic acid using the Lewis structure. **Calculate the percent error** of benzoic acid using your average molar mass value and the accepted molar mass value. *Recall:* **Percent error = absolute value{(actual - experimental)/ actual}*100%.**

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Molar Mass Determination by Freezing Point Depression

YOUR NAME: ____

DATA: *Watch the video (***http://mhchem.org/y/7.htm***) for assistance with this lab.*

Graph the following data in Excel, Numbers, Sheets or a similar computer program (hand drawn graphs will not be accepted for credit.) Use **Time (minutes)** as your x-axis and **Temperature °C** as your y-axis.

- You will have two different data streams ("**PDB**" and "**Solution I**" which use a common set of axes; use a color or point marking system to differentiate "PDB" from "Solution I".
- Ensure that your temperature axis reflects an appropriate range of values for the data provided (i.e. don't start at zero!) Attach the graph to this lab when turning it in to the instructor.

Time (minutes)	PDB (deg C)	Solution I (deg C)
0.5	63.7	61.3
1.0	62.1	59.7
1.5	60.8	58.2
2.0	59.1	56.5
2.5	57.5	54.8
3.0	55.8	53.4
3.5	54.0	52.0
4.0	53.0	50.8
4.5	53.0	50.2
5.0	53.0	50.2
5.5	53.0	50.2
6.0	53.0	50.1
6.5	53.0	50.0
7.0	53.0	50.0
7.5	53.0	49.9
8.0	53.0	49.8
8.5	53.0	49.7

ANALYSIS:

- On your graph, determine the freezing point for the "PDB" and "Solution I" curves. Look for the spot where two slopes on the same line come together; this is the official freezing point. **Mark the graph at the freezing point**, and enter the actual numbers below.

Freezing Point of PDB (°C): ________________ **Freezing Point of "Solution I" (°C):** ________________

- The "PDB" for your graph stands for "para-dichlorobenzene". Draw the structure of PDB here.

- Using the information from the graph and the freezing points, **determine the molar mass of the solute** in "Solution I" if **2.35 g** of the unknown compound were dissolved in **30.46 g** of PDB. (**k_{fp}** for PDB = **7.10 °C/m**) *Show work!*

POSTLAB QUESTIONS:

1. Determine the effect of the following on the final molar mass calculation in a freezing point depression experiment. *Briefly* give your reasoning. Indicate if the effect on molar mass will be **higher**, **lower** or **not change**.

   i. The thermometer you were using read temperatures consistently 1.2 ˚C higher than the real temperature.

   Effect on molar mass:

   Explain:

   ii. You knowingly added 3 g of unknown solute during your freezing point determination instead of 2 g.

   Effect on molar mass:

   Explain:

   iii. In adding your unknown solute to your solvent, you spilled some solute onto the table after weighing.

   Effect on molar mass:

   Explain:

2. A student performs a freezing point analysis. She determines that the freezing point of 21.00 g of stearic acid (where $k_f = 4.89°C/m$) is 68.20 ˚C. She adds 2.07 grams of an unknown compound to her sample and determines the freezing point to be 65.53 ˚C. She adds an additional 1.97 g of the unknown compound and determines the new freezing point to be 63.03 ˚C.

    a. Determine the molar mass of the unknown compound using the 2.07 g of sample.

    b. Determine the molar mass of the unknown compound using the *combined* samples (*hint:* 2.07 + 1.97 = 4.04 g total solute)

    c. Determine the average molar mass of the compound and the parts per thousand for the two trials.

CH 222 Winter 2025:
“Kinetics I - The Iodination Of Acetone *(online)*” ***Lab*** *- Instructions*

Note: **This is the lab for section W1 of CH 222 only.**

- *If you are taking section 01 or section H1 of CH 222*, please use this link: http://mhchem.org/r/8a.htm

Step One:

Watch the lab video for the “Kinetics I” lab, found here: **http://mhchem.org/y/8.htm**
Record the data found at the ***end*** of the lab video on pages Ib-8-4 and Ib-8-5.

Step Two:

Complete pages Ib-8-5 through Ib-8-15 using the “Kinetics I” video and the actual lab instructions on pages Ib-8-2 through Ib-8-3 (most of this lab is a “tutorial”, so you will also find instructions in the lab pages as well.) Include your name on page Ib-8-5!

Step Three:

Submit your lab (pages Ib-8-5 through Ib-8-15 *only* to avoid a point penalty**) as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, March 5 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.

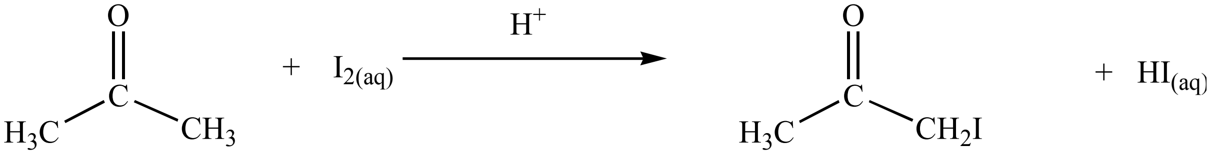
If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Kinetics I - The Iodination Of Acetone

Determining the Rate Constant for a Chemical Reaction

The **rate** of a chemical reaction depends on several factors: the nature of the reaction, the concentrations of the reactants, the temperature, and the presence of a possible catalyst. In *Part One* of this experiment we will determine the **rate law** for a reaction by changing some of the above variables and measuring the rate of the reaction. During *Part Two*, we will explore the relation between the rate constant and temperature to discover the **activation energy** for this reaction.

In this experiment, we will study the kinetics of the reaction between iodine and acetone:



The rate of this reaction is found to depend on the concentration of the hydrogen ion (acid, HCl) as well as the concentrations of the reactants (acetone and iodine). The **rate law** for this reaction is

$$\mathbf{rate = k[acetone]^m[H^+]^n[I_2]^p}$$

where **k** is the **rate constant** for the reaction and **m**, **n**, and **p** are the **orders of the reaction** with respect to acetone, hydrogen ions (acid), and iodine, respectively. Although orders of reaction can be any value, for this lab we will be looking only for **integer values** for the orders of reaction (0, 1, 2 are acceptable but not 0.5, 1.3, etc.)

The **rate** of the reaction can also be expressed as the change in the concentration of a reactant divided by the time interval:

$$rate = \frac{-\Delta[I_2]}{\Delta t}$$

The iodination of acetone is easily investigated because iodine (I_2) has a deep yellow/brown color. As the acetone is iodinated and the iodine converted to the iodide anion, this color will disappear, allowing the rate of the reaction to be easily monitored.

We can study the rate of this reaction by simply making I_2 the ***limiting reactant*** in a large excess of acetone and H^+ ion. By measuring the time required for the initial concentration of iodine (I_2) to be used up completely, the rate of the reaction can be determined by the equation

$$rate = \frac{-\Delta[I_2]}{\Delta t} = \frac{-([I_2]_{final} - [I_2]_{initial})}{t_{final} - t_{initial}} = \frac{-(0 - [I_2]_{initial})}{t_{final} - 0} = \frac{[I_2]_{initial}}{t_{final}}$$

or simply as

$$rate = \frac{[I_2]}{time}$$

From the rate information, we can determine the **orders** with respect to acetone (**m**), acid (**n**) and iodine (**p**) by varying the amounts of reactants and measuring the effect on the rate. Once the orders of reaction are known, we will be able to calculate the rate constant, k. In Part One of this experiment you will determine the rates of reactions, the orders of the reactants, and finally the rate constant at room temperature.

DIRECTIONS:

You will collect data during lab and complete the worksheets at the end. No formal typed lab report is required as long as your writing is legible and easy to read. You will perform a series of experiments that will examine the relationship between the concentration of reactants and the time for the iodination of acetone reaction.

For each trial listed below: measure out the appropriate quantities of **1.0 M HCl, 4.0 M acetone** and **water** using a 10.00 mL graduated cylinder and place them in a 125 mL Erlenmeyer flask. Now measure out the appropriate amount of **0.0050 M iodine** in a 10.00 mL graduated cylinder.

Start a timer (stopwatch) as you add the iodine to the 125 mL flask with the other chemicals. Swirl the flask **until the yellow color disappears**, **then halt the timer**. It may help to place the flask on a white piece of paper to help discern when the color disappears. Record the time elapsed in seconds.

Repeat this reaction mixture until **two** trials are within 20 seconds of each other.

Repeat this process for each of the four trials listed in the table below. Waste can be placed in the drain or in a waste bottle (probably the better option!) Get a stamp in your lab notebook before leaving lab, then complete the worksheet portions on your own.

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Name:

Kinetics I - The Iodination Of Acetone

Part I: Changing Concentration to Find the Rate Constant

For each mixture listed below, add all of the chemicals but iodine to a 125 mL Erlenmeyer flask. Add the iodine last, starting a stopwatch and measuring how long the reaction takes to turn the solution clear. Time should be recorded in seconds. **Repeat** each reaction mixture until **two** trials are within 20 seconds of each other.... repeat the trial again if the times vary too much. ***NOTE:*** **Record the data found at the end of the video in the places below for Trials #1 - #4.**

Trial #1:

HCl (mL)	Acetone (mL)	I_2 (mL)	Water (mL)	Total Volume (mL)
5	5	5	10	25.00

Time in seconds for yellow color to disappear, first time: ____________________ seconds

Time in seconds for yellow color to disappear, second time: ____________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ____________________ seconds

Trial #2:

HCl (mL)	Acetone (mL)	I_2 (mL)	Water (mL)	Total Volume (mL)
5	5	10	5	25.00

Time in seconds for yellow color to disappear, first time: ____________________ seconds

Time in seconds for yellow color to disappear, second time: ____________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ____________________ seconds

Trial #3:

HCl (mL)	Acetone (mL)	I_2 (mL)	Water (mL)	Total Volume (mL)
5	10	5	5	25.00

Time in seconds for yellow color to disappear, first time: ____________________ seconds

Time in seconds for yellow color to disappear, second time: ____________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ____________________ seconds

Trial #4:

HCl (mL)	Acetone (mL)	I_2 (mL)	Water (mL)	Total Volume (mL)
10	5	5	5	25.00

Time in seconds for yellow color to disappear, first time: ____________________ seconds

Time in seconds for yellow color to disappear, second time: ____________________ seconds

Time in seconds for yellow color to disappear, third time (if necessary): ____________________ seconds

Before you move on, check your data. Trial #1 should be about half the time of trial #2, and trial #3 and trial #4 should both be about the same time elapsed. If you don't see this trend, contact the instructor.

You are now ready to complete the Kinetics I lab!

Kinetics I - The Iodination Of Acetone – *Worksheet*

The ultimate goal of this section is to find the best value of the rate constant, k, for the iodination of acetone at room temperature. To get there, we need to first find the rates of each reaction, then the order of the reactants (acetone, HCl and iodine), and then finally, the rate constant k.

a. **Find the average time in seconds** for each Trial in Part One.

Example: the first experiment took 230 seconds, the second experiment took 250 seconds.

The *average time* would be: (230 + 250) / 2 = **240 seconds**

Record your reaction times from Part I (which were within 20 seconds of each other) and the final average times here:

	Experiment #1 (s)	Experiment #2 (s)	Average time (s)
Trial #1			
Trial #2			
Trial #3			
Trial #4			

Show an example of how you got the average values in seconds here:

b. **Find the concentration of each reactant** (acetone, HCl and iodine) before the reaction started.

Your group took bulk reactants (which were, as a reminder: **1.0 M HCl, 4.0 M acetone**, and **0.0050 M iodine**) then placed them (with water) in an Erlenmeyer flask. The final volume was always 25.00 mL. Mixing chemicals dilutes the concentrations from the "bulk" value to a smaller value.

We can determine these diluted values using: $\mathbf{M_1V_1 = M_2V_2}$

Example: Determine the concentration of iodine in trial #1.

Let $\mathbf{M_1}$ = initial (undiluted) concentration of iodine (0.0050 M), $\mathbf{V_1}$ = 5.00 mL (of undiluted iodine added to the mixture), and $\mathbf{V_2}$ = 25.00 mL (the total volume of the diluted solution once HCl, acetone and water are added). Solving for M_2, the concentration of iodine in the diluted solution, one gets:

> $\mathbf{M_2}$ = 0.0050 M * 5.00 mL / 25.00 mL = **0.0010 M**, which is the concentration of iodine used in the reaction in trial #1.

Example: Determine the concentration of iodine in trial #2.

Solution: Since $\mathbf{M_1}$ and $\mathbf{V_2}$ are the same as in the previous example, we see that only $\mathbf{V_1}$ has changed to 10.00 mL. Rearranging for $\mathbf{M_2}$ as before:

> $\mathbf{M_2}$ = 0.0050 M * 10.00 mL / 25.00 mL = **0.0020 M**, the concentration of iodine used in the reaction in trial #2.

i. **Calculate the concentration of acetone ($\mathbf{M_2}$) used in Trial 1 - Trial 4.**

The **concentration** of acetone was 4.0 M in the "bulk" solution (the "$\mathbf{M_1}$" value.)
The final volume ($\mathbf{V_2}$) is always 25.00 mL.
Trial 1, Trial 2 and Trial 4 used 5.00 mL of acetone from the bulk solution, but Trial 3 used 10.00 mL (your "$\mathbf{V_1}$" values)

	volume acetone (mL)	**Bulk Acetone (M)**	**Total Volume (mL)**	**acetone (M)**
Trial #1	5.00	4.0 M	25.00	__________
Trial #2	5.00	4.0 M	25.00	__________
Trial #3	10.00	4.0 M	25.00	__________
Trial #4	5.00	4.0 M	25.00	__________

ii. **Calculate the concentration of HCl (M_2) used in Trial 1 - Trial 4.**

The **concentration** of HCl was 1.0 M in the "bulk" solution (the "$\mathbf{M_1}$" value.)
The final volume ($\mathbf{V_2}$) is always 25.00 mL.
Trial 1, Trial 2 and Trial 3 used 5.00 mL of acetone from the bulk solution, but Trial 4 used 10.00 mL (your "$\mathbf{V_1}$" values)

	volume HCl (mL)	**Bulk HCl (M)**	**Total Volume (mL)**	**HCl (M)**
Trial #1	5.00	1.0 M	25.00	________
Trial #2	5.00	1.0 M	25.00	________
Trial #3	5.00	1.0 M	25.00	________
Trial #4	10.00	1.0 M	25.00	________

iii. **Complete the following table** showing the diluted concentrations of all reactants used in each trial. The I_2 concentrations have been completed for you (see the examples at the beginning of part b, above.)

	acetone (M) *step i, above*	**HCl (M)** *step ii, above*	**I_2 (M)**
Trial #1	________	________	0.0010
Trial #2	________	________	0.0020
Trial #3	________	________	0.0010
Trial #4	________	________	0.0010

Use this area to show relevant calculations and at least one example as to how you found a diluted concentration.

c. Find the rate of each trial

In this lab, rate is best described by: **rate = $[I_2]$/(average time in seconds)** (see first page of lab)
Find the rate for each reaction by using the calculated iodine concentration then dividing by the average time for that trial (Part Three, section a, above).

Example: Trial #1 required an average of 240 seconds to go to completion. Trial #2 required an average of 496 seconds to go to completion. Find the rate of reaction for both trial #1 and trial #2.

Solution: The rate of reaction is equal to the concentration of iodine divided by the average time elapsed for the reaction. (your values are in Part Three, section a.)

In this example, the first two times are 240 s and 496 s. Using the appropriate concentrations, we can calculate rate:

rate (trial #1) = $[I_2]$/(average time in seconds) = 0.0010 M / 240 s = **$4.2 * 10^{-6}$ M s^{-1}**.

rate (trial #2) = $[I_2]$/(average time in seconds) = 0.0020 M / 496 s = **$4.0 * 10^{-6}$ M s^{-1}**.

Now, **calculate your rate values** by filling in the table below:

	I_2 (M)	**average time (s)**	**rate (M s^{-1})**
Trial #1	0.0010	__________	__________
Trial #2	0.0020	__________	__________
Trial #3	0.0010	__________	__________
Trial #4	0.0010	__________	__________

Use this space to show at least one example of how you calculated the rate of the reaction.

d. **Find the order** of the reaction with respect to each reactant (**m**, **n**, and **p**)

To find the order of each reactant, we need to localize the effect that each reactant had on the rate. This is expressed in a **rate order** which, for CH 222 and CH 223, **can only be equal to zero, one or two** (no fractions, negative numbers, etc.)

In this lab, it is safe to assume that **water does not affect the rate**, so we can ignore its contribution to any changes that the rate might exhibit.

Notice that **trial 2 has twice as much iodine as trial 1**. Any changes to the rates of trial 1 and trial 2 are a direct result of the iodine (HCl and acetone are constant), and we will use these trials to calculate **p, the reaction order for iodine.**

Also notice that **trial 3 has twice as much acetone as trial 1**, so any changes to the rates of trials 3 and 1 will be the result of acetone only (HCl and iodine are constant), and we will use these trials to calculate **m, the reaction order for acetone.**

Lastly, notice that **trial 4 has twice as much HCl as trial 1**, yet acetone and iodine concentrations remain constant; we will use these trials to calculate **n, the reaction order for HCl.**

Example: Find the order of reaction with respect to iodine (p) if these values are used:

rate (trial #1) = $[I_2]$/(average time in seconds) = 0.0010 M / 240 s = $\mathbf{4.2 * 10^{-6}\ M\ s^{-1}}$.

rate (trial #2) = $[I_2]$/(average time in seconds) = 0.0020 M / 496 s = $\mathbf{4.0 * 10^{-6}\ M\ s^{-1}}$.

Solution: Notice how in trial #2 we doubled the concentration of $[I_2]$ while leaving the other reactants' concentrations (HCl, acetone) unchanged. An effect upon the rate of the reaction will reflect the influence of the iodine only, and this will allow us to determine **p**.

In these sample calculations, doubling the concentration of iodine (to 0.0020 M from 0.0010 M) spawned a negligible change in the rate ($4.0 * 10^{-6}$ M s^{-1} versus $4.2 * 10^{-6}$ M s^{-1}). Because we are only concerned with whole integer values of rate orders, this implies a **zero order reactant**, and **p = 0**.

A more formal approach to determining p would be as follows: divide the expression for rate 2 by the expression for rate 1; this results in the following:

$$\frac{\text{rate 2}}{\text{rate 1}} = \frac{k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p}{k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p}$$

The values of k, [acetone] and [HCl] remain constant between trial 1 and trial 2 (only $[I_2]$ changes), so the expression reduces to

$$\frac{\text{rate 2}}{\text{rate 1}} = \frac{4.0 * 10^{-6}}{4.2 * 10^{-6}} = \frac{[0.0020]^p}{[0.0010]^p} = 2^p$$

$$0.95 = (2)^p$$

Taking the logarithm of both sides leads to

$$\log 0.95 = \log 2^p = p \log 2$$

and solving for p:

$$p = \frac{\log 0.95}{\log 2} = -0.074 \approx 0$$

Therefore, the order with respect to iodine equals **zero**, or **p = 0**.

Now, **find the order of each reactant** by filling in the tables below. ***Note:*** **rate** values from Part 3 section c; **concentration** (M) values from Part 3, section b, subsection iii. Show your work on the following page. ***Remember*** that m, n and p can be **0, 1, or 2 only**! Round your answers as necessary!

For I_2 (p):

	I_2 (M)	**rate (M s^{-1})**	**My value of p is:**
Trial #1	0.0010	__________	__________
Trial #2	0.0020	__________	

For acetone (m):

	acetone (M)	**rate (M s^{-1})**	**My value of m is:**
Trial #1	__________	__________	__________
Trial #3	__________	__________	

For HCl (n):

	HCl (M)	**rate (M s^{-1})**	**My value of n is:**
Trial #1	__________	__________	__________
Trial #4	__________	__________	

Use this space to show **how** you got each of the orders of reaction (m, n and p):

e. **Find the value of k, the rate constant** for the iodination of acetone reaction.

You now have all the necessary information to calculate the rate constant, k, for each trial. For this reaction,

$$\text{rate} = k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p$$

Rate values appear in Part 3 section c
[acetone], [HCl] and [I2] are the concentrations for each trial (Part 3, section b, subsection iii)
m, n and p are the orders of reaction (Part 3, section d)

Example: Find the value of k for trial #1 if the rate = 4.2×10^{-6} M s^{-1} and assuming that the order with respect to acetone (m) and HCl (n) is two and the order with respect to I_2 (p) is zero.

Solution: In trial #1, the diluted concentration of acetone is 0.80 M, the HCl is 0.20 M and I_2 is 0.0010 M. Using the given values, we can calculate k from the following equation:

$$\text{rate} = k[\text{acetone}]^m[\text{HCl}]^n[I_2]^p$$

$$4.2 * 10^{-6} = k[0.80]^2[0.20]^2[0.0010]^0$$

$$k = 4.2 * 10^{-6} / [0.80]^2[0.20]^2$$

and solving for k we get a value of $\mathbf{k = 1.6 * 10^{-4}\ M^{-1}\ s^{-1}}$

Now **find the rate constant k for each trial** by completing the table:

My value of m = __________ *(these can be found in section d, above)*

My value of n = __________

My value of p = __________

	acetone (M)	**HCl (M)**	**I_2 (M)**	**rate ($M\ s^{-1}$)**	**value of k**
Trial #1	__________	__________	0.0010	__________	__________
Trial #2	__________	__________	0.0020	__________	__________
Trial #3	__________	__________	0.0010	__________	__________
Trial #4	__________	__________	0.0010	__________	__________

Concentrations in section b, subsection iii *Rate values in section c*

Average value of k: ______________________

parts per thousand of your four k values: ______________________

Use this space to show **a sample calculation for k** and also relevant **parts per thousand calculations**:

f. **You are done!** Finish the postlab questions (which are similar to the work you just completed) and you are good to go!

Kinetics I - The Iodination Of Acetone - Postlab Questions:

1. In a reaction involving the iodination of acetone, the following reaction mixture was used: 5.00 mL 4.0 M acetone, 5.00 mL 1.0 M HCl, 5.00 mL 0.0050 M I_2, and 10.0 mL water.

   a. What was the molarity of the acetone in the reaction mixture? (Recall that $M_1V_1 = M_2V_2$)

   b. The color of the above reaction mixture disappeared in 250 seconds. What was the rate of the reaction? (Hint: First determine the initial concentration of the iodine, then use the equation for rate from the lab.)

2. A second reaction mixture was made: 10.00 mL acetone, 5.00 mL HCl, 5.00 mL I_2, and 5.00 mL of H_2O.

   a. What was the molarity of the acetone in this reaction mixture?

   b. The iodine color disappeared in 120 seconds. What was the rate of the reaction?

   c. Determine the order of the reaction (m) with respect to acetone using the information from question 1 and 2. (Round off the value of m to the nearest integer)

3. A third reaction mixture is made: 10.00 mL acetone, 5.00 mL HCl, 10.00 mL I_2. If the reaction is zero order with respect to iodine, how long will it take for the iodine color to disappear? (*Hint:* rate = $[I_2]$/time, use the rate from question 2b and the new $[I_2]$ to solve for the time elapsed.)

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CH 222 Winter 2025: **"Kinetics II - The Iodination Of Acetone** *(online)"* ***Lab*** *- Instructions*

Note: **This is the lab for section W1 of CH 222 only.**

- *If you are taking section 01 or section H1 of CH 222*, please use this link:
  http://mhchem.org/r/9a.htm

Step One:

Watch the lab video for the "Kinetics II" lab, found here:
http://mhchem.org/y/9.htm
Record the data found at the ***end*** of the lab video on page Ib-9-5.

Step Two:

Complete pages Ib-9-5 through Ib-9-11 using the "Kinetics II" video and the actual lab instructions on pages Ib-9-2 through Ib-9-3 (most of this lab is a "tutorial", so you will also find instructions in the lab pages as well.) Include your name on page Ib-9-5!

Step Three:

Submit your lab (pages Ib-9-5 through Ib-9-11 *only* to avoid a point penalty) **as a *single* PDF file to the instructor via email (mike.russell@mhcc.edu) on Wednesday, March 12 by 11:59 PM. Be sure to include any necessary computer generated graphs as well.** I recommend a free program like CamScanner (https://camscanner.com) to convert your work to a PDF file. Do not include the graph as a separate file.... have all documents in one file to avoid a point penalty.

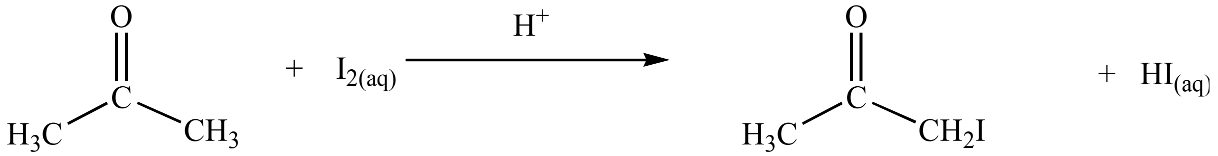
If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Kinetics II - The Iodination Of Acetone

Determining the Activation Energy for a Chemical Reaction

The **rate** of a chemical reaction depends on several factors: the nature of the reaction, the concentrations of the reactants, the temperature, and the presence of a possible catalyst. In *Part One* of this experiment we determined the **rate law** for a reaction by changing the concentrations and measuring the rate of the reaction. During *Part Two*, we will explore the relation between the rate constant and temperature to discover the **activation energy** for this reaction.

In this experiment, we continue to study the kinetics of the reaction between iodine and acetone:



In last week's lab, you discovered the **average rate constant, k**, value, as well as the **orders of reaction** (**m, n and p**) which apply to the **rate law** for this reaction:

$$\textbf{rate} = \textbf{k[acetone]}^{m}[\textbf{H}^{+}]^{n}[\textbf{I}_2]^{p}$$

In this experiment, you will study the rate of the reaction at different temperatures to find its **activation energy**, $\mathbf{E_a}$. The temperature at which the reaction occurs influences the rate of the reaction. An increase in temperature increases the rate.

As with concentration, there is a quantitative relation between reaction rate and temperature, but here the relation is somewhat more complicated. This relation is based on the idea that to react, the reactant species must have a certain minimum amount of energy present (and the correct geometry, if appropriate) at the time the reactants collide in the reaction step.

This amount of energy, which is typically furnished by the kinetic energy of the species present, is called the **activation energy** ($\mathbf{E_a}$, also known as the *energy of activation*) for the reaction. The formula (called the *Arrhenius equation*) relating the rate constant k to absolute Kelvin temperature T and E_a is:

$$\ln k = \frac{-E_a}{RT} + \ln A$$

In this equation, **R** is the gas constant (**8.3145 J/mole K**), and natural logarithms (**ln**) need to be used (*do not use base 10 logs!*) The quantity **A** is referred to as the **collision frequency** and A can be used to determine the fraction of molecules present with sufficient energy and geometry to become products at a given instant in time.

By measuring k at different temperatures, we can graphically determine the activation energy for a reaction. In Part Two of this experiment you will determine the effect of temperature on rate and calculate the activation energy for the reaction.

DIRECTIONS:

You will collect data during lab and complete the worksheets at the end. No formal typed lab report is required as long as your writing is legible and easy to read. In this experiment, you will see the effect of temperature upon the reaction.

We shall measure one of the reactions from the "Kinetics I" lab at different temperatures. You do not need to repeat these experiments twice (to be within 20 seconds of each other) as in Part One.

For each entry in trial #5 listed below: measure out the appropriate quantities of **1.0 M HCl, 4.0 M acetone** and **water** using a 10.00 mL graduated cylinder and place them in a 125 mL Erlenmeyer flask. Now measure out the appropriate amount of **0.0050 M iodine** in a 10.00 mL graduated cylinder.

Using ice and/or a hot plate, get the solution to a desired **temperature** before adding the iodine. **Record the temperature**, then add the iodine to the Erlenmeyer, start a stopwatch and measure how long the reaction takes to turn the solution clear. Time should be recorded in **seconds**. The iodine does not need to be at the same temperature as the solution in the Erlenmeyer flask.

Temperatures need to be higher than 15 °C (too slow!) and lower than 60 °C (keep the acetone from boiling), and the interval between measurements needs to be at least 5 °C apart.

Waste can be placed in the drain or in a waste bottle (which is probably the better option!) Get a stamp in your lab notebook before leaving lab.

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Name:

Kinetics II - The Iodination Of Acetone

As before, add all the chemicals but iodine to a 125 mL Erlenmeyer flask. Using ice and/or a hot plate, get the solution to a desired **temperature** before adding the iodine. **Record the temperature**, then add the iodine to the Erlenmeyer, start a stopwatch and measure how long the reaction takes to turn the solution clear. Time should be recorded in **seconds**. The iodine does not need to be at the same temperature as the solution in the Erlenmeyer flask.

* Record one trial at **room temperature**. *Hint:* you *may* be able to use some of your data from Part I!
* Record one trial at a **temperature lower than room temperature, but above 15 degrees Celsius**.
* Record **three trials at temperatures higher than room temperature, but under 60 degrees Celsius**.

Temperature differences should be at least 5 degrees Celsius apart (i.e. spread out your temperatures!)

***NOTE*: Record the data found at the end of the video in the places below:**

Trial #5:

HCl (mL)	**Acetone (mL)**	**I_2 (mL)**	**Water (mL)**	**Total Volume (mL)**
5	5	5	10	25.00

Temperature (°C): ____________________ Time (seconds): _______________

Temperature (°C): ____________________ Time (seconds): _______________

Temperature (°C): ____________________ Time (seconds): _______________

Temperature (°C): ____________________ Time (seconds): _______________

Temperature (°C): ____________________ Time (seconds): _______________

Hint: As temperature increases, the reaction time should decrease. If you do not see this trend, repeat one or more experiments.

You are now ready to complete the Kinetics II lab!

Kinetics II - The Iodination Of Acetone – *Worksheet*

Earlier you determined the time elapsed for a given set of concentrations as the temperature was altered. We shall use that information and techniques similar to that of the "Kinetics Part I" lab to determine the **energy of activation** and **collision frequency** for the iodination of acetone reaction using the Arrhenius equation.

a. **Find the inverse Kelvin temperature** for each value in Trial #5.

Convert your temperatures from °C to K, then take the *inverse* of your Kelvin temperatures.

Example: Convert 37.5 °C to an inverse Kelvin temperature.

Solution: 37.5 °C = 310.7 K. To find the inverse, calculate $(310.7\ \text{K})^{-1}$ = **$3.219 * 10^{-3}\ \text{K}^{-1}$**

Complete the table below. The first column (Temperature (°C) comes from your data collected in Part Two, Trial #5, while in lab.

	Temperature (°C)	**Temperature (K)**	**Temperature^{-1} (K^{-1})**
Temp #1	__________	__________	__________
Temp #2	__________	__________	__________
Temp #3	__________	__________	__________
Temp #4	__________	__________	__________
Temp #5	__________	__________	__________

Use this space to show **a sample calculation** for getting from Temperature (°C) to an inverse Kelvin temperature:

b. **Find the rate** for each temperature value in Trial #5.

Recall from the "Kinetics Part I" lab that, for this experiment:
rate = $[I_2]$/(time in seconds)

Use this equation to **find the rate of reaction** (in M s^{-1}) for each temperature. Time values come from Trial #5, above:

	I_2 (M)	**time (s)**	**rate (M s^{-1})**
Trial #1	0.0010	__________	__________
Trial #2	0.0010	__________	__________
Trial #3	0.0010	__________	__________
Trial #4	0.0010	__________	__________
Trial #5	0.0010	__________	__________

Use this space to show at least one example of how you calculated the rate of the reaction.

c. **Find the value of the rate constant, k,** for each temperature value in Trial #5.

We will use the process developed in "Kinetics Part I" lab to help us find the values of k for each temperature.

i. First, we need the **diluted concentrations**: *(these can be found in section b of the "Kinetics Part I" lab)*

Concentration (M) of **acetone** when 5.00 mL was used: _______________

Concentration (M) of **HCl** when 5.00 mL was used: _______________

Concentration (M) of **I_2** when 5.00 mL was used: _______________

ii. Next, we need the **reaction orders** for each reactant: *("Kinetics Part I", section d)*

My value of m (acetone) = __________

My value of n (HCl) = __________

My value of p (I_2) = __________

If you do not have m = 1, n = 1, p = 0, talk to the instructor!

iii. Now use the techniques from in Kinetics I", section e, to find the value of k, and then ln k.

Use the **rates from section b** *(above)* and the values for concentration and order (m, n and p) **to find k**. The only variable that will change is the rate; the orders and concentrations remain constant. Take the **natural log (ln) of each k** value as well (i.e. ln ($2.6 * 10^{-5}$) = **-10.56)** (Note: *report your ln k values to the **hundredths** place to satisfy significant figures.*)

	rate ($M\ s^{-1}$)	**k**	**ln k**
Trial #1	__________	__________	__________
Trial #2	__________	__________	__________
Trial #3	__________	__________	__________
Trial #4	__________	__________	__________
Trial #5	__________	__________	__________

Show a sample calculation for these steps on the next page.

Use this space to show at least one example of how you calculated the rate constant k and ln k.

d. **Create a graph** of ln k versus inverse Kelvin temperature values

You will be creating a graph of ln k versus inverse temperature to find the energy of activation.

First, collect your inverse temperature (the x-axis) and ln k values (the y-axis) here. Inverse Kelvin temperature values come from section a in this lab. ln k values come from section c, subsection iii, in this lab.

	Temperature^{-1} (K^{-1})	**ln k**
Trial #1	__________	__________
Trial #2	__________	__________
Trial #3	__________	__________
Trial #4	__________	__________
Trial #5	__________	__________

Using Excel or a similar program, create a **graph** of your ln k values versus the inverse Kelvin temperature values. Make the graph at least as big as half a sheet of paper, and be sure to include unit labels (ln k for the y-axis and (Temperature)$^{-1}$ for the x-axis. *Note* that when using your graphing program, you may need to enter values as decimals, i.e. enter 0.00315 instead of 3.15×10^{-3}.

Staple / attach the graph to the end of this lab report packet.

e. **Find the energy of activation** for the iodination of acetone using the data in Trial #5.

The data points on the graph from the last section should correspond roughly to a straight line with a negative slope. This is the behavior predicted by the Arrhenius equation:

$$\ln k = \frac{-E_a}{RT} + \ln A$$

where **ln k** is the y-axis, **(Temperature in Kelvin)$^{-1}$** is the x-axis, **-E_a/R** is the slope, **R = 8.3145 J K^{-1} mol^{-1}** (the "energy" gas constant), **E_a** is the energy of activation, and **A** is the collision frequency.

Perform a linear regression analysis using your calculator or graphing program (inverse Kelvin temperatures will be your x-axis, ln k values will be your y-axis.) Record the values that you collected here:

Slope = ___________ y-intercept = _________ correlation coefficient (r) = ____________

The **energy of activation**, E_a, can be determined from the slope. From the value of the slope determined through linear regression, **calculate the activation energy**.

Energy of activation = **–R*slope** = _____________

Units for your Energy of activation value = _____________

The collision frequency, A, can be determined from the y-intercept. From the value of the y-intercept determined through linear regression, **calculate the collision frequency**.

Collision frequency = **$e^{y\text{-int}}$** = _____________

Note that **e** is the anti natural logarithm.

f. **You are done!** Finish the postlab question (which is similar to the work you just completed) and you are good to go!

Kinetics II - The Iodination Of Acetone - Postlab Question:

The following reaction

$$2\ N_2O_{5(g)} \rightarrow 4\ NO_{2(g)} + O_{2(g)}$$

was studied at several temperatures, and the following values of k were obtained:

k (s^{-1})	**T (˚C)**
$2.0*10^{-5}$	20.0
$7.3*10^{-5}$	30.0
$2.7*10^{-4}$	40.0
$9.1*10^{-4}$	50.0
$2.9*10^{-3}$	60.0

Using linear regression and the techniques developed in this lab, calculate the **activation energy** and **collision frequency** for this reaction. **Include a computer generated graph** of ln k versus $(T)^{-1}$. *Hint:* make sure you use inverse Kelvin temperatures! Make sure the x-axis lists "0.003" numbers (and not whole integers, etc.)

Slope = ___________ y-intercept = _________ correlation coefficient (r) = ____________

Energy of activation = _____________

Units for your Energy of activation value = _____________

Collision frequency = _____________

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CH 222 Winter 2025:

Problem Set #1

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #1 by **reading Chapter 7** of the textbook and/or by watching the videos found on the website (https://mhchem.org/222)
- **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, January 13 at 1:10 PM**
- *Section H1:* due **Wednesday, January 15 at 1:10 PM**

Section W1: **Watch the recitation video** for Problem Set #1:

http://mhchem.org/y/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 #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, January 15.**

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

CH 222 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 Seven** *and* **Chapter Guide One**
*Important Tables and/or Constants: "***Bond Enthalpies and Electronegativities***" Handout* (after this problem set) and "**Geometry and Polarity Guide**" *(Handout,* https://mhchem.org/geopo*)*

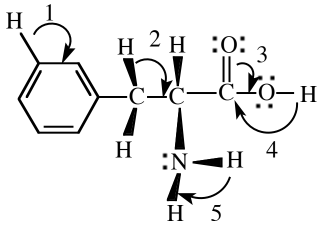
---

1. Give the periodic group number, the number of valence electrons and the number of core electrons for each of the following atoms.
   a. C
   b. Cl
   c. Ne
   d. Si
   e. Se
   f. Al
2. Which of the following elements are capable of forming compounds in which the indicated atom has more than four valence electron pairs?
   a. C
   b. P
   c. O
   d. F
   e. Cl
   f. B
   g. Se
   h. Sn

3. Which compound in each of the following pairs should require the higher temperature to melt?
   a. NaCl or RbCl
   b. BaO or MgO
   c. NaCl or MgS
4. Draw a Lewis structure for each of the following molecules or ions.
   a. CS_2
   b. BF_4^{-1}
   c. NO_2^{-1}
   d. $SOCl_2$
5. Draw a Lewis structure for each of the following molecules or ions.
   a. BrF_5
   b. IF_3
   c. IBr_2^{-1}
   d. BrF_2^{+1}
6. Draw a Lewis structure for each of the following molecules or ions. Describe the electron-pair geometry and the molecular geometry around the central atom.
   a. ClF_2^{+1} *(note: this is one Cl atom, 2 F atoms, and a +1 charge)*
   b. $SnCl_3^{-1}$
   c. BCl_3
   d. CS_2

Problem Set #1 continues on the next page

7. Draw a Lewis structure of each of the following molecules or ions. Describe the electron-pair geometry and the molecular geometry around the central atom.
    a. SiF_6^{2-}
    b. PF_5
    c. SF_4
    d. XeF_4
8. Give approximate values for the indicated bond angles.
    a. Cl-S-Cl in SCl_2
    b. N-N-O in N_2O
    c. Bond angles 1 - 5 in phenylalanine (*right*), one of the natural amino acids and a "breakdown" product of aspartame.



Phenylalanine

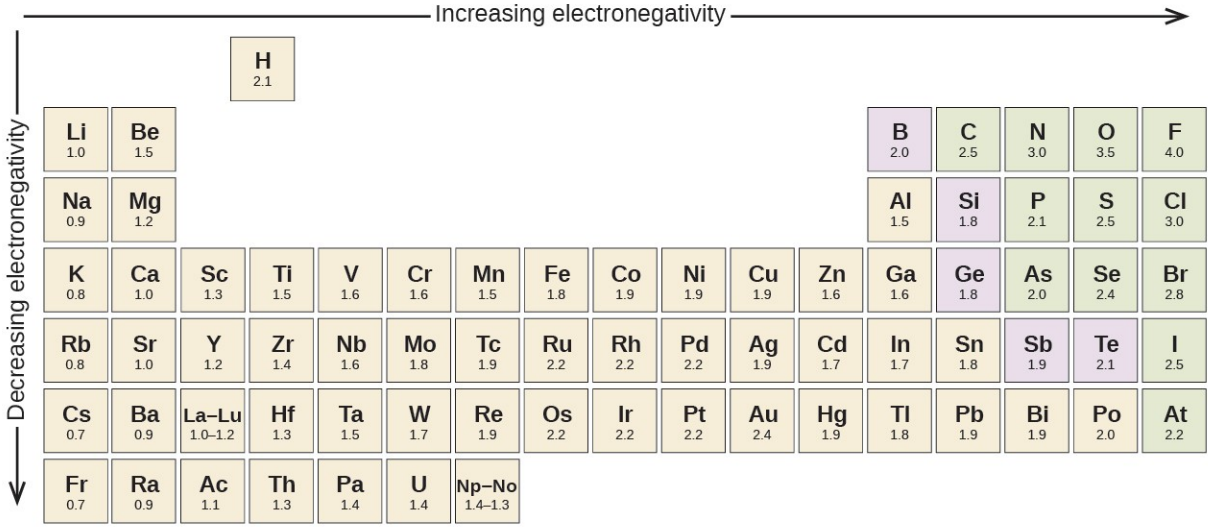
9. Determine the formal charge on each atom in the following molecules or ions:
    a. HCO_2^{-1} (formate ion)
    b. HCO_2H (formic acid)
    c. CO_2 (two resonance forms). Which resonance form should be more stable?
10. For each of the bonds listed below, tell which atom is the more negatively charged. Use values of electronegativity to support your answer.
    a. C-N
    b. C-H
    c. C-Br
    d. S-O
11. Considering both formal charge and electronegativities, predict on which atom or atoms the positive charge resides in the following cations:
    a. H_3O^+
    b. NH_4^+
    c. NO_2^+
    d. NF_4^+
12. Determine if the following molecules are polar or nonpolar. Use a table of electronegativities to determine the most polar bond in the molecules below.
    a. CS_2 b. CF_4 c. PCl_3 d. CO
13. Give the bond order for each bond in the following molecules or ions:
    a. CN^{-1}
    b. CH_3CN
    c. $CH_3CH{=}CH_2$
    d. SO_3^{2-}
14. Phosgene, $COCl_2$, is a highly toxic gas that was used as a weapon in World War I. Using bond energies, estimate the enthalpy change for the reaction of carbon monoxide and chlorine to produce phosgene. (*Hint:* First draw the electron dot structures of the reactants and products so you know the types of bonds involved.) The reaction:

$$CO(g) + Cl_2(g) \rightarrow COCl_2(g)$$

Bond Enthalpy Values (kJ/mol)

Bond	Bond Energy	Bond	Bond Energy	Bond	Bond Energy
H–H	436	C–S	260	F–Cl	255
H–C	415	C–Cl	330	F–Br	235
H–N	390	C–Br	275	Si–Si	230
H–O	464	C–I	240	Si–P	215
H–F	569	N–N	160	Si–S	225
H–Si	395	N=N	418	Si–Cl	359
H–P	320	N≡N	946	Si–Br	290
H–S	340	N–O	200	Si–I	215
H–Cl	432	N–F	270	P–P	215
H–Br	370	N–P	210	P–S	230
H–I	295	N–Cl	200	P–Cl	330
C-C	345	N–Br	245	P–Br	270
C=C	611	O–O	140	P–I	215
C≡C	837	O=O	498	S–S	215
C–N	290	O–F	160	S–Cl	250
C=N	615	O–Si	370	S–Br	215
C≡N	891	O–P	350	Cl–Cl	243
C–O	350	O–Cl	205	Cl–Br	220
C=O	741	O–I	200	Cl–I	210
C≡O	1080	F–F	160	Br–Br	190
C–F	439	F–Si	540	Br–I	180
C–Si	360	F–P	489	I–I	150
C–P	265	F–S	285		

Electronegativity Values



			H 2.1													
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La–Lu 1.0–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1	Th 1.3	Pa 1.4	U 1.4	Np–No 1.4–1.3										

CH 222 Winter 2025:

Problem Set #2

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #2 by **reading Chapter 8** of the textbook and/or by watching the videos found on our website (https://mhchem.org/222)
- **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 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 H1:* due **Wednesday, January 22 at 1:10 PM**

Section 01 ***and*** *Section W1:* **Watch the recitation video** for Problem Set #2 here:
http://mhchem.org/y/u.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.)

- *Section W1:* **S**ubmit 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, January 22.**
- *Section 01:* Due to MLK day, submit Problem Set #2 as a hard copy (nothing electronic) on **9 AM, Wednesday January 22 in AC 1303** (lecture.)

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

CH 222 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 Eight** *and* **Chapter Guide Two**
Important Tables and/or Constants: "**MO Diagram for B_2, C_2, and N_2**" / "**MO Diagram for O_2, F_2, and Ne_2**" *(Handouts,* http://mhchem.org/MO*)*, "**Geometry and Polarity Guide**" *(Handout,* https://mhchem.org/geopo*)*

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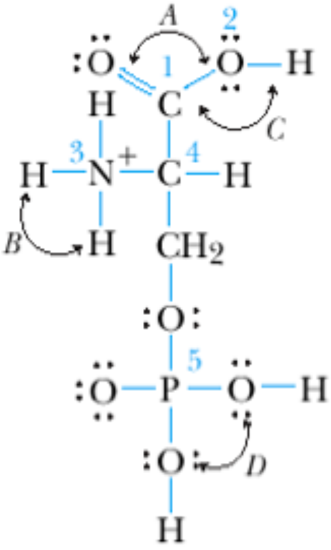
1. Draw the Lewis structure of NF_3. What are its electron pair and molecular geometries? What is the hybridization of the nitrogen atom? What orbitals on N and F overlap to form bonds between these elements?
2. Specify the electron pair and molecular geometry for each of the following. Describe the hybrid orbital set used by the central atom in each molecule or ion.
   a. CSe_2
   b. SO_2
   c. CH_2O
   d. NH_4^+
3. Draw the Lewis structure and then specify the electron pair and molecular geometries for each of the following molecules or ions. Identify the hybridization of the central atom.
   a. $XeOF_4$
   b. OSF_4
   c. BrF_5
   d. The central atom in Br_3^{-1}
4. The compound C_4H_8 has six isomers. Draw them. (*Note:* 4 of them have a double bond.)
5. Give the electron configurations for the Li_2, Li_2^{+1} and Li_2^{-1} in molecular orbital terms. Compare the Li-Li bond order in the three species; which has the shortest bond length?
6. Oxygen, O_2, can acquire one or two electrons to give O_2^{-1} (superoxide ion) or O_2^{2-} (peroxide ion.) Write the molecular orbital configuration for O_2, O_2^{-1} and O_2^{2-}. Remember to use the molecular orbital diagram for O_2, F_2 and Ne_2 when constructing the diagrams. For each species, determine the
   a. Magnetic character
   b. Net number of σ and π bonds
   c. Bond order
   d. Relative oxygen-oxygen bond length
7. The nitrosyl ion, NO^+, has an interesting chemistry. Use the "O_2, F_2 and Ne_2" molecular orbital diagram for this problem.
   a. Is NO^+ diamagnetic or paramagnetic? If paramagnetic, how many unpaired electrons does it have?
   b. What is the highest energy occupied molecular orbital (HOMO) in the molecule? What is the lowest unoccupied molecular orbital (LUMO) in the molecule?
   c. What is the nitrogen-oxygen bond order?
   d. Is the N-O bond in NO^+ stronger or weaker than the bond in NO? Explain.

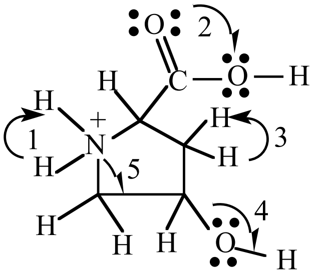
Problem Set #2 continues on the next page

8. Nitrogen, N_2, can ionize to form N_2^+ or add an electron to form N_2^{-1}. Using molecular orbital theory, compare these three species with regard to:
    a. Their magnetic character
    b. Net number of π bonds
    c. Bond order
    d. Bond length
    e. Bond strength
9. Phosphoserine is a less common amino acid with the structure shown to the right.
    a. Describe the hybridization of atoms 1 through 5.
    b. What are the approximate values of the bond angles *A*, *B*, *C* and *D*?



Phosphoserine

10. Sketch the Lewis structures of ClF_2^+ and ClF_2^{-1}. What are the electron pair and molecular geometries of each ion? Do both have the same F-Cl-F- angle? What hybrid set is used in each ion?
11. Compare the structure and bonding in CO_2 and CO_3^{2-} with regard to:
    a. The O-C-O bond angles
    b. The CO bond order
    c. The C atom hybridization.
    d. Does the molecule CO have a stronger bond than CO_2 and/or CO_3^{2-}? Explain.
12. Hydroxyproline is an unusual amino acid with the structure shown to the right.
    a. What are the approximate values for the bond angles for *1*, *2*, *3*, *4* and *5*?
    b. Describe the hybridization around the central atom for *1*, *2*, *3*, *4* and *5*.



Hydroxyproline

13. Iodine and oxygen form a complex series of ions, among them IO_4^{-1} and IO_5^{-3}. Draw the Lewis structures for these ions and specify their electron pair and molecular geometries. What is the hybridization of the I atom in these ions?
14. Which of the following molecules or ions should be paramagnetic? What is the highest occupied molecular orbital (HOMO) in each one? Note that if an O, F or Ne is present in the molecule, you should use the molecular orbital diagram for O_2, F_2 and Ne_2 to construct the molecule.
    a. NO
    b. OF^{-1}
    c. O_2^{2-}
    d. Ne_2^{+1}
    e. CN

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CH 222 Winter 2025:

Problem Set #3

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #3 by **reading Chapter 20** of the textbook and/or by watching the videos found on our website (https://mhchem.org/222)
- **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, January 27 at 1:10 PM**
- *Section H1:* due **Wednesday, January 29 at 1:10 PM**

Section W1: **Watch the recitation video** for Problem Set #3 here:

http://mhchem.org/y/v.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, January 29.**

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

CH 222 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 Twenty** *and* **Chapter Guide Three**
Important Tables and/or Constants: "**Organic Chemistry Nomenclature Guide**" *(Handout,* http://mhchem.org/OCG*)*, "**Organic Chemistry Lab**"

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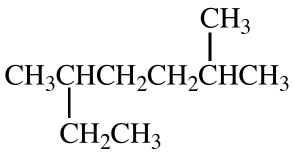
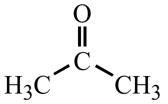
1. What is the molecular formula for an alkane with 6 carbon atoms? Draw and name the five isomers.
2. Give the systematic name for the alkane shown to the right:
3. Draw the structure for cycloheptane. Is the seven-member ring planar? Is this an isomer of n-heptane? Explain.
4. What structural requirement is necessary for an **alkene** to have *cis* and *trans* isomers? Can *cis* and *trans* isomers exist for an alkyne or an alkane? Draw and name the five alkenes with the formula C_7H_{14} and a seven carbon ("straight") chain.

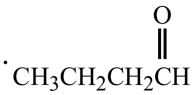


$CH_3\underset{\displaystyle CH_2CH_3}{\underset{|}{C}}HCH_2CH_2\overset{\displaystyle CH_3}{\overset{|}{C}}HCH_3$

Structure for Problem #2

5. Draw structural formulas for the following compounds:
   a. butan-1-ol *and* butan-2-ol
   b. 1,1-dibromoethane
   c. 3,3-dimethylbutan-2-ol
   d. 3-methyl-1-butyne
6. Draw the structures of the following compounds:
   a. diethyl ether
   b. 2-methoxypropane
7. Name the following compounds:

a. 

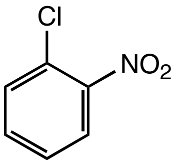
b. 

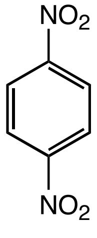
$CH_3CH_2CH_2\overset{O}{\overset{\|}{C}}H$

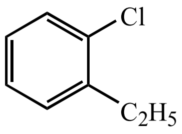
c. 

$CH_3\overset{O}{\overset{\|}{C}}CH_2CH_2CH_3$

8. Regarding structural isomers:
   a. Draw all the possible isomers for C_3H_8O. Give the systematic name for each compound.
   b. Draw the structural formula for an aldehyde and a ketone with the molecular formula C_4H_8O. Name each compound.
9. Give the systematic name for each of the following compounds:

a. 

b. 

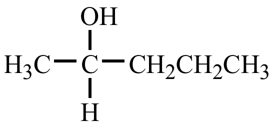
c. 

10. Draw structural formulas for the following carboxylic acids:
    a. 2-methylhexanoic acid
    b. 3-methylpentanoic acid
    c. acetic acid

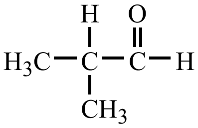
Problem Set #3 continues on the next page

Problem Set #3, Continued from previous page

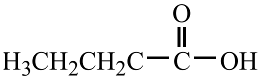
11. Name the following amines:
    a. $CH_3CH_2CH_2NH_2$
    b. $(CH_3)_3N$
    c. $(CH_3)(C_2H_5)NH$
    d. $CH_3CH_2CH_2CH_2CH_2CH_2NH_2$
12. Name each compound:

a. 

$H_3C-CH(OH)-CH_2CH_2CH_3$

b. 

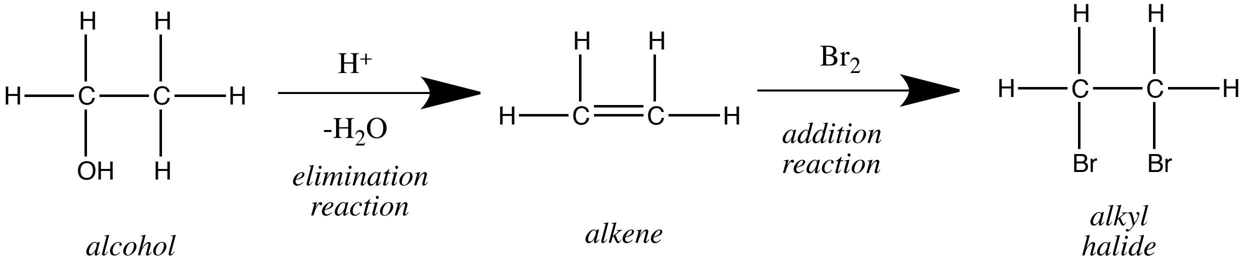
$H_3C-CH(CH_3)-C(=O)-H$

c. 

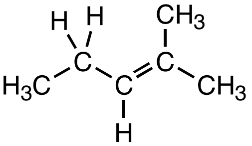
$H_3CH_2CH_2C-C(=O)-OH$

13. Give the official IUPAC organic chemistry name for the following chemical compounds:
    a. acetone
    b. toluene
    c. acetylene
    d. formaldehyde
    e. acetaldehyde
    f. ethylene

The following diagram might prove useful for the following three questions:



14. Draw the structure and give the systematic name for the products of the following addition reactions:

a. 

$H_3C-CH_2-CH=C(CH_3)-CH_3$ + H_2 →

b. $CH_3C\equiv CCH_2CH_3$ + 2 Br_2 →

15. The compound 2,3-dibromo-2-methylhexane is formed by addition of Br_2 to an alkene. Identify the alkene and write an equation for this reaction.
16. Addition of acid (H^+) to propan-2-ol creates an alkene through an elimination reaction. Draw and name the alkene. If the alcohol was propan-1-ol, would the synthesized alkene be different? Explain.

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CH 222 Exam Prep I Worksheet Name: ______________________________

Worksheet due dates: **Mon, 2/3, 1:10 PM** *(01)* **, Wed, 2/5, 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.*

Problem 1: Using the bond energies in the table below, estimate ΔH for the reaction: $\mathbf{H_2O_2 + CH_3OH \rightarrow H_2CO + 2\ H_2O}$

Bond:	C-C	C=C	C-O	C=O	C-H	O-H	O-O
Energy (kJ/mol)	347	614	358	799	413	463	146

Answer to Problem #1: **-345 kJ**

Problem 2: One of the following species could probably not exist as a stable molecule as predicted by Lewis structures. Draw the structures and **explain why one of them is an improbable structure**.

$\mathbf{NH_3, N_2H_2, N_2H_4, N_2H_6, N_2O_4}$

Hint to Problem #2: **Look for too many bonds around the central atom.**

Problem 3: What is the molecular geometry for the following structure: $\mathbf{BeF_3^{-1}}$ (draw the structure to get credit!)

Answer to Problem #3: **trigonal planar**

Problem 4: Draw the molecular orbital description of the NO^{-1} anion. Is NO^{-1} paramagnetic? What is the bond order for NO^{-1}? Is NO^{-1} isoelectronic with CO? **Define isoelectronic to receive credit.**

Partial answer to Problem #4: **NO^{-1} is paramagnetic, has a bond order = 2 and it is not isoelectronic with CO.**

Problem 5: A teacher gave a molecule the following name: **2-ethyl-3-methyl-5-isopropylhexane** (note that "isopropyl" is a propyl group where the connection is made at the center carbon.) However, an alert student recognized that although the molecule can be correctly drawn from this name, the name violates the systematic rules for organic chemistry. **What is the correct (systematic) name for the molecule**? Draw the structure and circle the longest chain to get credit.

Hint to Problem #5: **"isopropyl" is also known as "2-propyl"**

CH 222 Winter 2025:
Problem Set #4
Instructions

Step One (all sections):

- **Learn the material** for Problem Set #4 by **reading Chapter 9 and Chapter 10** of the textbook and/or by watching the videos found on our website (https://mhchem.org/222)
- **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, February 10 at 1:10 PM**
- *Section H1:* due **Wednesday, February 12 at 1:10 PM**

Section W1: **Watch the recitation video** for Problem Set #4 here:
http://mhchem.org/y/w.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, February 12.**

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

CH 222 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 Nine, Chapter Ten** *and* **Chapter Guide Four**
Important Tables and/or Constants: **R = 0.082057 L atm mol^{-1} K^{-1}, 760 mm Hg = 1 atm = 1013 mbar, 1 mbar = 1 hPa, 1 torr = 1 mm Hg**

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1. A sample of CO_2 gas has a pressure of 56.5 mm Hg in a 125 mL flask. What is the volume of this gas sample when it has a pressure of 62.3 torr at the same temperature?
2. Define STP for gases. A 5.0 mL sample of CH_4 gas at 1 atm is enclosed in a gas tight syringe at 22 °C. If the syringe is immersed in an ice bath at STP, what is the new gas volume?
3. A steel cylinder holds 1.50 g of ethanol, CH_3CH_2OH. What is the pressure of the ethanol vapor if the cylinder has a volume of 251 cm^3 and the temperature is 250 °C? Assume all the ethanol is in the vapor phase at this temperature.
4. Diethyl ether, $(CH_3CH_2)_2O$, vaporizes easily at room temperature. If the vapor exerts a pressure of 311 mbar in a flask at 25 °C, what is the density of the vapor?
5. Chloroform is a common liquid used in the laboratory. It vaporizes readily. If the pressure of the chloroform vapor in a flask is 195 mm Hg at 25.0 °C, and the density of the vapor is 1.25 g/L, what is the molar mass of the chloroform?
6. A 0.0125 g sample of a gas with an empirical formula of CHF_2 is placed in a 165 mL flask. It has a pressure of 13.7 mm Hg at 22.5 °C. What is the molecular formula for this compound?
7. Silane, SiH_4, reacts with O_2 to give silicon dioxide and water according to the equation below. A 5.20 L sample of silane gas at 356 mm Hg and 25 °C is allowed to react with O_2 gas. What volume of O_2 gas, in liters, is required for the complete reaction if the oxygen has a pressure of 425 mm Hg at 25 °C? The reaction:
$$\mathbf{SiH_4(g) + 2\,O_2(g) \rightarrow SiO_2(s) + 2\,H_2O(\ell)}$$
8. **a)** A cylinder of compressed gas is labeled "Composition (mole %): 4.5% H_2S, 3.0% CO_2, balance N_2." The pressure gauge attached to the cylinder reads 46 atm. Calculate the partial pressure of each gas, in atmospheres, in the cylinder. **b)** If oxygen gas is collected at 20 °C over water (with a vapor pressure of 17.5 torr), what is the pressure of the oxygen if the total pressure is exactly 1 atm?
9. Equal masses of gaseous N_2 and Ar are placed in separate flasks of equal volume at the same temperature. Tell whether each of the statements is true or false, and briefly explain your answer.
   a. There are more molecules of N_2 present than atoms of Ar.
   b. The pressure is greater in the Ar flask.
   c. The Ar atoms have a greater average speed than the N_2 molecules.
   d. The N_2 molecules collide more frequently with the walls of the flask than do the Ar atoms.
10. The reaction of SO_2 with Cl_2 gives dichlorine monoxide (see the reaction below) that is used to bleach wood pulp and to treat wastewater. All of the compounds involved in the reaction are gases. List them in order of increasing average speed.
$$\mathbf{SO_2(g) + 2\,Cl_2(g) \rightarrow OSCl_2(g) + Cl_2O(g)}$$

Problem Set #4 continues on the next page

11. In each pair of gases below, tell which will effuse faster:
    a. CO_2 or F_2
    b. O_2 or N_2
    c. C_2H_4 or C_2H_6
    d. two chlorofluorocarbons: $CFCl_3$ or $C_2Cl_2F_4$
12. Analysis of a gaseous chlorofluorocarbon ($C_xCl_yF_z$) shows that it contains 11.79% C and 69.57% Cl. In another experiment you find that 0.107 g of the compound fills a 458 mL flask at 25 °C with a pressure of 21.3 mm Hg. What is the molecular formula of the compound?
13. You are given 1.56 g of a mixture of $KClO_3$ and KCl. When heated, $KClO_3$ decomposes to KCl and O_2 according to the reaction shown below. If the 1.56 g of mixture creates 327 mL of O_2 collected in a flask with a pressure of 735 mm Hg at 19 °C, what is the weight percent of $KClO_3$ in the mixture?

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

14. What type of intermolecular forces must be overcome in converting each of the following from a liquid to a gas?
    a. CO_2
    b. $CHCl_3$
    c. NH_3
    d. SCl_4
    e. $I_2(CH_3OH)$
    f. $Na^+(aq)$
15. Rank the following in order of increasing intermolecular force strength. At 25 °C and 1 atm, which exist as gases and which exist as liquids?
    a. $CH_3CH_2CH_3$
    b. CH_3CH_2OH
    c. He
    d. water
16. The enthalpy of vaporization of liquid mercury is 59.11 kJ/mol. What quantity of heat is required to vaporize 0.500 mL of mercury at 357 °C, its normal boiling point? The density of Hg is 13.6 g/mL.

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CH 222 Winter 2025:

Problem Set #5

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #5 by **reading Chapter 10 and Chapter 11** of the textbook and/or by watching the videos found on our website (https://mhchem.org/222)
- **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:

Section 01 and H1: 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, February 17 at 1:10 PM**
- *Section H1:* due **Wednesday, February 19 at 1:10 PM**

Section W1: **Watch the recitation video** for Problem Set #5 here:

http://mhchem.org/y/x.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, February 19.**

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

CH 222 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 Ten, Chapter Eleven** *and* **Chapter Guide Five**
Important Tables and/or Constants: **R = 8.3145 J mol^{-1} K^{-1}**, "**Cubic Unit Cells Guide**" *(Handout)*, "**Solids**" *(Lab)*

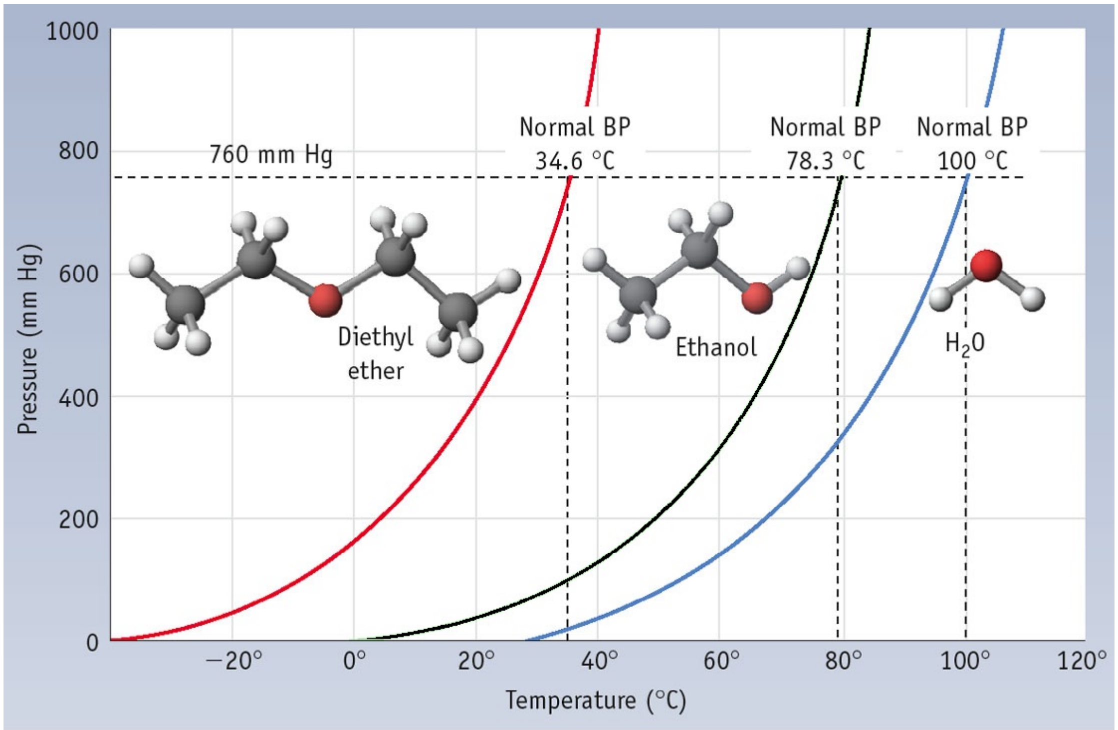
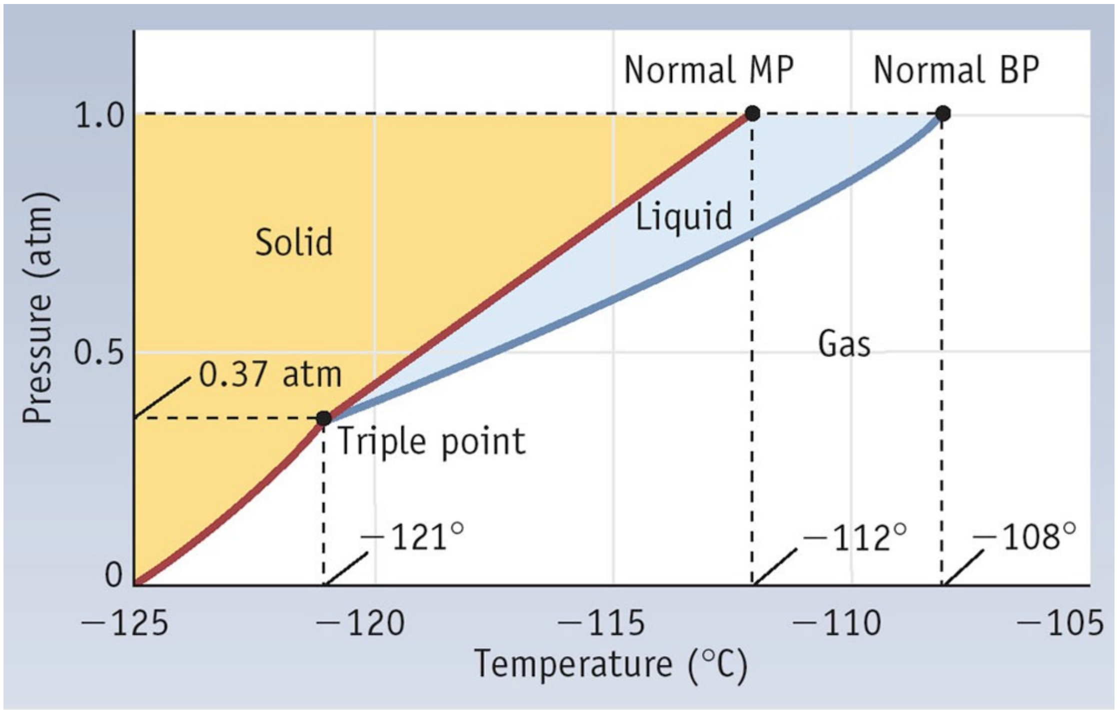


Figure for Problem One

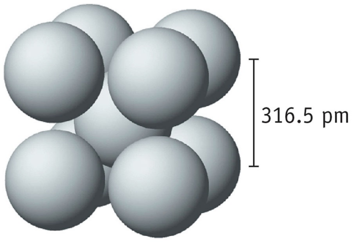
1. Use the figure above to answer the following questions:
    a. What is the equilibrium vapor pressure of diethyl ether at room temperature (approximately 20 °C)?
    b. Place diethyl ether, ethanol and water in order of increasing intermolecular forces.
    c. If the pressure in a flask is 400 mm Hg and the temperature is 40 °C, which of the three compounds are liquids and which are gases?
2. Answer each of the following questions with *increases, decreases* or *does not change*.
    a. If the intermolecular forces in a liquid increase, the normal boiling point of the liquid ______.
    b. If the intermolecular forces in a liquid decrease, the vapor pressure of the liquid ______.
    c. If the surface area of a liquid decreases, the vapor pressure ______.
    d. If the temperature of a liquid increases, the equilibrium vapor pressure ______.

Problem Set #5 continues on the next page



Phase Diagram for Xenon

3. Use the phase diagram for xenon given above to answer the following questions:
    a. In what phase is the xenon found at room temperature and 1.0 atm pressure?
    b. If the pressure exerted on a sample is 0.75 atm and the temperature is -114 °C, in what phase does the substance exist?
    c. If you measure the vapor pressure of a liquid sample and find it to be 380 mm Hg, what is the temperature of the liquid phase?
    d. What is the vapor pressure of the solid at -122 °C?
    e. Which is the denser phase, solid or liquid? Explain.
4. The specific heat capacity of silver is 0.235 J g^{-1} K^{-1}. Its melting point is 962 °C and its heat of fusion is 11.3 kJ/mol. What quantity of heat, in Joules, is required to change 5.00 g of silver from a solid at 25 °C to a liquid at 962 °C?
5. If your air conditioner is more than several years old, it may use the chlorofluorocarbon CCl_2F_2 as the heat transfer fluid. The normal boiling point of CCl_2F_2 is -29.8 °C, and the enthalpy of vaporization is 20.11 kJ/mol. The gas and the liquid have specific heats of 117.2 J mol^{-1} K^{-1} and 72.3 J mol^{-1} K^{-1}, respectively. How much heat is evolved when 20.0 g of CCl_2F_2 is cooled from +40.0 °C to -40.0 °C?
6. Tungsten crystallizes in the unit cell shown to the right.
    a. What type of unit cell is this?
    b. How many tungsten atoms occur per unit cell?
    c. If the edge of the unit cell is 316.5 pm, what is the radius of a tungsten atom?



Tungsten unit cell

Problem Set #5 continues on the next page

Problem Set #5, Continued from previous page

7. Calcium metal crystallizes in a face-centered cubic unit cell. The density of the solid is 1.54 g/cm^3. What is the radius of a calcium atom?
8. Equilibrium vapor pressures of dichlorodimethylsilane, $SiCl_2(CH_3)_2$, are given below.

Temperature (°C)	**Vapor Pressure (mm Hg)**
-0.4	40.
+17.5	100.
+51.9	400.
+70.3	760.

   a. What is the normal boiling point of dichlorodimethylsilane?
   b. Plot these data as ln P versus $1 / T$. Perform a linear regression on the data. At what temperature does the liquid have an equilibrium vapor pressure of 250 mm Hg?
   c. Calculate the molar enthalpy of vaporization for dichlorodimethylsilane using the Clausius-Clapeyron equation.

9. Fill in the blanks in the table. All solutions are aqueous.

Compound	**Molality**	**Weight Percent**	**Mole Fraction**
KNO_3	______	10.0	_____
CH_3CO_2H	0.0183	_____	_____
$HOCH_2CH_2OH$	_____	_____	0.0599

10. Concentrated aqueous sulfuric acid has a density of 1.84 g/cm^3 and is 95.0% by weight H_2SO_4. What is the molarity of this acid? What is the molality?
11. Silver ion has an average concentration of 28 ppb (parts per billion) in U.S. water supplies.
    a. What is the molality of the silver ion?
    b. If you wanted 1.0×10^2 g of silver and could recover it chemically from water supplies, what volume of water, in liters, would you have to treat? Assume the density of water is 1.0 g/cm^3.
12. Hydrogen gas has a Henry's law constant of 1.07×10^{-6} M/ mm Hg at 25 °C when dissolving in water. If the total pressure of the gas (H_2 plus water vapor) over water is 1.00 atm, what is the concentration of H_2 in the water in grams per milliliter? The vapor pressure of water at 25 °C is 23.8 mm Hg.
13. Nonvolatile urea, $(NH_2)_2CO$, which is widely used in fertilizers and plastics, is quite soluble in water. If you dissolve 9.00 g of urea in 10.0 mL of water, what is the vapor pressure of the solution at 24 °C? Assume the density of water is 1.00 g/ mL, and the vapor pressure of water at 24 °C is 22.4 mm Hg.
14. What is the boiling point of a solution composed of 0.755 g of caffeine, $C_8H_{10}O_2N_4$, in 95.6 g of benzene, C_6H_6? The normal boiling point for benzene is 80.10 °C and K_{bp} for benzene = 2.53 °C/ m.

Problem Set #5 continues on the next page

15. Some ethylene glycol, $HOCH_2CH_2OH$, is added to your car's cooling system along with 5.0 kg of water. If the freezing point of the water-glycol solution is -15.0 °C, what mass of $HOCH_2CH_2OH$ must have been added? The normal freezing point for water is 0.0 °C and K_{fp} for water = 1.86 °C/ *m*.
16. Butylated hydroxyanisole (BHA) is used as an antioxidant in margarine and other fats and oils; it prevents oxidation and prolongs the shelf life of the food. What is the molar mass of BHA if 0.640 g of the compound, dissolved in 25.0 g of chloroform ($CHCl_3$), produces a solution whose boiling point is 62.22 °C? The normal boiling point for chloroform is 61.70 °C and K_{bp} for chloroform = 3.63 °C/ *m*.
17. The organic compound aluminon is used as a reagent to test for the presence of the aluminum ion in aqueous solution. A solution of 2.50 g of aluminon in 50.0 g of water freezes at -0.197 °C. What is the molar mass of aluminon? The normal freezing point for water is 0.000 °C and K_{fp} for water = 1.86 °C/ *m*.
18. Estimate the osmotic pressure of human blood at 37 °C. Assume blood is isotonic with a 0.154 M NaCl solution, and use a theoretical van't Hoff *i* factor for NaCl.

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CH 222 Exam Prep II Worksheet Name: ____

Worksheet due dates: **<u>Mon, 2/24</u>, 1:10 PM** *(01)* **, <u>Wed, 2/26</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>: A 3.31 g sample of lead(II) nitrate is heated in an evacuated cylinder with a volume of 1.62 L. The salt decomposes when heated according to the equation below. Assuming complete decomposition, what is the pressure in the cylinder after decomposition and cooling to a temperature of 300. K? Assume that the PbO(s) takes up negligible volume.

$$2\ Pb(NO_3)_2(s) \rightarrow 2\ PbO(s) + 4\ NO_2(g) + O_2(g)$$

Answer to Problem #1: **0.380 atm**

<u>Problem 2</u>: How much energy is needed to convert 64.0 g of ice at 0.00 °C to liquid water at 75.0 °C? Note that the Heat of fusion for water = 333 J/g.

Answer to Problem #2: **41.4 kJ**

Problem 3: Concentrated nitric acid is 70.0% by mass HNO_3 in water. The density of this acid is 1.42 g/cm^3. What is the molarity of the acid?

Answer to Problem #3: **15.8 M**

Problem 4: A 5.50 g sample of a compound is dissolved in 250. g of benzene. The freezing point of this solution is 1.02 °C below that of pure benzene. What is the molar mass of the compound? (k_f for benzene = -5.12 °C/m)

Answer to Problem #4: **110. g/mol**

Problem 5: Polyethylene is a synthetic polymer or plastic with many uses. 1.40 g of a polyethylene sample was dissolved in enough benzene to make 100. mL of solution, and the osmotic pressure was found to be 1.86 torr at 25 °C. What is the molar mass of the polyethylene?

Answer to Problem #5: **1.40 * 10^5 g/mol**

CH 222 Winter 2025:

Problem Set #6

Instructions

Step One (all sections):

- **Learn the material** for Problem Set #6 by **reading Chapter 12 and Chapter 21** of the textbook and/or by watching the videos found on our website (https://mhchem.org/222)
- **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, March 10 at 1:10 PM**
- *Section H1:* due **Wednesday, March 12 at 1:10 PM**

Section W1: **Watch the recitation video** for Problem Set #6 here:
http://mhchem.org/y/y.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, March 12.**

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

CH 222 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 Twelve, Chapter Twenty-one*** *and* ***Chapter Guide Six***
Important Tables and/or Constants: **R = 8.3145 J mol^{-1} K^{-1}**, "**Reaction Mechanisms Guide**" *(Handout)*

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1. Give the equation for the relative rates of disappearance of reactants and formation of products for the following reaction:
    a. $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$
    b. If $\Delta[H_2]/\Delta t = -4.5 \times 10^{-4}$ M min^{-1}, what is $\Delta[NH_3]/\Delta t$?
2. Nitrosyl bromide, NOBr, is formed from NO and Br_2. Experiments show that this reaction is second order in NO and first order in Br_2. The equation:
$$2\ NO(g) + Br_2(g) \rightarrow 2\ NOBr(g)$$
    a. Write the rate law equation for the reaction.
    b. How does the initial reaction rate change if the concentration of Br_2 is changed from 0.0022 M to 0.0066 M?
    c. What is the change in the initial rate if the concentration of NO is changed from 0.0024 M to 0.0012 M?
3. The reaction:
$$2\ NO(g) + 2\ H_2(g) \rightarrow N_2(g) + 2\ H_2O(g)$$
was studied at 904 °C and the data in the table below were collected.

Reactant Concentration (M)		
[NO]	**[H_2]**	**Rate of Appearance of N_2 (M s^{-1})**
0.420	0.122	0.136
0.210	0.122	0.0339
0.210	0.244	0.0678
0.105	0.488	0.0339

    a. Determine the order of the reaction for each reactant.
    b. Write the rate law equation for the reaction.
    c. Calculate the rate constant for the reaction.
    d. Find the rate of appearance of N_2 at the instant when [NO] = 0.350 M and [H_2] = 0.205 M.
4. The decomposition of N_2O_5 in CCl_4 is a first order reaction. If 2.56 mg of N_2O_5 is present initially, and 2.50 mg is present after 4.26 min at 55 °C, what is the value of the rate constant, *k*, at 55 °C?
5. The conversion of cyclopropane to propene occurs with a first order rate constant equal to 5.4×10^{-2} h^{-1}. How long will it take for the concentration of cyclopropane to decrease from an initial concentration of 0.080 M to 0.020 M?
6. Gold-198 is used in the diagnosis of liver problems. The half-life of ^{198}Au is 2.69 days. If you begin with 2.8 μg of this gold isotope, what mass remains after 10.8 days?

Problem Set #6 continues on the next page

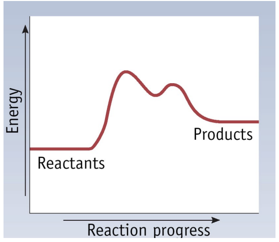
Problem Set #6, Continued from previous page

7. Strontium-90 is a hazardous radioactive isotope that resulted from atmospheric testing. A sample of strontium carbonate containing ^{90}Sr is found to have an activity of 1.00 x 10^3 dpm. One (1.00) year later the activity of this sample is 975 dpm.
   a. Calculate the half-life of strontium-90 from this information.
   b. How long will it take for the sample activity to drop to 1.00% of its initial value?
8. Ammonia decomposes when heated according to the equation shown below. The data in the table for this reaction were collected at a high temperature. Plot [NH_3] versus time, ln [NH_3] versus time and 1/[NH_3] versus time. What is the order with respect to NH_3? Find the rate constant for the reaction from the appropriate slope.

The reaction: $NH_3(g) \rightarrow NH_2(g) + H(g)$

Time (h)	**[NH_3] (M)**
0	8.00 x 10^{-7}
25	6.75 x 10^{-7}
50	5.84 x 10^{-7}
75	5.15 x 10^{-7}

9. Answer the following questions based on the reaction coordinate diagram shown to the right.
   a. Is the reaction exothermic or endothermic?
   b. Does this reaction occur in more than one step? If so, how many?



Reaction Diagram

10. What is the rate law equation for each of the following *elementary* reactions?
    a. $Cl(g) + ICl(g) \rightarrow I(g) + Cl_2(g)$
    b. $O(g) + O_3(g) \rightarrow 2\,O_2(g)$
    c. $2\,NO_2(g) \rightarrow N_2O_4(g)$
11. The reaction of $NO_2(g)$ and CO(g) is thought to occur in two steps:

Step 1 (slow) $NO_2(g) + NO_2(g) \rightarrow NO(g) + NO_3(g)$
Step 2 (fast) $NO_3(g) + CO(g) \rightarrow NO_2(g) + CO_2(g)$

   a. Add the elementary steps to find the overall, stoichiometric equation.
   b. What is the molecularity of each step? Which step is rate determining?
   c. For this mechanism to be consistent with kinetic data, what must be the experimental rate law equation?
   d. Identify any intermediates and/or catalysts in this reaction.
12. The data in the table below shows the temperature dependence of the rate constant for the reaction $N_2O_5(g) \rightarrow 2\,NO_2(g) + {}^1/_2\,O_2(g)$. Plot these data in the appropriate way to derive the activation energy and frequency factor for the reaction.

T(K)	***k*(s^{-1})**
338	4.87 x 10^{-3}
328	1.50 x 10^{-3}
318	4.98 x 10^{-4}
308	1.35 x 10^{-4}
298	3.46 x 10^{-5}
273	7.87 x 10^{-7}

Problem Set #6 continues on the next page

13. Complete the following nuclear equations. Write the mass number, atomic number and symbol for the remaining particle(s).
    - a. ${}^{9}_{4}\text{Be} + ? \rightarrow {}^{6}_{3}\text{Li} + {}^{4}_{2}\text{He}$
    - b. ${}^{241}_{95}\text{Am} + {}^{4}_{2}\text{He} \rightarrow {}^{243}_{97}\text{Bk} + ?$
    - c. ${}^{238}_{92}\text{U} + ? \rightarrow {}^{249}_{100}\text{Fm} + 5\,{}^{1}_{0}\text{n}$
    - d. Gallium-67 decays by electron capture.
    - e. Potassium-38 decays with positron decay.
    - f. Technetium-99m decays with γ emission.
14. Calculate the binding energy in kilojoules per mole of nucleons of P for the formation of ${}^{30}\text{P}$ and ${}^{31}\text{P}$. The required masses (in grams per mole) are ${}^{1}_{1}\text{H} = 1.00783$, ${}^{1}_{0}\text{n} = 1.00867$, ${}^{30}_{15}\text{P} = 29.97832$ and ${}^{31}_{15}\text{P} = 30.97376$.

CH 222 Final Exam Prep Worksheet Name: ____________________

Worksheet due dates: **At the time of your Lecture Final** *(01, H1)* **, Wed, 3/19, 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.*

Problem 1: The following data was collected for the reaction shown below. Determine the value of the rate constant (k). *Be sure to show the* ***orders*** *of each reactant in this reaction (and how you got them!)*

$$\mathbf{2\ MnO_4^{-1}(aq) + 5\ H_2C_2O_4(aq) + 6\ H^+(aq) \rightarrow 2\ Mn^{2+}(aq) + 10\ CO_2(g) + 8\ H_2O(l)}$$

$[MnO_4^{-1}]$	$[H_2C_2O_4]$	$[H^+]$	Rate (M/s)
$1 * 10^{-3}$	$1 * 10^{-3}$	1.0	$2 * 10^{-4}$
$2 * 10^{-3}$	$1 * 10^{-3}$	1.0	$8 * 10^{-4}$
$2 * 10^{-3}$	$2 * 10^{-3}$	1.0	$1.6 * 10^{-3}$
$2 * 10^{-3}$	$2 * 10^{-3}$	2.0	$1.6 * 10^{-3}$

Answer to Problem #1: **k = 2 * 10^5**

Problem 2: The decomposition of N_2O_5 ($2\ N_2O_5(g) \rightarrow O_2(g) + 4\ NO_2(g)$) is first order in N_2O_5 with $k = 1.0 * 10^{-5}\ s^{-1}$. If the initial concentration of N_2O_5 is $1.0 * 10^{-3}$ M, calculate the concentration of N_2O_5 after $1.0 * 10^5$ seconds.

Answer to Problem #2: **3.7 * 10^{-4} M**

Problem 3: The rate constant *k* for a reaction is 2.6 * 10^{-8} when the reaction proceeds at 300.0 K, and the activation energy is 98 kJ/mol. Determine the frequency factor, A, for the reaction. What is the value of *k* at 310 K?

Answer to Problem #3: **A = 3.0 * 10^{9}, k = 9.2 * 10^{-8}**

Problem 4: A sample of wood from an Egyptian mummy case gives a ^{14}C count of 9.4 cpm/gC (counts per minute per gram of carbon.) How old is the wood? (The initial decay rate of ^{14}C is 15.3 cpm/gC, and the ^{14}C half-life is 5730 years.)

Answer to Problem #4: **4.0 * 10^{3} years**

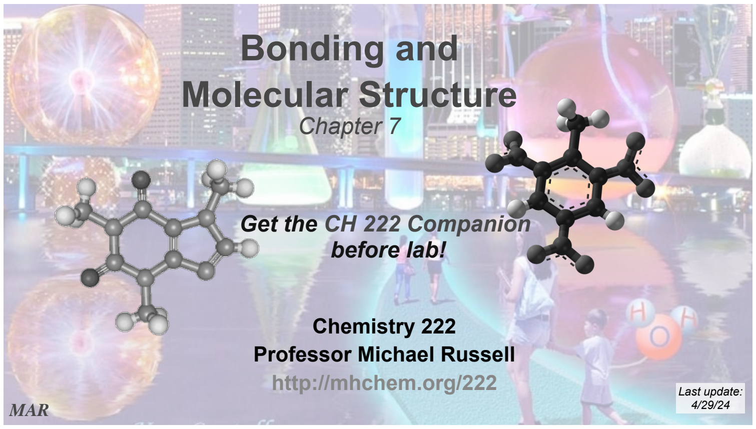
Problem 5: One of the hopes for solving the world's energy problem is to make use of the following fusion reaction:

$$^{2}H + {}^{3}H \rightarrow {}^{4}He + {}^{1}n + \text{energy}$$

How much energy (in kJ) is released when one mole of deuterium is fused with one mole of tritium according to the above reaction? The masses of the atoms and neutrons are:

^{2}H: 2.0140 amu ^{3}H: 3.01605 amu ^{4}He: 4.002603 amu ^{1}n: 1.008665 amu

Answer to Problem #5: **-1.69 * 10^{9} kJ**



Bonding and Molecular Structure

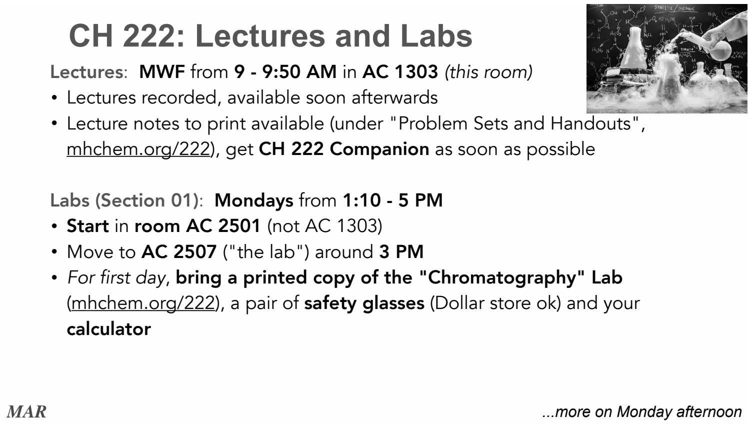
Chapter 7

Get the CH 222 Companion before lab!

Chemistry 222
Professor Michael Russell
http://mhchem.org/222

Last update: 4/29/24

MAR



CH 222: Lectures and Labs

Lectures: **MWF** from **9 - 9:50 AM** in **AC 1303** *(this room)*

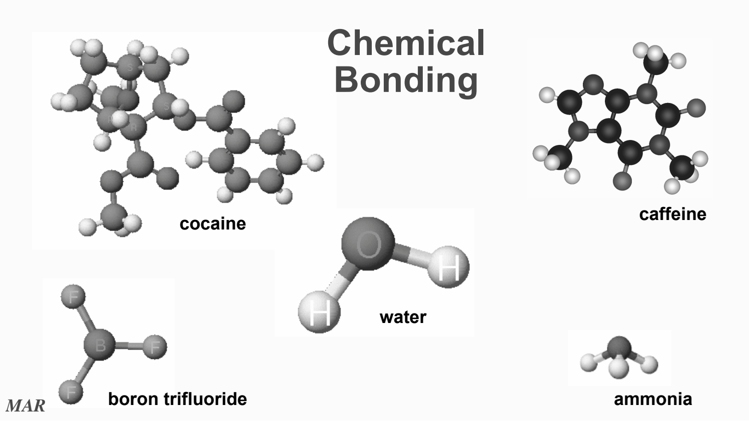
- Lectures recorded, available soon afterwards
- Lecture notes to print available (under "Problem Sets and Handouts", mhchem.org/222), get **CH 222 Companion** as soon as possible

Labs (Section 01): **Mondays** from **1:10 - 5 PM**

- **Start** in **room AC 2501** (not AC 1303)
- Move to **AC 2507** ("the lab") around **3 PM**
- *For first day*, **bring a printed copy of the "Chromatography" Lab** (mhchem.org/222), a pair of **safety glasses** (Dollar store ok) and your **calculator**

MAR

...more on Monday afternoon



Chemical Bonding

cocaine

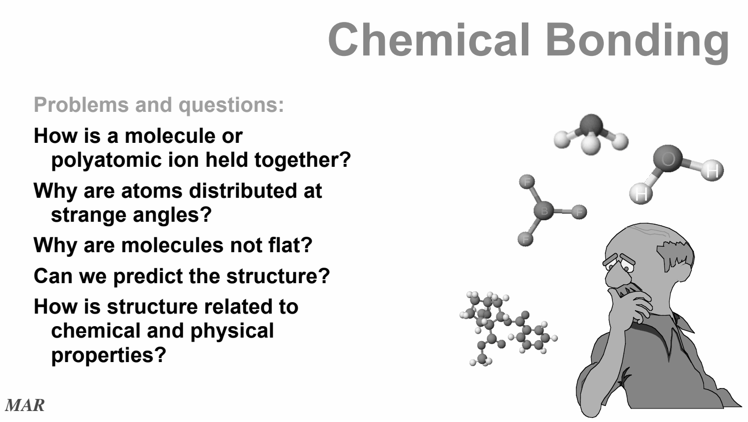
caffeine

water

boron trifluoride

ammonia

MAR



Chemical Bonding

Problems and questions:

How is a molecule or polyatomic ion held together?

Why are atoms distributed at strange angles?

Why are molecules not flat?

Can we predict the structure?

How is structure related to chemical and physical properties?

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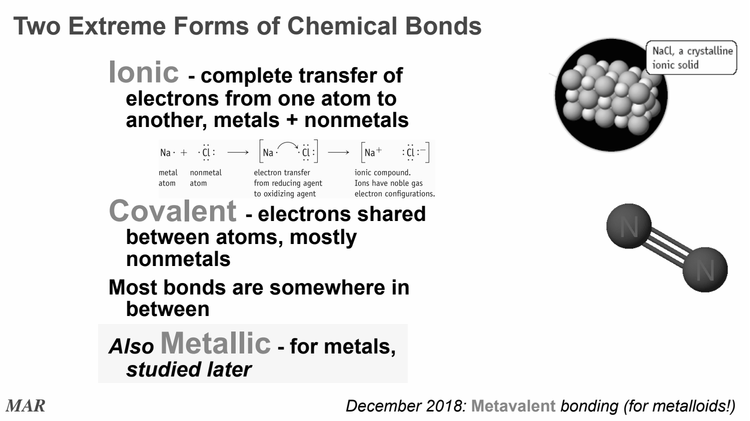


Two Extreme Forms of Chemical Bonds

Ionic - complete transfer of electrons from one atom to another, metals + nonmetals

$$Na\cdot + \cdot\ddot{\underset{..}{Cl}}: \longrightarrow [Na\cdot \ \cdot\ddot{\underset{..}{Cl}}:] \longrightarrow [Na^{+} \ :\ddot{\underset{..}{Cl}}:^{-}]$$

metal atom | nonmetal atom | electron transfer from reducing agent to oxidizing agent | ionic compound. Ions have noble gas electron configurations.

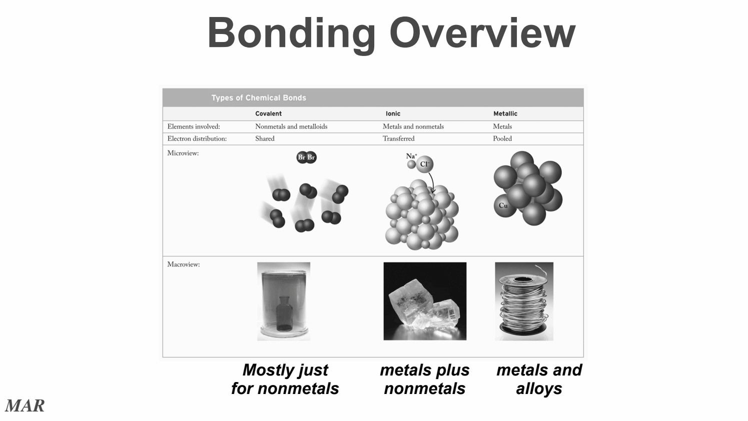
Covalent - electrons shared between atoms, mostly nonmetals

Most bonds are somewhere in between

Also* Metallic - *for metals, studied later

MAR

December 2018: Metavalent bonding (for metalloids!)



Bonding Overview

Types of Chemical Bonds

	Covalent	Ionic	Metallic
Elements involved:	Nonmetals and metalloids	Metals and nonmetals	Metals
Electron distribution:	Shared	Transferred	Pooled
Microview:			
Macroview:			
	Mostly just for nonmetals	***metals plus nonmetals***	***metals and alloys***

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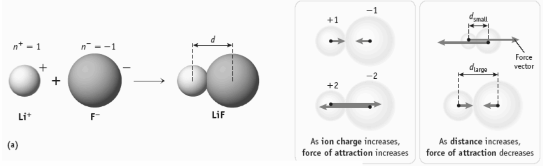
Ionic Forces - Coulomb's Law

CH 221 Flashback:

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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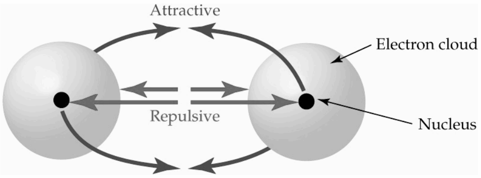
Higher ionic force, higher melting point, etc.

Covalent Bonding

Covalent bonds arise from the mutual attraction of 2 nuclei for the same electrons.



A covalent bond is a balance of <u>attractive</u> and <u>repulsive</u> forces.

Interatomic Interactions

H• •H

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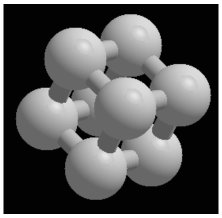
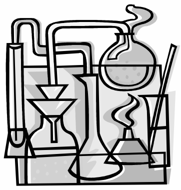
Covalent Bonding

- **Covalent bonding will be the focus of the first two chapters**
- **We will re-visit ionic bonding and Metallic bonding in a future chapter**
- **Important to know when a compound is ionic, covalent or metallic!**





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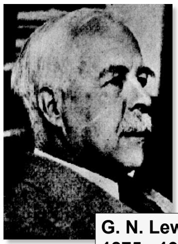
Electron Distribution in Molecules



G. N. Lewis
1875 - 1946

- **Electron distribution is depicted with Lewis electron dot structures**
- **Valence electrons are distributed as shared or BOND PAIRS and unshared or LONE PAIRS.**

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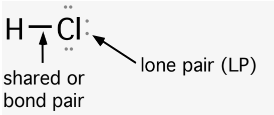
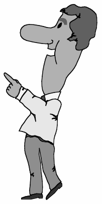
Bond and Lone Pairs

Valence electrons are distributed as shared or BOND PAIRS and unshared or LONE PAIRS.





This is called a LEWIS ELECTRON DOT structure

3 lone pairs + 1 bond pair = 4 pairs total

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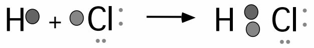
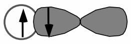
Bond Formation

A bond can result from a "head-to-head" overlap of atomic orbitals on neighboring atoms, making a sigma (σ) bond.





Overlap of H (1s) and Cl (3p)

Note that each atom has a single, unpaired electron in their atomic orbital.

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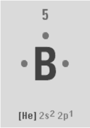
Valence Electrons

Electrons are divided between core and valence electrons

B $1s^2 2s^2 2p^1$

Core = [He] , valence = $2s^2 2p^1$

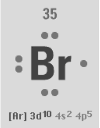
2 core e^-, 3 valence e^-



Br   [Ar] $3d^{10} 4s^2 4p^5$

Core = [Ar] $3d^{10}$, valence = $4s^2 4p^5$

28 core e^-, 7 valence e^-



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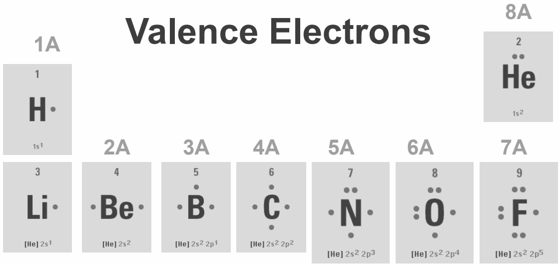
Valence Electrons



∴ Number of valence electrons is equal to the Group number

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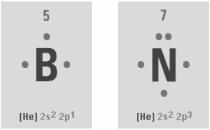
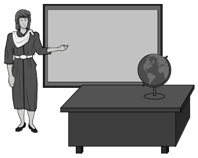
Building Lewis Structures

No. of valence electrons of a main group atom = Group number

For Groups 1A - 4A, no. of bond pairs = group number.

For Groups 5A - 7A, BPs = 8 - Grp. No.





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Building a Lewis Dot Structure

No. of valence electrons of an atom = Group number

For Groups 1A - 4A (14), no. of bond pairs = group number

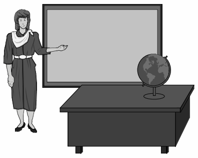
For Groups 5A (15) - 7A (17), BPs = 8 - Grp. No.

Except for H (and sometimes atoms of 3rd and higher periods),

BPs + LPs = 4

This observation is called the

OCTET RULE



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Building a Lewis Dot Structure

Ammonia, NH_3

1. Count valence electrons

   H = 1 and N = 5

   Total = (3 x 1) + 5

   = 8 electrons *or*

   4 pairs of electrons

2. Decide on the central atom; never H.

   Central atom is atom of lowest affinity for electrons.

   Therefore, N is central

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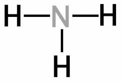
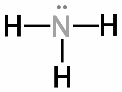
Building a Lewis Dot Structure

3. Form a sigma bond *(single bond)* between the central atom and surrounding atoms.



4. Remaining electrons form LONE PAIRS to complete octet as needed.



3 BOND PAIRS and 1 LONE PAIR. ***Note that N has a share in 4 pairs (8 electrons), while H shares 1 pair.***

Unshared electron pairs ("lone pairs") take up more volume than shared electron pairs ("bonding pairs")

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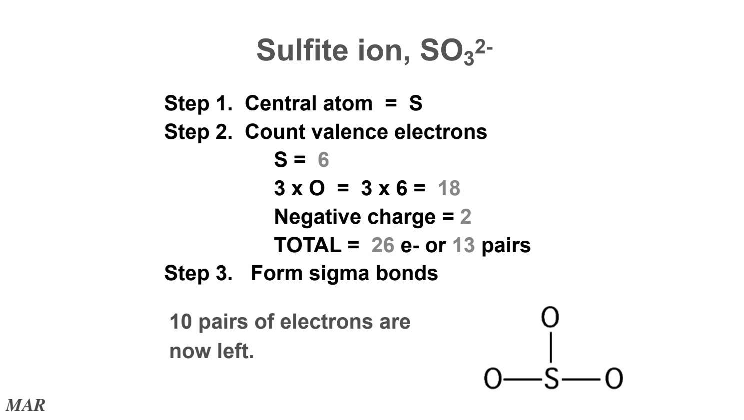
See Building Lewis Structures handout



Sulfite ion, SO_3^{2-}

Step 1. Central atom = S

Step 2. Count valence electrons

S = 6

3 x O = 3 x 6 = 18

Negative charge = 2

TOTAL = 26 e- or 13 pairs

Step 3. Form sigma bonds

10 pairs of electrons are now left.

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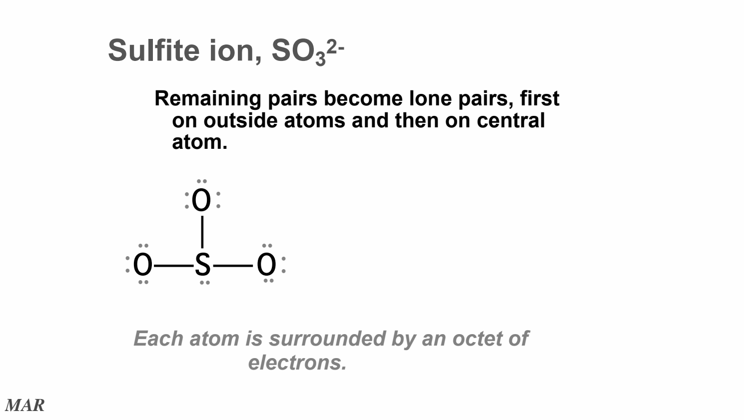

Sulfite ion, SO_3^{2-}

Remaining pairs become lone pairs, first on outside atoms and then on central atom.

Each atom is surrounded by an octet of electrons.

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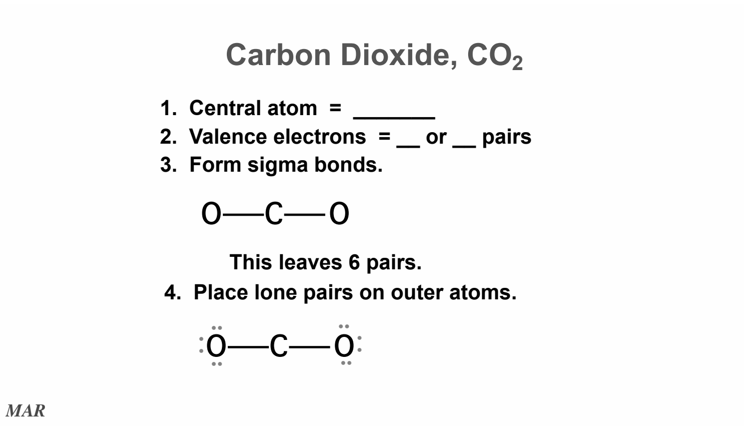

Carbon Dioxide, CO_2

1. **Central atom = _______**
2. **Valence electrons = __ or __ pairs**
3. **Form sigma bonds.**

This leaves 6 pairs.

4. **Place lone pairs on outer atoms.**

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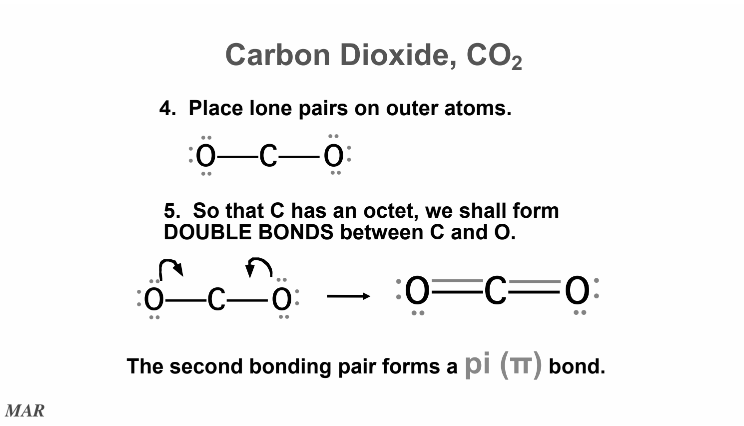

Carbon Dioxide, CO_2

4. **Place lone pairs on outer atoms.**

5. **So that C has an octet, we shall form DOUBLE BONDS between C and O.**

The second bonding pair forms a pi (π) bond.

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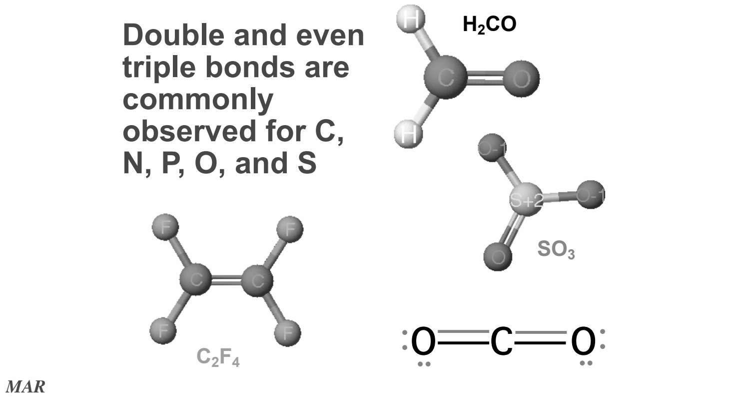

Double and even triple bonds are commonly observed for C, N, P, O, and S

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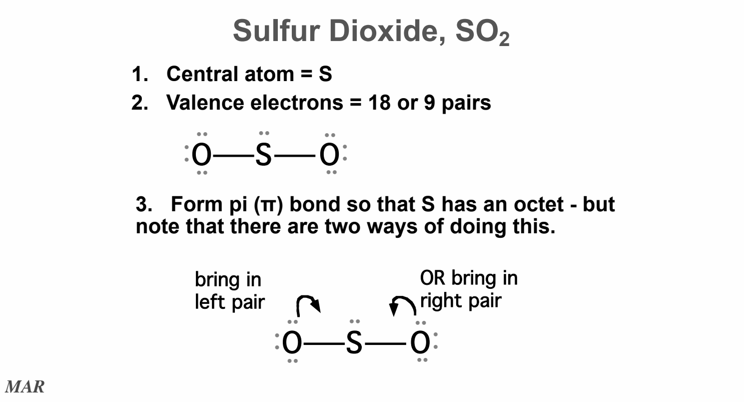

Sulfur Dioxide, SO_2

1. **Central atom = S**
2. **Valence electrons = 18 or 9 pairs**

3. **Form pi (π) bond so that S has an octet - but note that there are two ways of doing this.**

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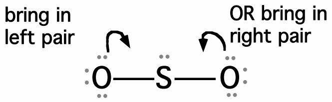
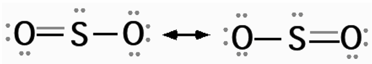
Sulfur Dioxide, SO_2



This leads to the following structures.



These equivalent structures are called RESONANCE STRUCTURES. The true electronic structure is a HYBRID of the two.

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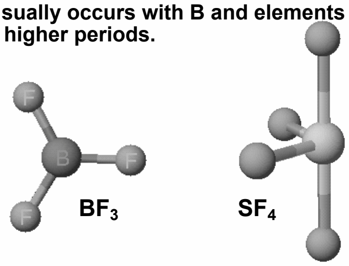
Violations of the Octet Rule

Usually occurs with B and elements of higher periods.



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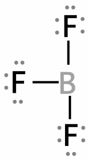
Boron Trifluoride

Central atom =

Valence electrons =
or electron pairs =

Assemble dot structure



The B atom has a share in only 6 electrons (or 3 pairs). B atom in many molecules is electron deficient.

Also common for Al and Be

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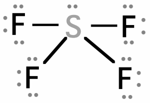
Sulfur Tetrafluoride, SF_4

Central atom =

Valence electrons = ___ or ___ pairs.

Form sigma bonds and distribute electron pairs.



5 pairs around the S atom. A common occurrence outside the 2nd period.

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Odd # of electrons: NO_2

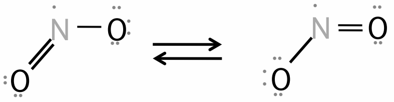
Paramagnetic compounds & free radicals

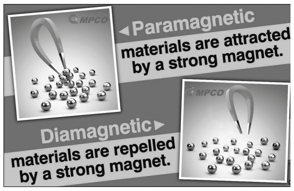
For NO_2, central atom = ___

Valence electrons = ___ or ___ pairs.

Odd e$^-$ occupies its own "space"

Form sigma bonds and distribute electron pairs.





Paramagnetic substances often more reactive than diamagnetic substances

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Formal Atom Charges

Atoms in molecules often bear a charge (+ or -).

The *predominant resonance structure* of a molecule is the one with *charges as close to 0* as possible.

Formal charge = Group number - $^1/_2$ (number bonding electrons) - (number lone pair electrons (lpe)),

or

FC = GN - bonds - lpe

Sum of all formal charges in a molecule **must equal ionic charge**

See Guide to Formal Charges

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Formal Atom Charges

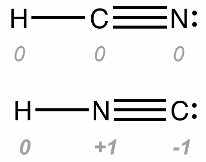
Formal charge = Group number - $^1/_2$ (number bonding electrons) - (number lone pair electrons)

or

FC = GN - bonds - lpe

Formal Charges In Isomers

more stable structure



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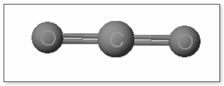
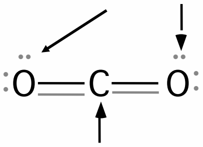
Carbon Dioxide, CO_2



FC = GN - bonds - lpe

6 - (1/2)(4) - 4 = 0



4 - (1/2)(8) - 0 = 0

*Sum of formal charges = = 0 + 0 + 0 = **0** which equals ionic charge on molecule*

This is a good structure!

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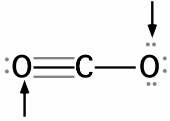
Formal Charge Comparison with CO_2

FC = GN - bonds - lpe

6 - (1/2)(2) - 6 = -1



6 - (1/2)(6) - 2 = +1

C atom formal charge is still 0.

***Sum** of formal charges = = -1 + 0 + 1 = **0** which equals ionic charge on molecule*

This is a valid resonance form, but not as good as previous slide

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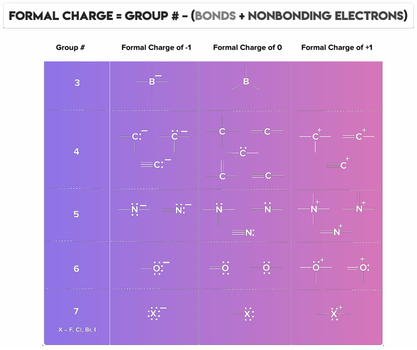
Formal Charge



FORMAL CHARGE = GROUP # – (BONDS + NONBONDING ELECTRONS)

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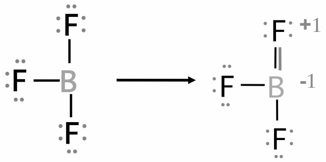
Boron Trifluoride, BF_3



What if we form a B-F double bond to satisfy the B atom octet?

F *never* makes double bonds!

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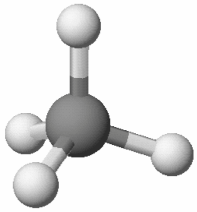
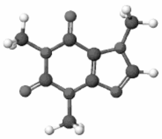
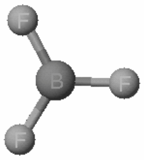
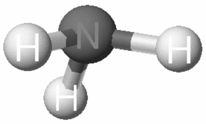
Molecular Geometry









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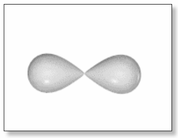
MOLECULAR GEOMETRY

VSEPR

Valence Shell Electron Pair Repulsion theory.

Most important factor in determining geometry is relative repulsion between electron pairs.



Molecule adopts the shape that minimizes the electron pair repulsions.

See Geometry and Polarity Guide

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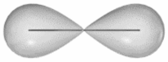
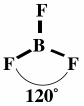
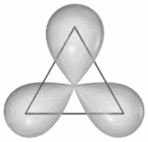
No. of e- Pairs Around Central Atom	Example	Geometry
2	F—Be—F 180˚	linear

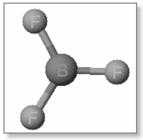


MAR *See Geometry and Polarity Guide*

VSEPR

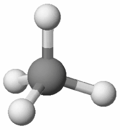
No. of e- Pairs Around Central Atom	Example	Electron Pair Geometry
2	F—Be—F 180˚	linear
3		trigonal planar

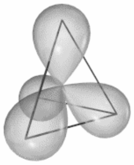


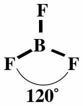
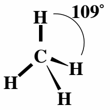


MAR *See Geometry and Polarity Guide*

VSEPR





No. of e- Pairs Around Central Atom	Example	Electron Pair Geometry
2	F—Be—F 180˚	linear
3		trigonal planar
4		tetrahedral

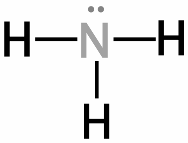
MAR *See Geometry and Polarity Guide*

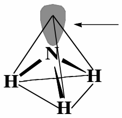
VSEPR

Structure Determination by VSEPR

Ammonia, NH_3

1. **Draw electron dot structure**
2. **Count BPs and LPs = 4**
3. **The 4 electron pairs are at the corners of a tetrahedron.**





lone pair of electrons in tetrahedral position

The ELECTRON PAIR GEOMETRY is tetrahedral.

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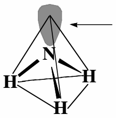
Structure Determination by VSEPR

Ammonia, NH_3

The electron pair geometry is tetrahedral.

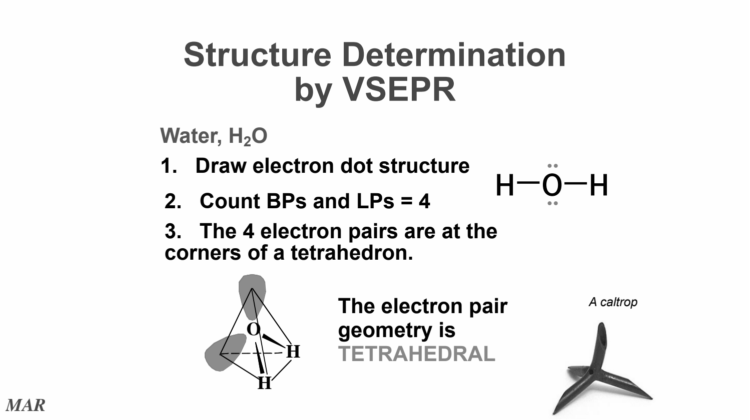


lone pair of electrons in tetrahedral position

The MOLECULAR GEOMETRY - the positions of the atoms - is TRIGONAL PYRAMID

See Geometry and Polarity Guide

MAR



Structure Determination by VSEPR

Water, H_2O

1. Draw electron dot structure
2. Count BPs and LPs = 4
3. The 4 electron pairs are at the corners of a tetrahedron.

H—Ö—H

The electron pair geometry is TETRAHEDRAL

A caltrop

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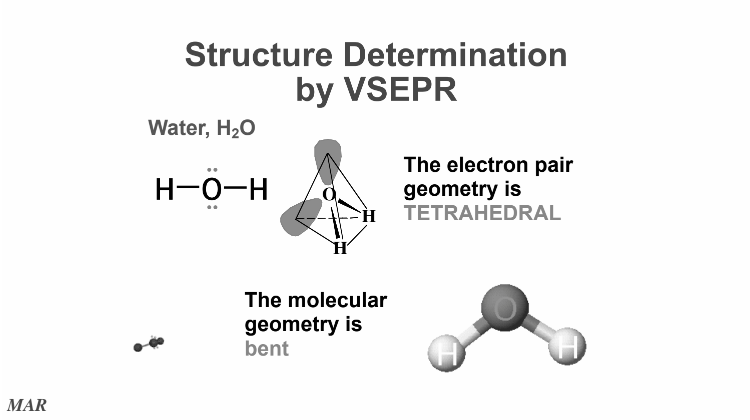

Structure Determination by VSEPR

Water, H_2O

H—Ö—H

The electron pair geometry is TETRAHEDRAL

The molecular geometry is bent

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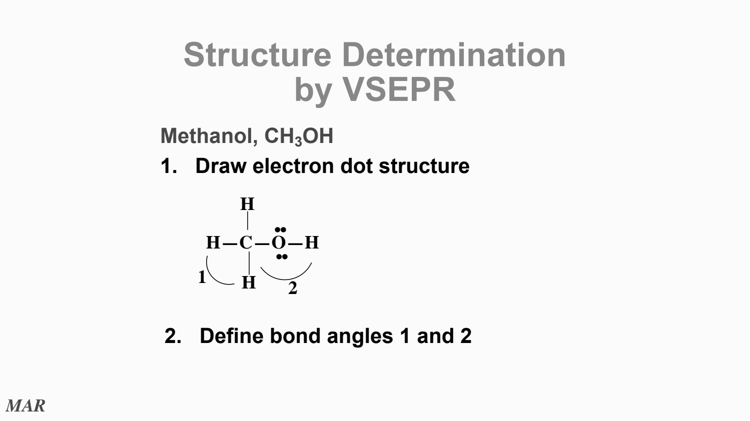

Structure Determination by VSEPR

Methanol, CH_3OH

1. Draw electron dot structure

2. Define bond angles 1 and 2

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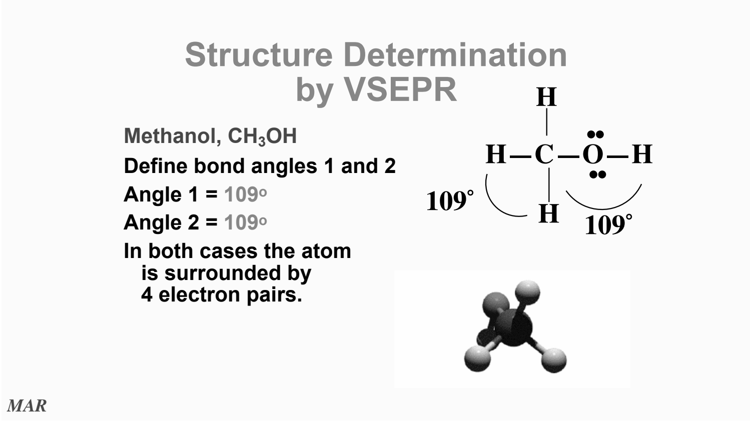

Structure Determination by VSEPR

Methanol, CH_3OH

Define bond angles 1 and 2

Angle 1 = 109°

Angle 2 = 109°

In both cases the atom is surrounded by 4 electron pairs.

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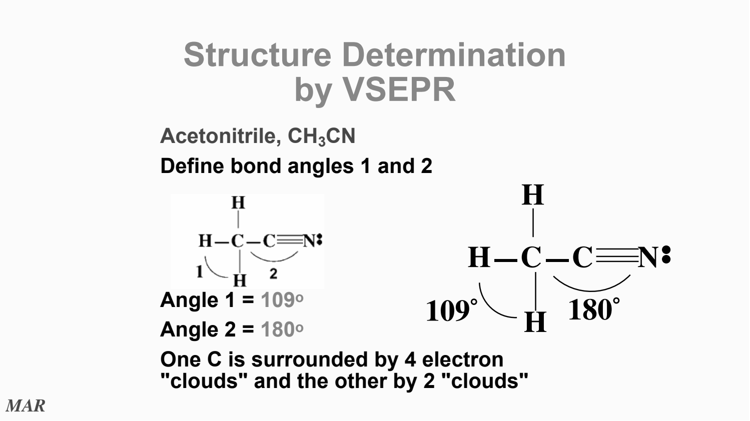

Structure Determination by VSEPR

Acetonitrile, CH_3CN

Define bond angles 1 and 2

Angle 1 = 109°

Angle 2 = 180°

One C is surrounded by 4 electron "clouds" and the other by 2 "clouds"

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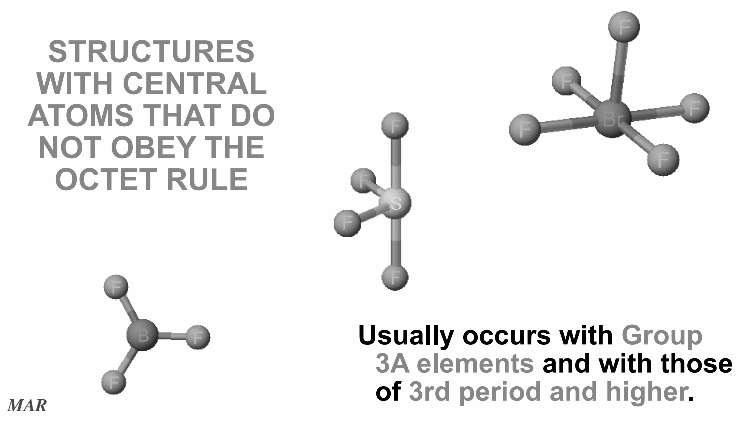

STRUCTURES WITH CENTRAL ATOMS THAT DO NOT OBEY THE OCTET RULE

Usually occurs with Group 3A elements and with those of 3rd period and higher.

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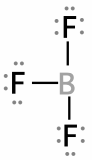
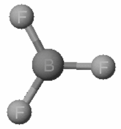
Compounds with 3 Pairs Around the Central Atom

The B atom is surrounded by only 3 electron pairs.

Bond angles are 120°



Geometry described as planar trigonal or trigonal planar



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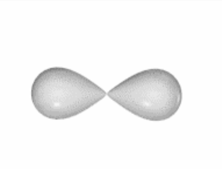
Compounds with 5 or 6 Pairs Around the Central Atom



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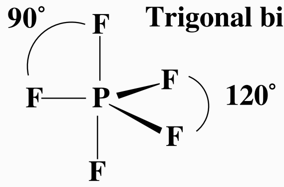
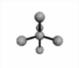
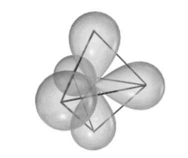
Compounds with 5 or 6 Pairs Around the Central Atom

Trigonal bipyramid



5 electron pairs





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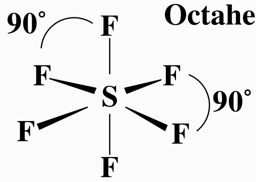
See Geometry and Polarity Guide

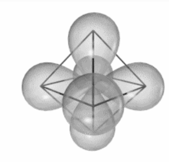
Compounds with 5 or 6 Pairs Around the Central Atom

Octahedron



6 electron pairs

The game "Jacks"



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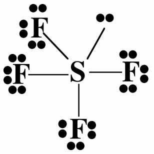
See Geometry and Polarity Guide

Sulfur Tetrafluoride, SF_4

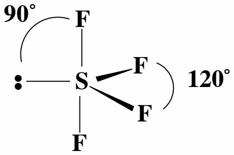
Number of valence electrons = 34

Central atom = S

Dot structure



Electron pair geometry = trigonal bipyramid (because there are 5 pairs around the S)



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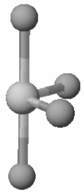
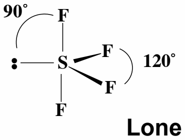
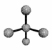
Sulfur Tetrafluoride, SF_4

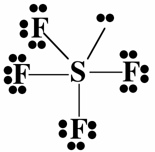
Molecular geometry = seesaw





Lone pair is in the equator because it requires more room.





Unshared electron pairs ("lone pairs") take up more volume than shared electron pairs ("bonding pairs")

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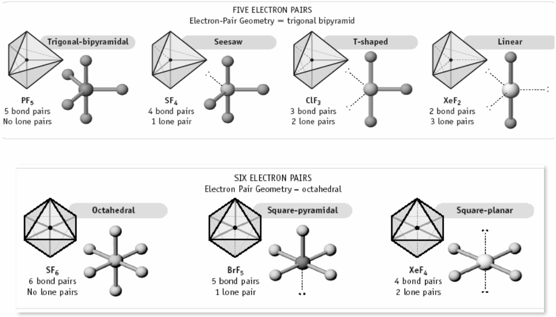
Other Molecular Geometries



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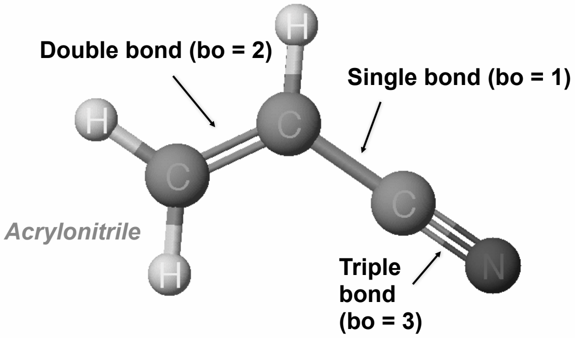
Bond Order

of bonds between a pair of atoms



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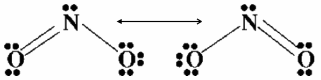
Bond Order

Fractional bond orders **occur in molecules with resonance structures.**

Consider NO_2^-



Bond order = $\dfrac{\text{Total \# of e- pairs used for a type of bond}}{\text{Total \# of bonds of that type}}$

Bond order = $\dfrac{\text{3 e- pairs in N — O bonds}}{\text{2 N — O bonds}}$

The N-O bond order = 1.5

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Bond Order

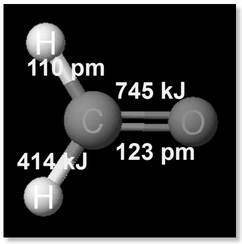
Bond order is related to two important bond properties:

(a) bond length

(b) bond energy

Bond length ***is*** inversely proportional ***to bond order***

Bond energy ***is*** proportional ***to bond order***



Bond lengths measured in pm (1 pm = 10^{-12} m) or Angstroms (1 Å = 10^{-10} m)

Bond energies measured in kJ (1 kJ = 10^{3} J)

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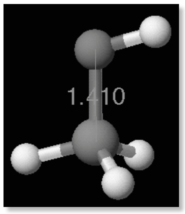
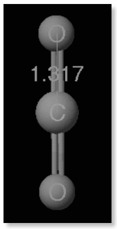
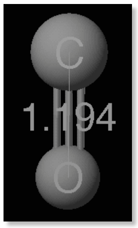
Bond Length







Bond length depends on bond order

As bond order increases, bond length decreases

Bond length is inversely proportional to bond order

Bond distances measured in Angstroms (1 Å = 10^{-10} m)

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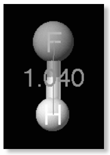
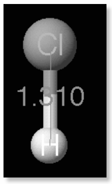
Bond Length

Bond length, the distance between two nuclei, also depends on the size of bonded atoms**.**

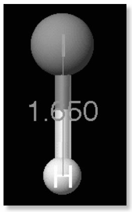
Bond length is inversely proportional to bond order



H-F



H-Cl



H-I

Bond distances measured in Angstroms (1 Å = 10^{-10} m)

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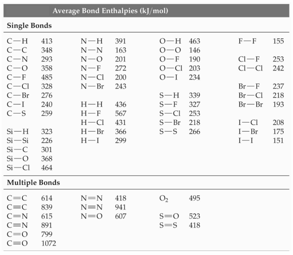
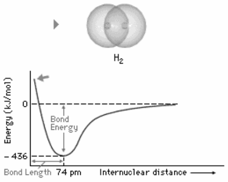
Bond Energy is measured by the energy required to break a bond

The **GREATER** the number of bonds (*bond order*) the **LARGER** the bond energy and the **SHORTER** the bond.



Average Bond Enthalpies (kJ/mol)

Single Bonds							
C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						
Si—Cl	464						
Multiple Bonds							
C=C	614	N=N	418	O_2	495		
C≡C	839	N≡N	941				
C=N	615	N=O	607	S=O	523		
C≡N	891			S=S	418		
C=O	799						
C≡O	1072						



Bond Energy

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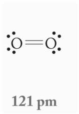
Bond Energy



Bond order =2
498 kJ/mol

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Using Bond Energies

Estimate the energy of the reaction:

H-H + Cl-Cl → 2 H-Cl

Net energy = ΔH_{rxn} = energy required to break bonds - energy evolved when bonds formed

H-H = 436 kJ/mol
Cl-Cl = 243 kJ/mol
H-Cl = 431 kJ/mol

ΔH = bonds broken - bonds formed

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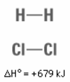
Using Bond Energies

ΔH = bonds broken - bonds formed

Estimate the energy of the reaction: H-H + Cl-Cl ----> 2 H-Cl

H-H = 436 kJ/mol
Cl-Cl = 243 kJ/mol
H-Cl = 431 kJ/mol



"Bonds broken" *or* "Reactant bonds":
H-H + Cl-Cl bond energies = 436 kJ + 243 kJ = 679 kJ

"Bonds formed" *or* "Product bonds":
2 mol H-Cl bond energies = 2 x 431 kJ = 862 kJ

ΔH = bonds broken - bonds formed
ΔH = 679 kJ - 862 kJ = -183 kJ

***exothermic*!**

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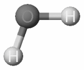
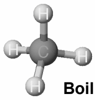
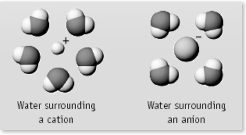
Molecular Polarity

Boiling point = 100 ˚C





Boiling point = -161 ˚C



Why do ionic compounds dissolve in water?

Why do water and methane differ so much in their boiling points?

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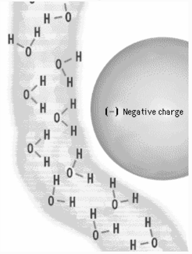
Molecular Polarity



Water molecules are attracted to balloons that have a static electric charge

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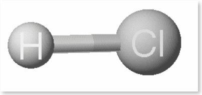
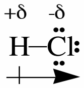
Bond Polarity



HCl is POLAR because it has a positive end and a negative end (dipoles).



Cl has a greater share in bonding electrons than does H.

Cl has slight negative charge (-δ) and H has slight positive charge (+δ)

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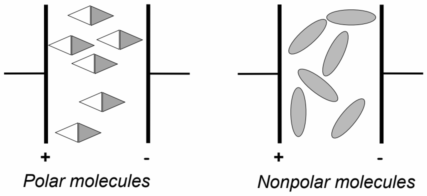
Bond Polarity

Dipole moment, μ, can measure dipole strength by placing molecules in electrical field. Polar molecules will align when the field is on. Nonpolar molecules will not.



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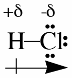

Bond Polarity

Due to polarity, the H-Cl bond energy is GREATER than expected for a "pure" covalent bond.

BOND	ENERGY
"pure" bond	339 kJ/mol calc'd
real bond	432 kJ/mol measured

Difference = 92 kJ. This difference is proportional to the difference in ELECTRONEGATIVITY, χ.

See Polarity Guide

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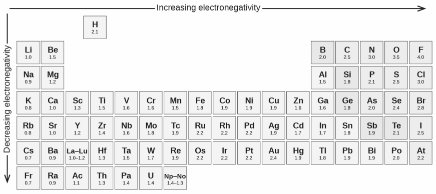
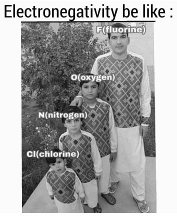
Electronegativity, χ

χ is a measure of the ability of an atom in a molecule to attract electrons to itself.





Electronegativities tend to increase up and to the right on the periodic table

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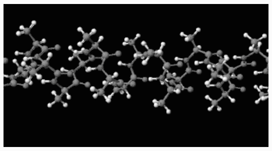
Linus Pauling, 1901-1994



The only person to receive <u>two</u> unshared Nobel prizes (for Peace *and* Chemistry)

Chemistry areas: bonding, electronegativity, protein structure

A great Oregonian *and* a great Scientist

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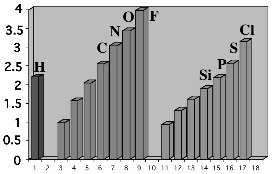

Electronegativity, χ

F has maximum χ.

Atom with lowest χ is the center atom in most molecules.

Relative values of χ determine BOND POLARITY (and point of attack on a molecule).

We are using "traditional" electronegativity values, but a new system has been introduced (January 2019)

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Bond Polarity

Which bond is more polar (or DIPOLAR)?

	O-H	O-F
χ	3.5 - 2.1	3.5 - 4.0
Δχ	1.4	0.5

∴ OH is *more polar* than OF

and polarity is "reversed"

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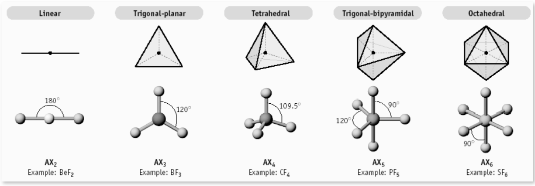
Molecular Polarity

Molecules will be polar *if*

a) bonds are polar

AND

b) the molecule is NOT "symmetric"



All above are symmetric and NOT polar (nonpolar)

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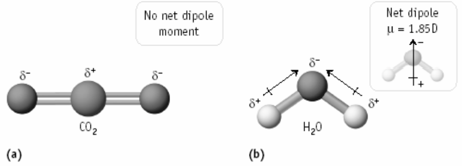
Polar or Nonpolar?

Compare CO_2 and H_2O. Which one is polar?



Polar bonds, nonpolar molecule

Polar bonds, polar molecule

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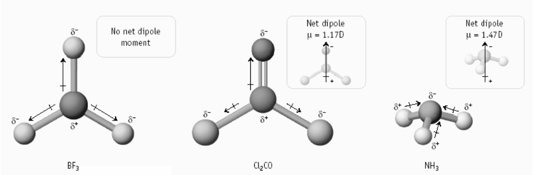
Polar or Nonpolar?

Consider AB_3 molecules: BF_3, Cl_2CO, and NH_3.



Polar and nonpolar molecules of the type AB_3.

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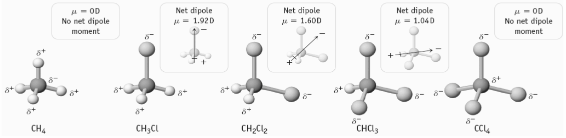
CH_4 through CCl_4 Polarity



Methane (CH_4) and carbon tetrachloride (CCl_4) are symmetrical and NOT polar.

All other compounds asymmetrical and polar.

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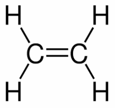
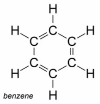
More on Molecular Polarity



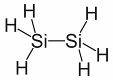
ethene

$H-C\equiv C-H$

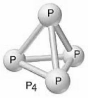
acetylene (ethyne)



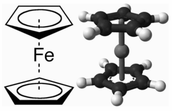
benzene



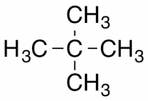
disilane



white phosphorus



ferrocene



$$H_3C-\overset{\displaystyle CH_3}{\underset{\displaystyle CH_3}{\overset{|}{\underset{|}{C}}}}-CH_3$$

2,2-dimethylpropane

***All* of these molecules are nonpolar due to their symmetry.**

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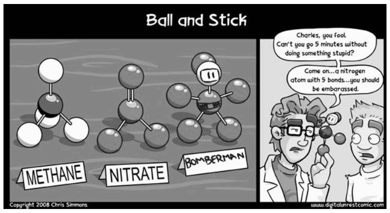
End of Chapter 7



See:

- *Chapter Seven Study Guide*
- *Chapter Seven Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*

Important Equations, Constants, and Handouts from this Chapter:

- *know how to determine if* **ionic, covalent** *or* **metallic bonds** *are present*
- **ionic bond strength** *determined by* **Coulomb's Law**
- **# valence electrons = group number** *(US periodic table!)*
- *know the relationship between* **bond order, bond length** *and* **bond energy**
- *see* **Geometry and Polarity Guide** *and* **Bond Enthalpies and Electronegativities** *(handouts)*

Formal Charge = Group Number - bonds - lone pair electrons

FC = GN - bonds - lpe

ΔH_{rxn} **= bonds broken - bonds formed**

Lewis Structures / VSEPR: bonding pairs, lone pairs, valence electrons, core electrons, total electrons, sigma bond, pi bond, VSEPR names (EPG & MG), formal charge, bond angles, polar, nonpolar, paramagnetic, diamagnetic, resonance structures, isomers

$$\text{bond order (resonance)} = \frac{\#\ of\ e^-\ pairs\ used\ for\ a\ type\ of\ bond}{\#\ of\ bonds\ of\ that\ type}$$

End of Chapter Problems: *Test Yourself*

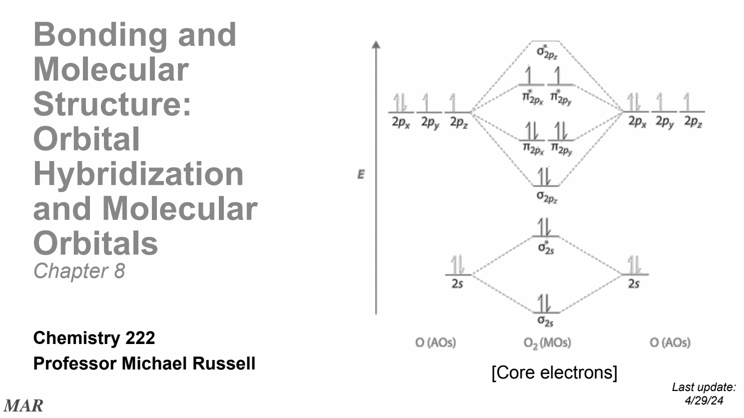
See practice problem set #1 and self quizzes for ***Lewis Structure / VSEPR*** *examples and practice*

1. Which of the following elements are capable of forming compounds in which the indicated atom has more than four valence electron pairs?
   N, As, C, O, Br, Be, S, Se
2. Which compound in each of the following pairs should require the higher temperature to melt?
   a. KBr or CsBr
   b. SrS or CaS
   c. LiF or BeO
3. Describe the EPG and MG around N in NH_2Cl.
4. Describe the EPG and MG around Cl in ClF_5.
5. Describe the EPG and MG around Te in TeF_4.
6. Which molecules are polar and which are nonpolar? H_2O, NH_3, CO_2, ClF, CCl_4
7. Give the bond order for each bond in the following molecules or ions: CH_2O, CO_2, NO_2^{+1}, CH_4
8. Oxygen difluoride is quite reactive with water, giving oxygen and HF:
   $OF_2(g) + H_2O(g) \rightarrow O_2(g) + 2\ HF(g) \quad \Delta H°_{rxn} = -318$ kJ
   Using bond energies, calculate the bond dissociation energy of the O-F bond in OF_2.

End of Chapter Problems: *Answers*

1. As, Br, S and Se
2. a. KBr b. CaS c. BeO
3. tetrahedral and trigonal pyramid
4. octahedral and square pyramid
5. trigonal bipyramid and seesaw
6. polar: H_2O, NH_3, ClF nonpolar: CO_2, CCl_4
7. **CH_2O** (2xBO=1 (C-H), 1xBO=2 (C=O)) , **CO_2** (2xBO=2 (C-O)), **NO_2^{+1}** (2xBO=2 (N-O)), **CH_4** (4xBO=1 (C-H))
8. D(O-F) = 195 kJ/mol

See practice problem set #1 and self quizzes for ***Lewis Structure / VSEPR*** *examples and practice*



Bonding and Molecular Structure: Orbital Hybridization and Molecular Orbitals

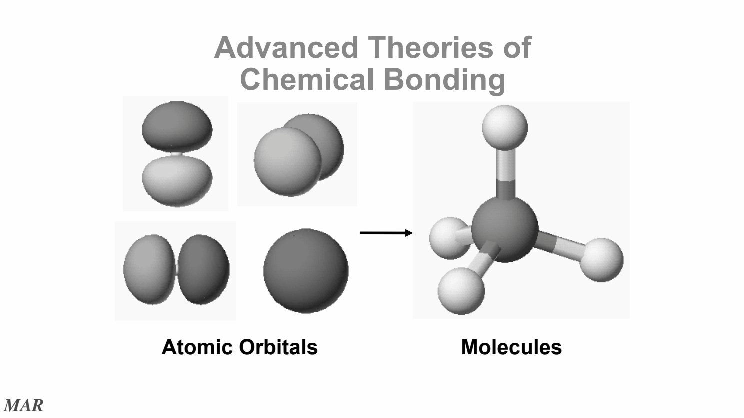
Chapter 8

Chemistry 222
Professor Michael Russell

[Core electrons]

Last update: 4/29/24

MAR



Advanced Theories of Chemical Bonding

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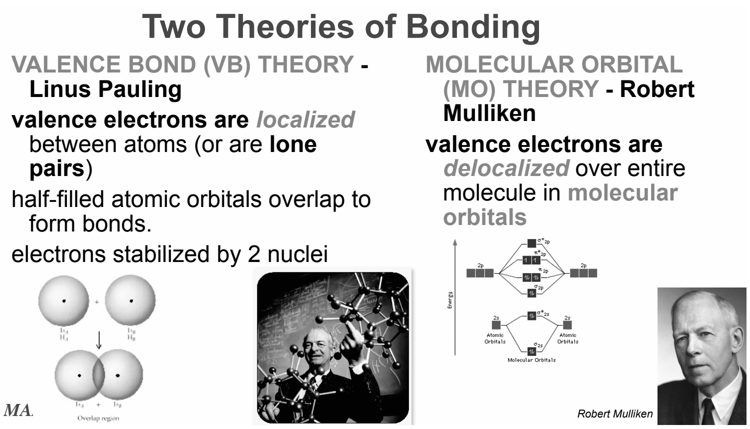


Two Theories of Bonding

VALENCE BOND (VB) THEORY - **Linus Pauling**

valence electrons are *localized* between atoms (or are **lone pairs**)

half-filled atomic orbitals overlap to form bonds.

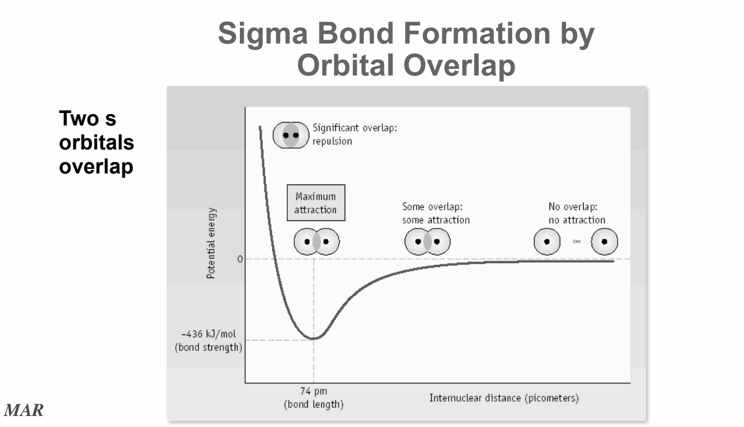
electrons stabilized by 2 nuclei

MOLECULAR ORBITAL (MO) THEORY - **Robert Mulliken**

valence electrons are *delocalized* over entire molecule in molecular orbitals

Robert Mulliken

MA.



Sigma Bond Formation by Orbital Overlap

Two s orbitals overlap

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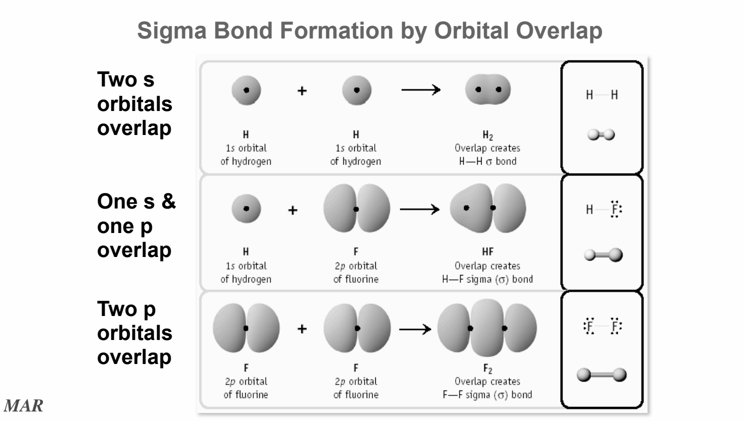


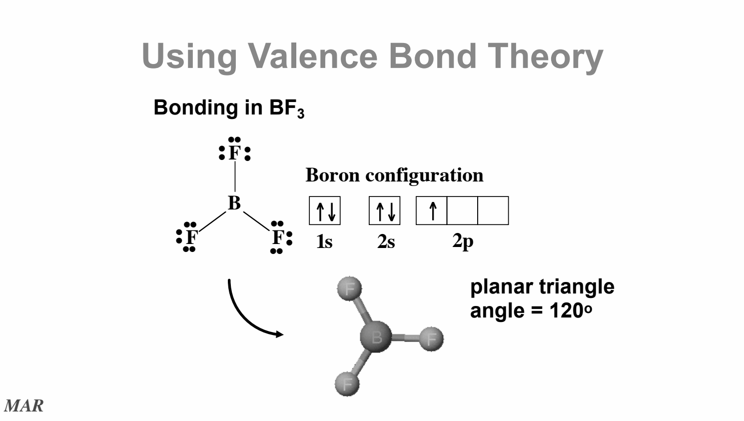
Sigma Bond Formation by Orbital Overlap

Two s orbitals overlap

One s & one p overlap

Two p orbitals overlap

MAR



Using Valence Bond Theory

Bonding in BF_3

Boron configuration

planar triangle
angle = 120°

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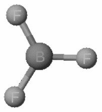

Bonding in BF_3

How to account for 3 bonds 120° apart using a spherical s orbital and p orbitals that are 90° apart?

Pauling said to modify VB approach with ORBITAL HYBRIDIZATION

- mix available orbitals to form a new set of orbitals - HYBRID ORBITALS - that will give the maximum overlap in the correct geometry.

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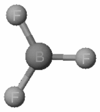
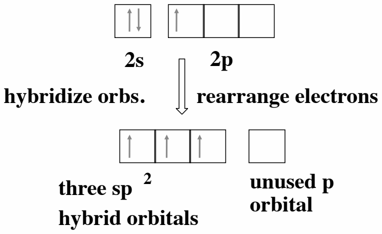


Bonding in BF_3



2s 2p

hybridize orbs. rearrange electrons

three sp^2 hybrid orbitals

unused p orbital

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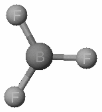
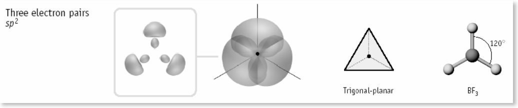


Bonding in BF_3

The three hybrid orbitals are made from 1 s orbital and 2 p orbitals create 3 sp^2 hybrids.



Now we have 3, half-filled HYBRID orbitals that can be used to form *planar* B-F sigma bonds.

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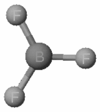
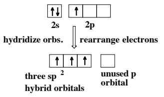
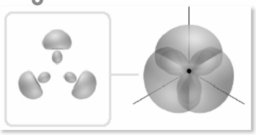
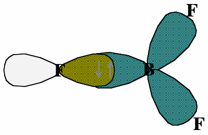


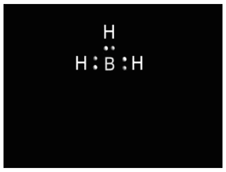
Bonding in BF_3





An orbital from each F overlaps one of the sp^2 hybrids to form a B-F σ bond.





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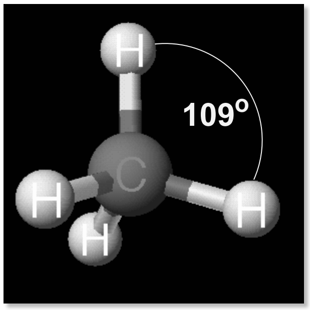
Bonding in CH_4

How do we account for 4 C-H sigma bonds 109° apart?

Need to use 4 atomic orbitals - s, p_x, p_y, and p_z - to form 4 new hybrid orbitals pointing in the correct direction.



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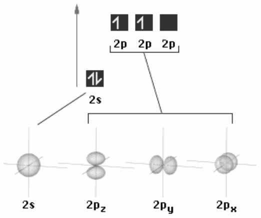
Bonding in a Tetrahedron - Formation of Hybrid Atomic Orbitals



4 C atom orbitals hybridize to form four equivalent sp^3 hybrid atomic orbitals.

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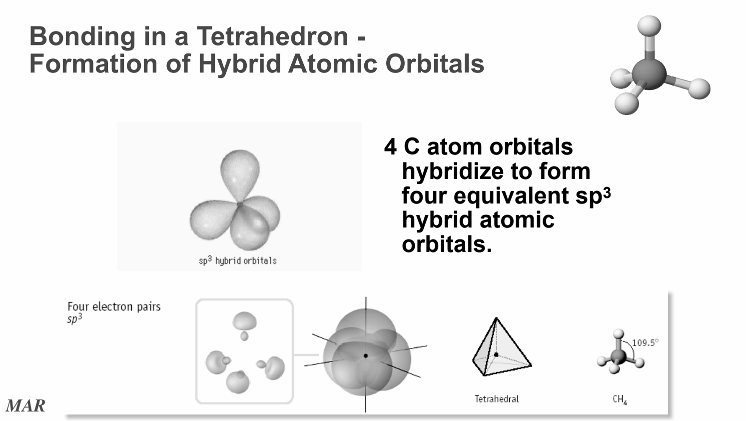

Bonding in a Tetrahedron - Formation of Hybrid Atomic Orbitals

4 C atom orbitals hybridize to form four equivalent sp^3 hybrid atomic orbitals.

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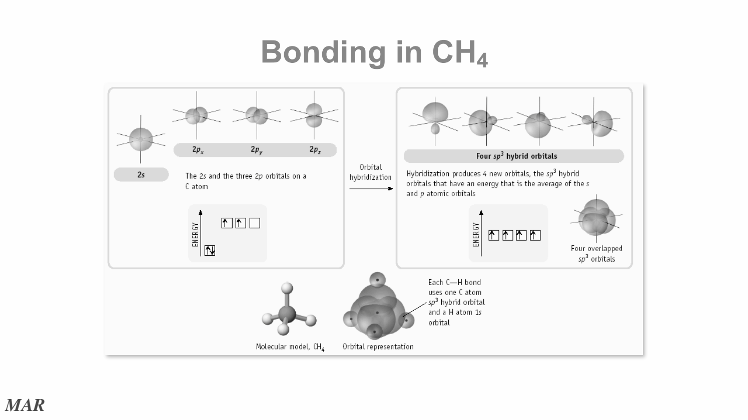

Bonding in CH_4

The 2s and the three 2p orbitals on a C atom

Hybridization produces 4 new orbitals, the sp^3 hybrid orbitals that have an energy that is the average of the s and p atomic orbitals

Each C—H bond uses one C atom sp^3 hybrid orbital and a H atom 1s orbital

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Orbital Hybridization

Bonds	EPG	Hybrid	REMAINING p orbs?
2	linear	sp	2 p
3	trigonal planar	sp^2	1 p
4	tetrahedral	sp^3	none
5	trigonal bipyramid	sp^3d	---
6	octahedral	sp^3d^2	---

see: VSEPR Guide

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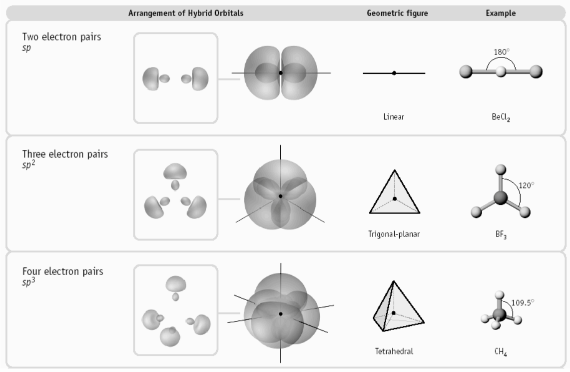

Arrangement of Hybrid Orbitals	Geometric figure	Example
Two electron pairs *sp*	Linear	$BeCl_2$ (180°)
Three electron pairs sp^2	Trigonal-planar	BF_3 (120°)
Four electron pairs sp^3	Tetrahedral	CH_4 (109.5°)

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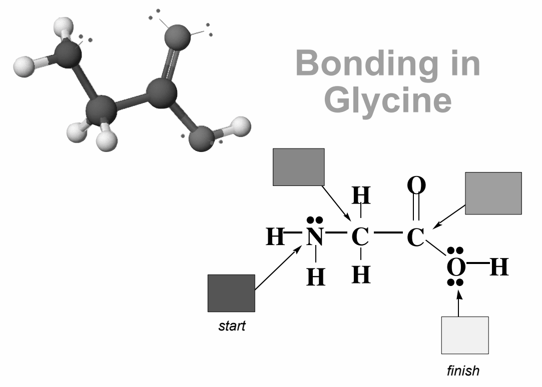

Bonding in Glycine

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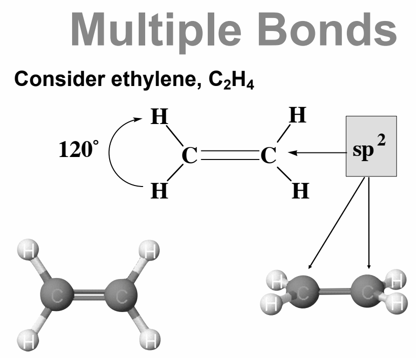

Multiple Bonds

Consider ethylene, C_2H_4

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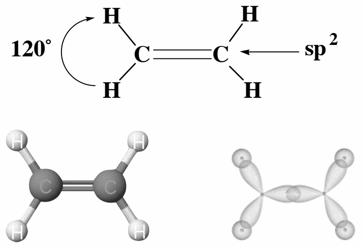
Sigma Bonds in C_2H_4



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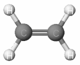
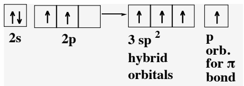
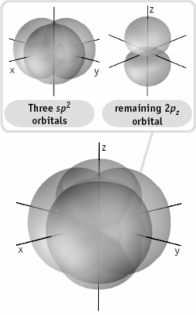


π Bonding in C_2H_4

The unused p orbital on each C atom contains an electron and this p orbital overlaps the p orbital on the neighboring atom to form the π bond.





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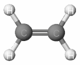
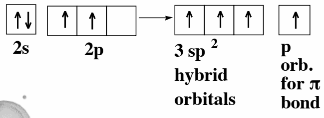
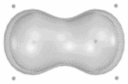


π Bonding in C_2H_4

The unused p orbital on each C atom contains an electron, and this p orbital overlaps the p orbital on the neighboring atom to form the π bond.





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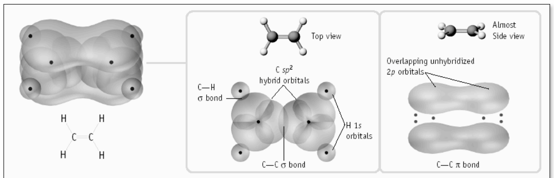
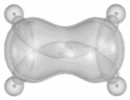
Multiple Bonding in C_2H_4



(a) Lewis structure and bonding of ethylene, C_2H_4.

(b) The C—H σ bonds are formed by overlap of C atom sp^2 hybrid orbitals with H atom 1*s* orbitals. The σ bond between C atoms arises from overlap of sp^2 orbitals.

(c) The carbon-carbon π bond is formed by overlap of an unhybridized 2*p* orbital on each atom. Note the lack of electron density along the C—C bond axis.



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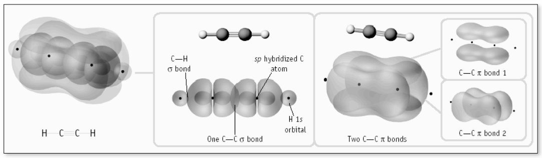
σ and π Bonding in C_2H_2



C_2H_2 has a *triple bond*

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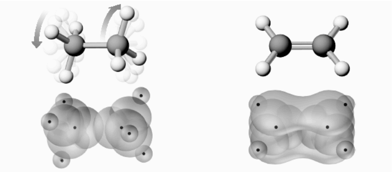
Consequences of Multiple Bonding

There is restricted rotation around C=C bond.



(a) Free rotation can occur around the axis of a single (σ) bond.

(b) In contrast, rotation is severely restricted around double bonds because doing so would break the π bond, a process generally requiring a great deal of energy.

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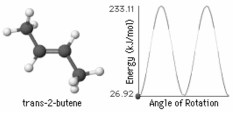
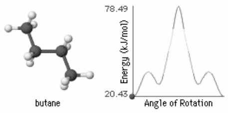
Consequences of Multiple Bonding

Restricted rotation around C=C bond.



trans-but-2-ene



n-butane

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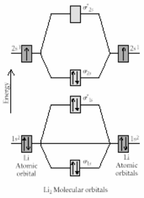
Molecular Orbital Theory

- **Accounts for paramagnetism, color, bonding**
- **Atomic orbitals delocalize into molecular orbitals**
- **Bonding, Antibonding and Nonbonding orbitals**
- **Quite complicated, need computers; we will only look at diatomics (*2 atom systems*) from the first and second periods *only***



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Advantages of MO Theory

Dioxygen should be electron paired (*diamagnetic*) by VB Theory, but dioxygen is actually *paramagnetic*.
MO Theory accounts for paramagnetism of O_2

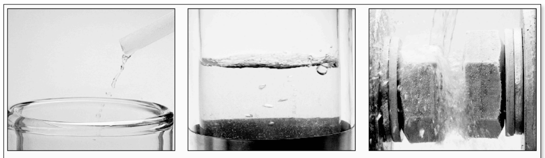

Figure 10.16 Liquid oxygen. Oxygen gas condenses to a liquid at −183 °C (*left*). Notice that liquid oxygen is very pale blue (*middle*). Oxygen in the liquid state is paramagnetic and clings to the poles of a magnet (*right*). *(Charles D. Winters)*

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Four Principles of MO Theory

Principle #1:
Number of Molecular Orbitals = Number of Atomic Orbitals

Two 1s orbitals from two hydrogen atoms create two molecular orbitals in H_2

Two 1s orbitals and two 2s orbitals from two lithium atoms create four molecular orbitals in Li_2

See Four Principles of MO Handout

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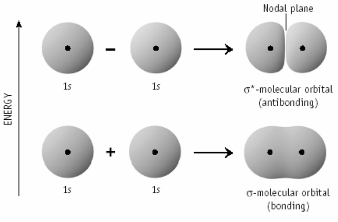
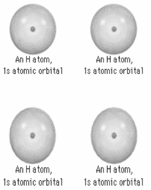
Molecular Orbital Type

When two atomic 1s H orbitals combine, a bonding (σ) and antibonding (σ*) molecular orbital forms



antibonding σ*



bonding σ

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Four Principles of MO Theory

Principle #2:
Bonding MO lower in energy than the parent orbital
Antibonding MO higher in energy than the parent orbital

Principle #3:
Electrons of molecule assigned to successively higher MOs

Use ***Pauli Exclusion Principle*** and ***Hund's Rule*** when assigning electrons

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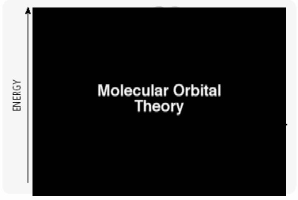
MO Diagram for H_2



Antibonding MO

Atomic Orbitals

Bonding MO

Two 1s electrons from two H atoms occupy the σ orbital in H_2

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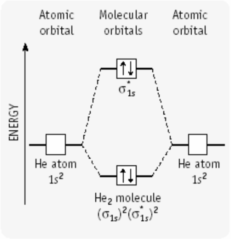
MO Diagram for He_2



Antibonding MO

Atomic Orbitals

Bonding MO

Two 1s electrons in σ,
Two 1s electrons in σ^*

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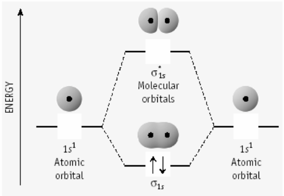
Bond Order in MO Theory

Bond Order = $^1/_2$ (# bonding e^- - # antibonding e^-)



Bond Order **> 0,**
stable molecule
Bond Order **= 0** or **< 0,**
unstable molecule

In H_2,
Bond Order =
$^1/_2$ (2 - 0) = 1;
stable

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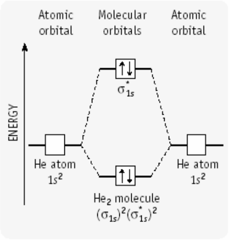
Bond Order in MO Theory

Bond Order = $^1/_2$ (# bonding e^- - # antibonding e^-)



Bond Order **> 0,**
stable molecule
Bond Order **= 0** or **< 0,**
unstable molecule

In He_2,
Bond Order =
$^1/_2$ (2 - 2) = 0;
unstable

∴ He_2 does not exist

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Four Principles of MO Theory

Principle #4:

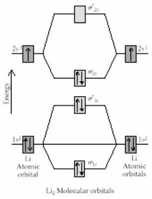
Atomic orbitals combine to give molecular orbitals only when the atomic orbitals are of similar energy

Similar energy = better overlap

1s + 1s = good MO
1s + 2s = poor MO
2s + 2s = good MO
2s + 2p = poor MO
3s + 2s = poor MO
... etc. ...

MAR

Example: Dilithium, Li_2



Note: no overlap between 1s and 2s

Bond Order = $^1/_2$ (4 - 2) = 1

Stable molecule

Would you expect Be_2 to exist? Why?

MAR

p orbitals and π bonds

Three possible p orbitals on each atom - *six total* p MO orbitals

Two p orbitals create 2 σ MO bonds

Four remaining p orbitals create 4 π MO bonds

Four p atomic orbitals create four π molecular orbitals,
π = bonding (2)
π* = antibonding (2)

... but there's a catch!

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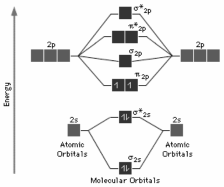
p orbitals and π bonds

For B, C and N,
π orbitals lower energy than σ orbital
π^* orbitals lower energy than σ^* orbital



***Example:* B_2**

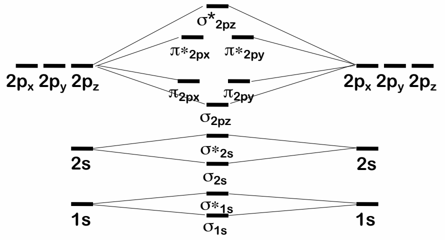
***Bond Order* = 1**

[Core electrons]

MAR *See MO Diagram (B_2 - N_2) Handout*

p orbitals and π bonds

For O, F and Ne,
σ orbital lower energy than π orbitals
π^* orbitals lower energy than σ^* orbital



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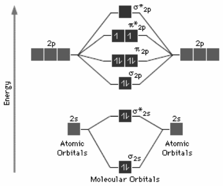
p orbitals and π bonds

For O, F and Ne,
σ orbital lower energy than π orbitals
π^* orbitals lower energy than σ^* orbital



***Example:* O_2**

***Bond Order* = 2**

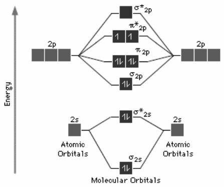
[Core electrons]

MAR *See MO Diagram (B_2 - N_2) Handout*

Paramagnetism

Paramagnetism exists when unpaired electrons in MO diagram

O_2 is paramagnetic;
unpaired electrons
in two π^* orbitals



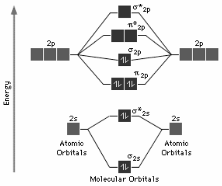
[Core electrons]

MAR

Paramagnetism

Paramagnetism exists when unpaired electrons in MO diagram

N_2 is diamagnetic;
all electrons paired



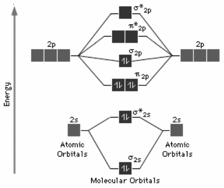
[Core electrons]

MAR

Molecular Orbital Notation

Used to abbreviate the MO diagrams

- **Ignore core electrons**
- **Write in order of increasing energy**



[Core electrons]

For N_2:

[core electrons]$(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^4(\sigma_{2p})^2$

Sigma and Pi Bonds

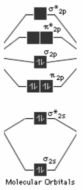
Determine sigma and pi bonds using:

$\#\ \sigma$ bonds = $^1/_2$ ($\#\ \sigma$ bonding e^- - $\#\ \sigma$ antibonding e^-)

$\#\ \pi$ bonds = $^1/_2$ ($\#\ \pi$ bonding e^- - $\#\ \pi$ antibonding e^-)

and

$\#\ \sigma$ bonds + $\#\ \pi$ bonds = bond order



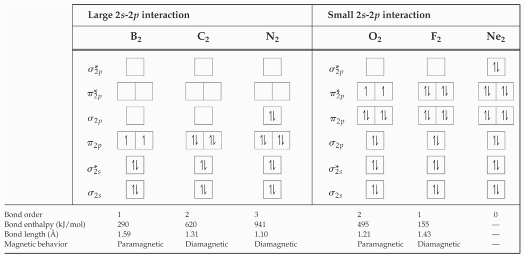
[Core electrons]

For N_2: **[core electrons]$(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^4(\sigma_{2p})^2$**

$\#\ \sigma$ bonds = $^1/_2$(4 - 2) = 1 σ bond
$\#\ \pi$ bonds = $^1/_2$(4 - 0) = 2 π bonds

bond order = $^1/_2$(8 - 2) = 3 = 1 σ + 2 π bonds

MO Diagram for Diatomics



	Large 2*s*-2*p* interaction				Small 2*s*-2*p* interaction		
	B_2	C_2	N_2		O_2	F_2	Ne_2
σ^*_{2p}	☐	☐	☐	σ^*_{2p}	☐	☐	↑↓
π^*_{2p}	☐☐	☐☐	☐☐	π^*_{2p}	↑ ↑	↑↓ ↑↓	↑↓ ↑↓
σ_{2p}	☐	☐	↑↓	π_{2p}	↑↓ ↑↓	↑↓ ↑↓	↑↓ ↑↓
π_{2p}	↑ ↑	↑↓ ↑↓	↑↓ ↑↓	σ_{2p}	↑↓	↑↓	↑↓
σ^*_{2s}	↑↓	↑↓	↑↓	σ^*_{2s}	↑↓	↑↓	↑↓
σ_{2s}	↑↓	↑↓	↑↓	σ_{2s}	↑↓	↑↓	↑↓
Bond order	1	2	3		2	1	0
Bond enthalpy (kJ/mol)	290	620	941		495	155	—
Bond length (Å)	1.59	1.31	1.10		1.21	1.43	—
Magnetic behavior	Paramagnetic	Diamagnetic	Diamagnetic		Paramagnetic	Diamagnetic	—

Note: all should have [core electrons]

Changes in MO diagrams due to
s-p mixing and/or electron repulsion

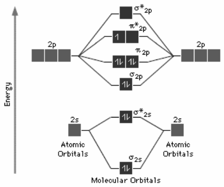
Ionic Diatomic Molecules

Predicting Ionic Diatomic MO diagrams simple
Use Hund and Pauli

Example: **O_2^+**

Remove electron from
π^*_{2p} orbital

Check bond order,
paramagnetism



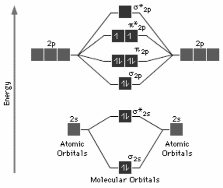
[Core electrons]

Ionic Diatomic Molecules

Predicting Ionic Diatomic MO diagrams simple
Use Hund and Pauli

Example: **O_2^-**

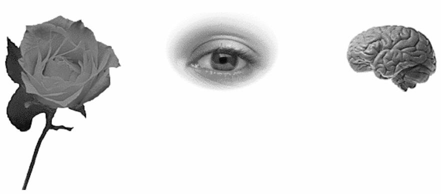
Using the O_2 diagram on
the right,
where would you place
the extra electron?
Is O_2^- more or less stable
than O_2? Why?



[Core electrons]

Application: Vision

Molecular Orbital Theory helps to describe the process of vision - *photochemistry*

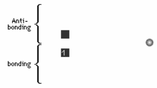


Application: Band Theory

In metallic bonding, electrons delocalized over metallic lattice - a sea of electrons

MO energies identical, excellent overlap

Helps explain conductivity, malleability, more



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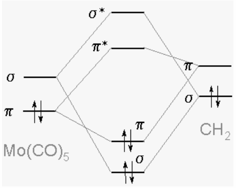
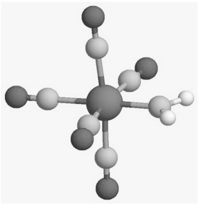
End of Chapter 8

See:

- *Chapter Eight Study Guide*
- *Chapter Eight Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*



MO Diagram for Mo(CO)$_5$=CH$_2$



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Important Equations, Constants, and Handouts from this Chapter:

- *the* **bond order, bond energy and bond length** *relationships still apply to both theories*
- *know the advantages and disadvantages of the Valence Bond and Molecular Orbital theories*
- *see the* Geometry and Polarity Guide *and the two* Molecular Orbital Theory *diagrams* (NBC *and* FONe) *(handouts)*

Valence Bond / Hybridization Theory: *types of* **hybridization (sp, sp^2, etc.), sigma and pi bonds**

Molecular Orbital Theory: **bonding** *and* **antibonding orbitals, sigma bonds** *and* **pi bonds, paramagnetic** *and* **diamagnetic,** *the* **"NBC" vs. "FONe"** *diagrams*

bond order (MO theory) = $\frac{1}{2}(\#\ bonding\ e^- - \#\ antibonding\ e^-)$

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End of Chapter Problems: ***Test Yourself***

Be sure to view practice problem set #2 and self quizzes for
***MO theory and Valence Bond theory** examples and practice*

1. Specify the electron-pair and molecular geometry for each of the following. Describe the hybrid orbital set used by the central atom in each molecule or ion.
   a. BBr_3 b. CO_2 c. CH_2Cl_2 d. XeF_4
2. Use MO theory to tell which has the largest bond order: C_2 or F_2. Are either species paramagnetic?
3. Use MO theory to speculate on the existence of dilithium.
4. Which compound is stronger by MO theory: Be_2 or B_2^{+1}.
5. Describe the hybridization change on carbon as methane (CH_4) is burned to create carbon dioxide.

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End of Chapter Problems: ***Answers***

1. a. trigonal planar, trigonal planar, sp^2. b. linear, linear, sp. c. tetrahedral, tetrahedral, sp^3. d. octahedral, square planar, sp^{3}d^2
2. BO(C_2) = 2, diamagnetic. BO(F_2) = 1, diamagnetic.
3. By MO theory, dilithium (Li_2) should exist (BO = 1, diamagnetic.)
4. MO theory would predict that B_2^{+1}(bond order = 0.5, paramagnetic) is stronger than Be_2 (bond order = 0, this should not exist at all.)
5. sp^3 to sp

Be sure to view practice problem set #2 and self quizzes for
***MO theory and Valence Bond theory** examples and practice*

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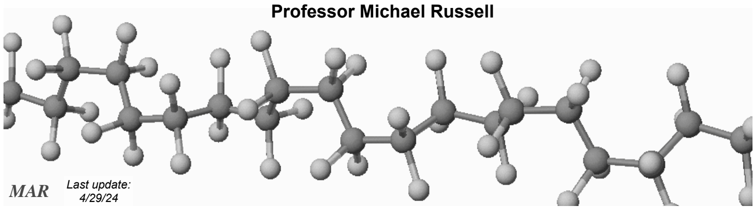
Carbon: Not Just Another Element

Chapter 20:
The Organic Chemistry chapter!

Chemistry 222
Professor Michael Russell



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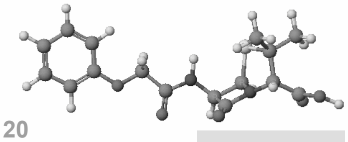
Last update: 4/29/24

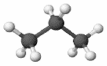
Organic Chemistry

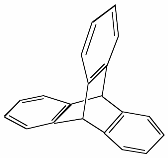


Vast majority of over 20 million known compounds are based on Carbon: organic compounds.

Generally contain C, H + other elements

Great variety of compounds





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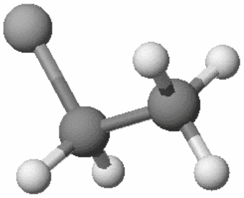
Two Major Concepts



NOMENCLATURE - Naming the molecules correctly and knowing the general classes of organic compounds

REACTIVITY - Studying patterns of reactivity within classes of compounds

We will focus primarily on nomenclature but also show examples of reactivity

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Nomenclature

Need to know Alkyl Groups -

methyl = CH_3

ethyl = CH_3CH_2

propyl = $CH_3CH_2CH_2$

butyl = $CH_3CH_2CH_2CH_2$

Also **pentyl, hexyl, heptyl, octyl,** *etc.*

R is "generic" alkyl group

Alkyl groups may be combined with other elements or alkyl groups to give classes of compounds

See the Organic Chemistry Nomenclature Guide

MAR

ALKANES

Generic Alkane Representation: R-H

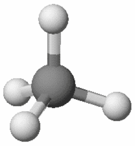
Generic Alkane Formula: C_nH_{2n+2}

-yl +ane

Ex: **methane = CH_4**

(methyl group + H)

CH_3-H



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ALKANES

Generic Alkane Representation: R-H

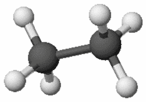
Generic Alkane Formula: C_nH_{2n+2}

-yl +ane

Ex: **ethane = CH_3CH_3**

(ethyl group + H)

CH_3CH_2-H



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ALKANES

Generic Alkane Representation: R-H

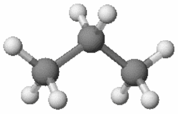
Generic Alkane Formula: C_nH_{2n+2}

-yl +ane

Ex: propane = $CH_3CH_2CH_3$

(propyl group + H)

$CH_3CH_2CH_2$-H



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ALKANES

Generic Alkane Representation: R-H

Generic Alkane Formula: C_nH_{2n+2}

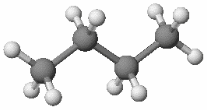
-yl +ane

Ex: butane = $CH_3CH_2CH_2CH_3$

(butyl group + H)

$CH_3CH_2CH_2CH_2$-H

Alkanes often called "saturated hydrocarbons" – all carbons "saturated" with H



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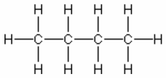
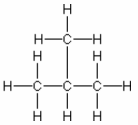
ALKANES - Isomers

But we can write butane in two ways:



n-butane



2-methylpropane

These are isomers (same formula, structurally different). Name using "longest chain" alkane preceded by numbered alkyl groups

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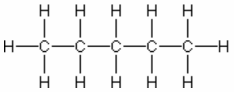
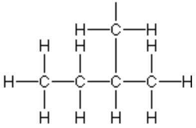
Isomers of Pentane

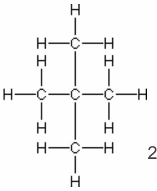


n-pentane



2-methylbutane

"3-methylbutane" would be incorrect - use smallest number possible



2,2-dimethylpropane

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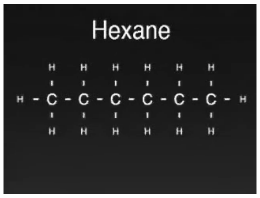
Isomers of Hexane



Number of isomers grows as number of carbons increases

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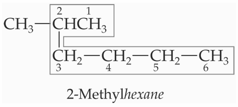
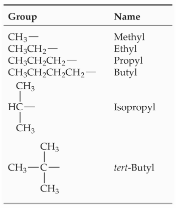
How to Name a Compound



2-Methyl*hexane*



Group	Name
CH_3—	Methyl
CH_3CH_2—	Ethyl
$CH_3CH_2CH_2$—	Propyl
$CH_3CH_2CH_2CH_2$—	Butyl
CH_3 HC— CH_3	Isopropyl
CH_3 CH_3—C— CH_3	*tert*-Butyl

1. **Find the longest chain in the molecule.**
2. **Number the chain from the end nearest the first substituent encountered.**
3. **List the substituents as a prefix along with the number(s) of the carbon(s) to which they are attached.**

"Longest chain, smallest number"

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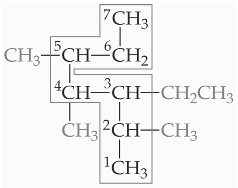
How to Name a Compound



3-Ethyl-2,4,5-trimethylheptane

If there is more than one type of substituent in the molecule, list them alphabetically.

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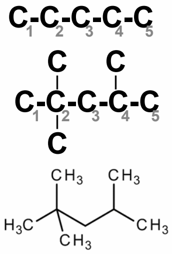
"Longest chain, smallest number"

How to Create a Structure from a Name

Example: provide the structure for the following name: **2,2,4-trimethylpentane**



1. Start at the end of the name to find the chain of carbons; write them "in a row" and number them
2. Groups not in the chain will be listed at the beginning of the name (methyl = CH_3, etc.)
3. Fill in hydrogen atoms at the end if necessary

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"Longest chain, smallest number"

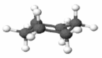
CYCLOALKANES (C_nH_{2n})

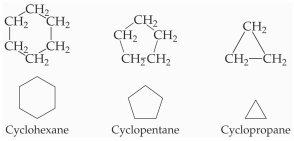
+cyclo -yl +ane

Ex: **cyclohexane = C_6H_{12}**

Generic Cycloalkane Formula: C_nH_{2n}

(cyclohexyl group + H)





Also cyclobutane, etc.

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Alkyl Halides

Generic Alkyl Halide Representation: R-X

X = halogen (F, Cl, Br or I)

Ex: **methyl iodide = CH_3-I**

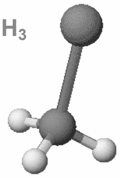
(methyl + iodide)

also known as iodomethane

Ex: **2-iodopropane = CH_3-CHI-CH_3**

(2-propyl + iodide)

also known as 2-propyl iodide



Many other possibilities

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ALCOHOLS

Generic Alcohol Representation: R-OH

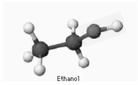
-yl +anol

Generic Alcohol Formula: $C_nH_{2n+2}O$

Ex: **ethanol = CH_3CH_2OH**

(ethyl group + OH)

CH_3CH_2-OH



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ALCOHOLS

Generic Alcohol Representation: R-OH

-yl +anol

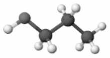
Generic Alcohol Formula: $C_nH_{2n+2}O$

Ex: **1-propanol = $CH_3CH_2CH_2OH$**

(propyl group + OH)

$CH_3CH_2CH_2$-OH

new: **propan-1-ol**



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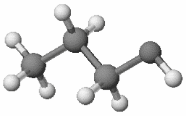
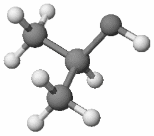
ALCOHOLS

Generic Alcohol Representation: R-OH

Note that both 1-propanol and 2-propanol exist



1-propanol
CH_3-CH_2-CH_2OH
propan-1-ol



2-propanol
CH_3-CH(OH)-CH_3
propan-2-ol

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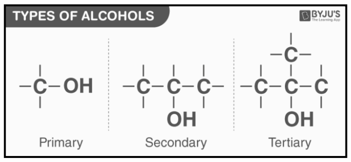
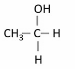
TYPES OF ALCOHOLS



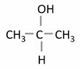
Primary alcohols



$CH_3-\underset{H}{\overset{OH}{C}}-H$

ethanol

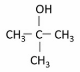
Secondary alcohols



$CH_3-\underset{H}{\overset{OH}{C}}-CH_3$

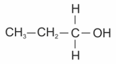
propan-2-ol

Tertiary alcohols



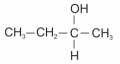
$CH_3-\underset{CH_3}{\overset{OH}{C}}-CH_3$

2-methyl-propan-2-ol



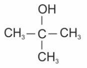
$CH_3-CH_2-\underset{H}{\overset{H}{C}}-OH$

Primary Alcohol
propan-1-ol



$CH_3-CH_2-\underset{H}{\overset{OH}{C}}-CH_3$

Secondary Alcohol
butan-2-ol



$CH_3-\underset{CH_3}{\overset{OH}{C}}-CH_3$

Tertiary Alcohol
2-methyl-propan-2-ol

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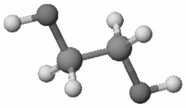
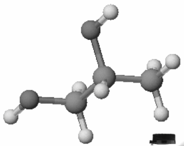
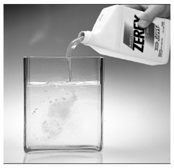
GLYCOLS
Alcohols (diols) with Two OH Groups







Ethylene glycol
(ethane-1,2-diol)

Propylene glycol
(propane-1,2-diol)

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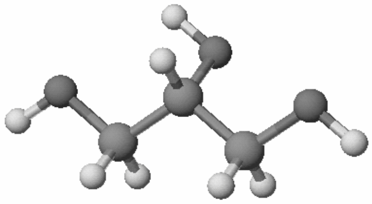
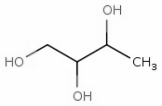
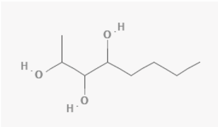
Glycerol (propane-1,2,3-triol)
Alcohol with 3 OH Groups





butane-1,2-3-triol



octane-2,3-4-triol

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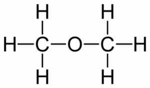
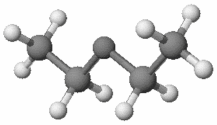
ETHERS - "old school"

Generic Ether Representation: R-O-R
***Generic Ether Formula:* $C_nH_{2n+2}O$**
***Ex:* diethyl ether = CH_3CH_2-O-CH_2CH_3**
(ethyl + O + ethyl)



$H-\underset{H}{\overset{H}{C}}-O-\underset{H}{\overset{H}{C}}-H$

dimethyl ether



CH_3—O—CH_2CH_3

ethyl methyl ether

Many other possibilities

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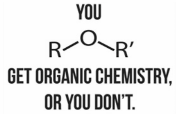
ETHERS - "IUPAC"

Two nomenclatures for ethers!

Ethers also use "IUPAC" Nomenclature
(shorter alkyl group -yl +oxy) (longer group -yl +ane)

***Ex:* 1-ethoxypropane = CH_3CH_2-O-$CH_2CH_2CH_3$**
(ethyl -yl + oxy)(propyl -yl +ane)
***old school:* ethyl 1-propyl ether**

***Ex:* 2-methoxypropane = $CH_3CH(OCH_3)$-CH_3**
2-(methyl -yl + oxy)(propyl -yl +ane)
***old school:* methyl 2-propyl ether**



CH_3—O—CH_2CH_3

methoxy ethane

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KETONES

Generic Ketone Representation: R-CO-R

R = alkyl group, CO = carbonyl (C=O)

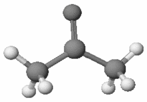
Generic Ketone Formula: $C_nH_{2n}O$

Two carbons connected to carbonyl

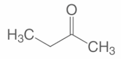
Ex: **propanone = CH_3-C(=O)-CH_3**

(methyl + C=O + methyl)

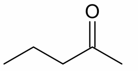
3 carbon atoms like *propane*



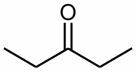
Propanone is also known as Acetone



butanone



2-pentanone



3-pentanone

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ALDEHYDES

Generic Aldehyde Representation: R-CO-H

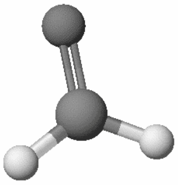
R = alkyl group or H, CO = carbonyl (C=O)

Generic Aldehyde Formula: $C_nH_{2n}O$

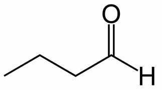
At least one H connected to carbonyl

Ex: **methanal = H_2-C=O**

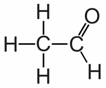
1 carbon atom like *methane*



Methanal is also known as Formaldehyde



butanal



ethanal (acetaldehyde)

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ALKYNES

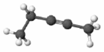
Generic Alkyne Representation: R-C≡C-R

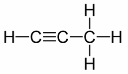
R = alkyl group or H

Generic Alkyne Formula: C_nH_{2n-2}

Ex: **2-pentyne = H_3C-C≡C-CH_2CH_3**

5 carbon atoms like *pentane*

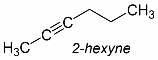




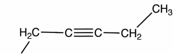
$H-C\equiv C-CH_3$

propyne

Ethyne is also known as Acetylene



2-hexyne



3-hexyne

ALKENES

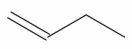
Generic Alkene Representation: R-HC=CH-R

R = alkyl group or H

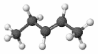
Generic Alkene Formula: C_nH_{2n}

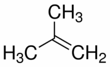
Ex: ***trans*-pent-2-ene = H_3C-HC=CH-CH_2CH_3**

5 carbon atoms like *pentane*

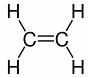


but-1-ene





2-methyl propene



ethene or ethylene

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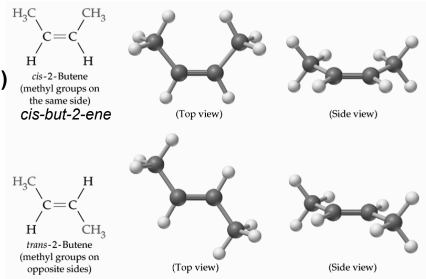
ALKENES - *cis* and *trans*

cis
(same side)

trans
(opposite)



cis-but-2-ene

trans-but-2-ene

Important!

Differing reactivity in *cis* and *trans* isomers

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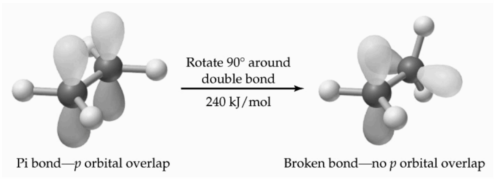
ALKENES - *cis* and *trans*

***Cis - trans* isomerism** occurs because the electronic structure of the carbon–carbon double bond makes rotation energetically unfavorable.



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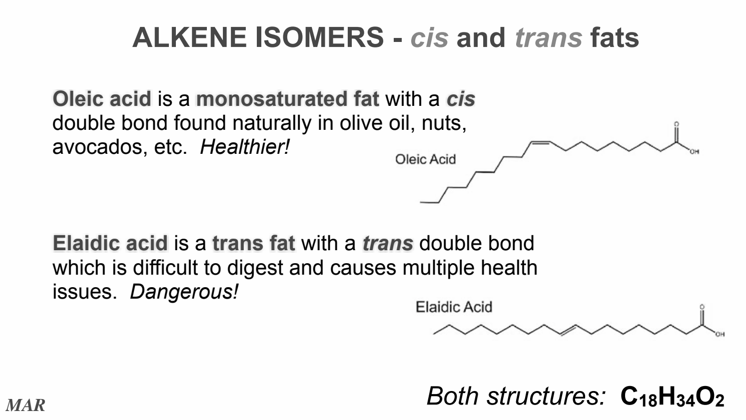

ALKENE ISOMERS - *cis* and *trans* fats

Oleic acid is a **monosaturated fat** with a ***cis*** double bond found naturally in olive oil, nuts, avocados, etc. *Healthier!*

Oleic Acid

Elaidic acid is a **trans fat** with a ***trans*** double bond which is difficult to digest and causes multiple health issues. *Dangerous!*

Elaidic Acid

Both structures: **$C_{18}H_{34}O_2$**

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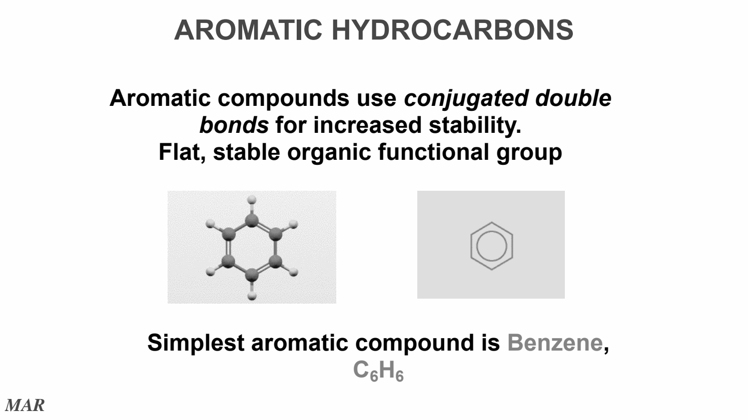

AROMATIC HYDROCARBONS

Aromatic compounds use *conjugated double bonds* for increased stability.
Flat, stable organic functional group

Simplest aromatic compound is Benzene, C_6H_6

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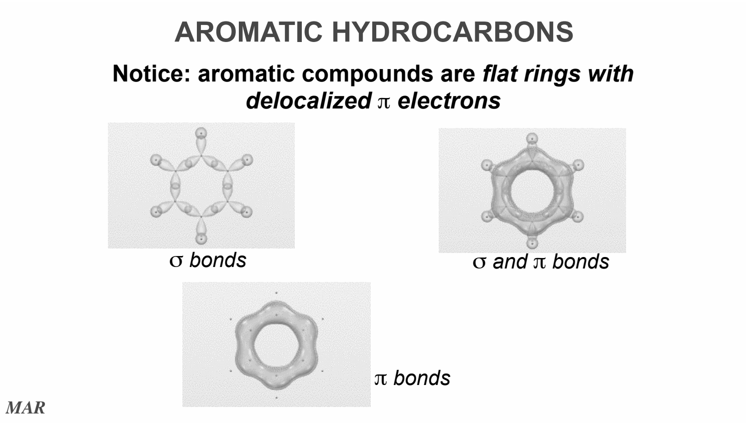

AROMATIC HYDROCARBONS

Notice: aromatic compounds are *flat rings with delocalized* π *electrons*

σ *bonds*

σ *and* π *bonds*

π *bonds*

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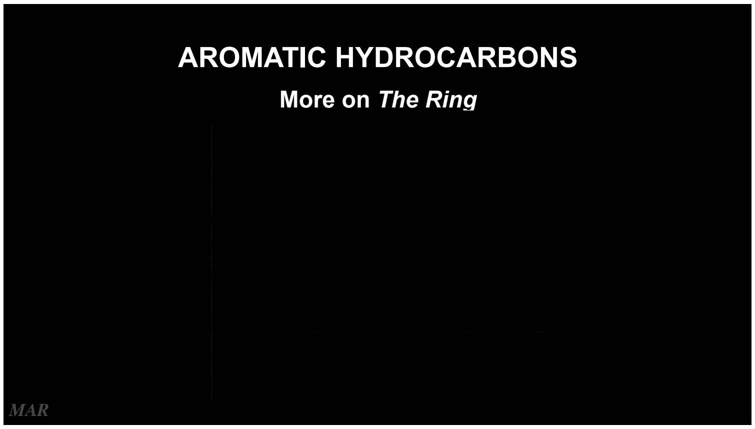

AROMATIC HYDROCARBONS

More on *The Ring*

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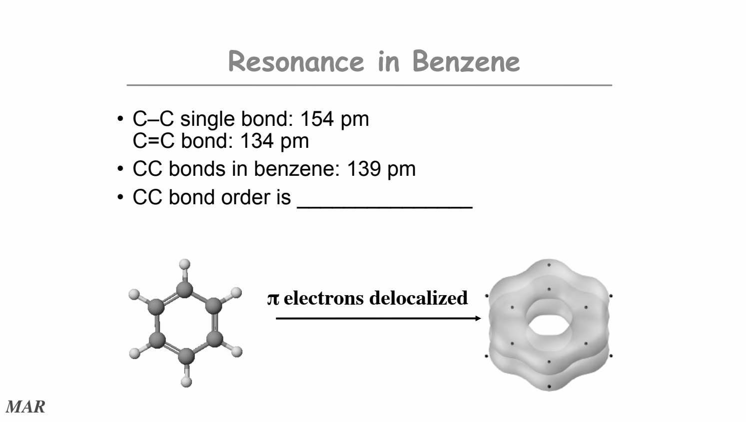

Resonance in Benzene

- C–C single bond: 154 pm
  C=C bond: 134 pm
- CC bonds in benzene: 139 pm
- CC bond order is _______________

π electrons delocalized

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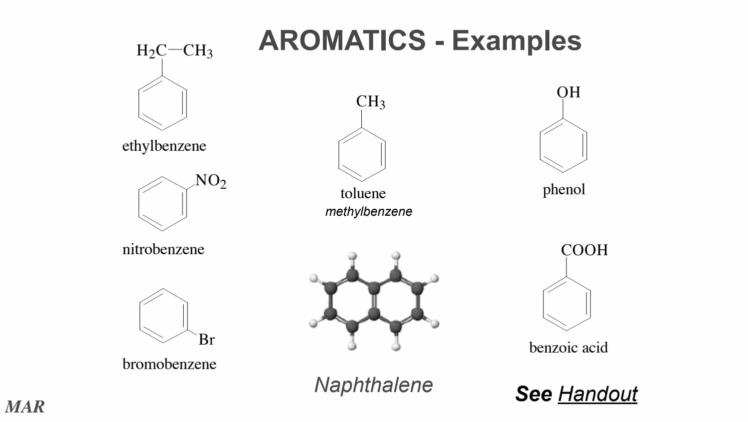

AROMATICS - Examples

ethylbenzene

toluene
methylbenzene

phenol

nitrobenzene

bromobenzene

Naphthalene

benzoic acid

***See** Handout*

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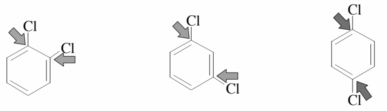
AROMATICS - Examples

ortho position — *meta* position — *para* position



1,2-dichlorobenzene or o-dichlorobenzene

1,3-dichlorobenzene or m-dichlorobenzene

1,4-dichlorobenzene or p-dichlorobenzene

See Handout

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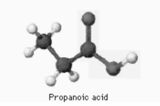
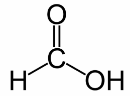
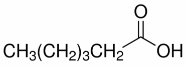
CARBOXYLIC ACIDS

Generic Carboxylic Acid Representation:
R-COOH
Generic Carboxylic Acid Formula: $C_nH_{2n}O_2$
Ex: **propanoic acid = CH_3CH_2-CO-OH**
(propyl -2H + =O + OH)

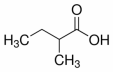




methanoic acid or formic acid



hexanoic acid



2-methyl-butanoic acid

Many other possibilities!

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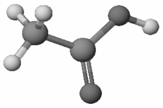
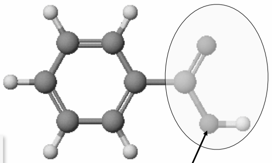
Carboxylic Acids

Acetic acid



Benzoic acid



Acids are found in many natural substances: bread, fruits, milk, wine

Carboxylic acid group with acidic H⁺
All are WEAK acids

Vinegar *is essentially acetic acid!*

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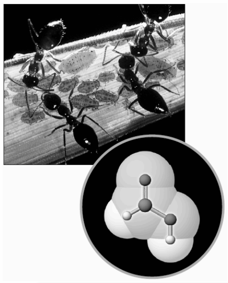
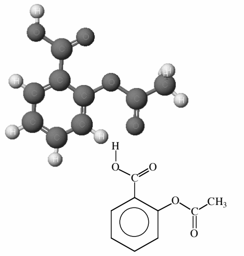
Carboxylic Acids



Formic acid, HCO_2H, gives the sting to ants.



Aspirin, acetylsalicylic acid

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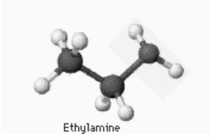
AMINES

Generic Amine Representation: **$R_{(3-x)}$-NH_x**

Ex: **ethylamine = CH_3CH_2-NH_2**
(ethyl + NH_2)



Also diethylamine and triethylamine

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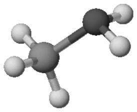
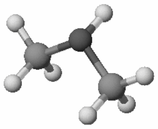
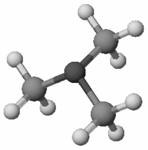


Methylamine

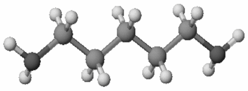


Dimethylamine

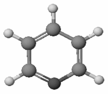


Trimethylamine

Amines generally have terrible odors!



Cadaverine



Pyridine

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ESTERS

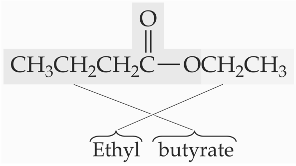
Generic Ester Representation: **R_1-COO-R_2**

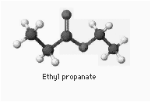
Generic Ester Formula: $C_nH_{2n}O_2$

Ex: **ethyl propanate = CH_3CH_2-COO-CH_2CH_3**

(propyl -2H + =O + O + ethyl)

Similar to carboxylic acids




Ethyl propanate

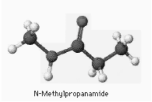
Name R_2 first, then R_1

MAR

AMIDES

Generic Amide Representation:
R_x-CO-$NH_yR_{(2-y)}$

Ex: ***N*-methylpropanamide**
CH_3NH-CO-CH_2CH_3

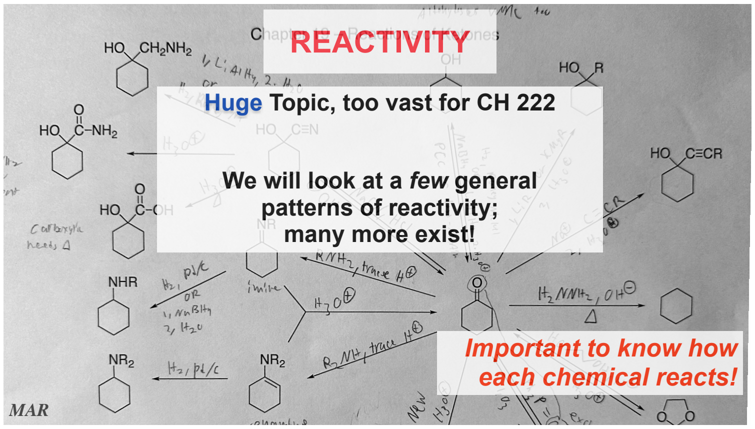

N-Methylpropanamide

N-groups off of nitrogen atom

MAR

Nomenclature and Models

MAR



REACTIVITY

Huge Topic, too vast for CH 222

We will look at a *few* general patterns of reactivity; many more exist!

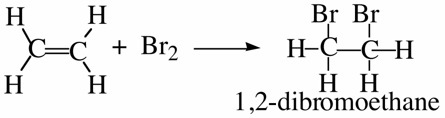
Important to know how each chemical reacts!

MAR

Reactions of Alkenes: ADDITION REACTIONS

Alkenes are unsaturated — more bonds can form to the C atoms

Molecules such as Br_2, H_2, HCl, HBr, and H_2O add to the double bond



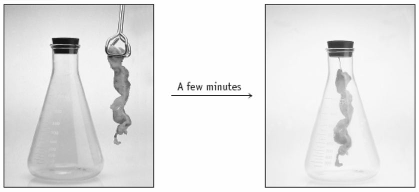
$$H_2C{=}CH_2 + Br_2 \longrightarrow H\text{-}CHBr\text{-}CHBr\text{-}H$$

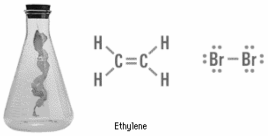
1,2-dibromoethane

MAR

C_2H_4 + Br_2 →
CH_2Br-CH_2Br

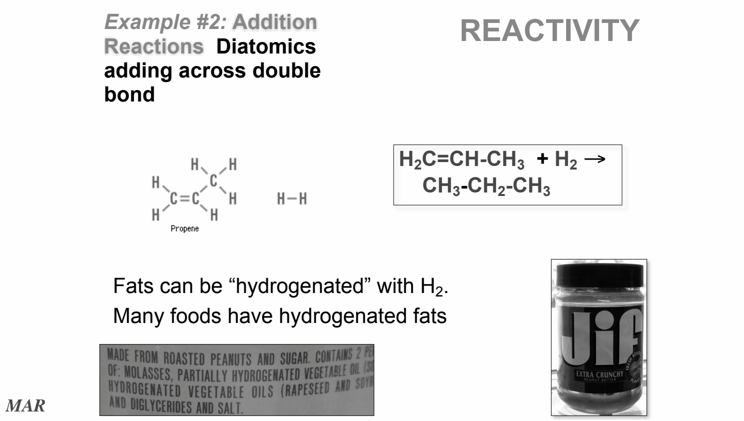
REACTIVITY





Example #1:
Addition Reactions
Diatomics adding across double bond

MAR



REACTIVITY

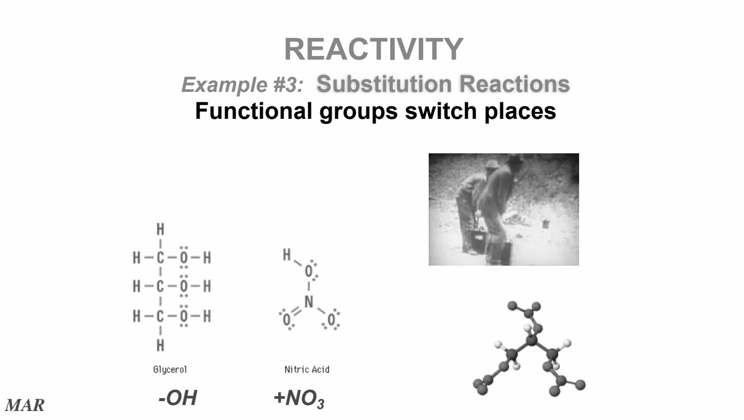
Example #2: **Addition Reactions** **Diatomics adding across double bond**

$H_2C{=}CH\text{-}CH_3 + H_2 \rightarrow CH_3\text{-}CH_2\text{-}CH_3$

Fats can be “hydrogenated” with H_2.
Many foods have hydrogenated fats

MADE FROM ROASTED PEANUTS AND SUGAR. CONTAINS 2 P
OF: MOLASSES, PARTIALLY HYDROGENATED VEGETABLE OIL (R
HYDROGENATED VEGETABLE OILS (RAPESEED AND SOY
AND DIGLYCERIDES AND SALT.

MAR



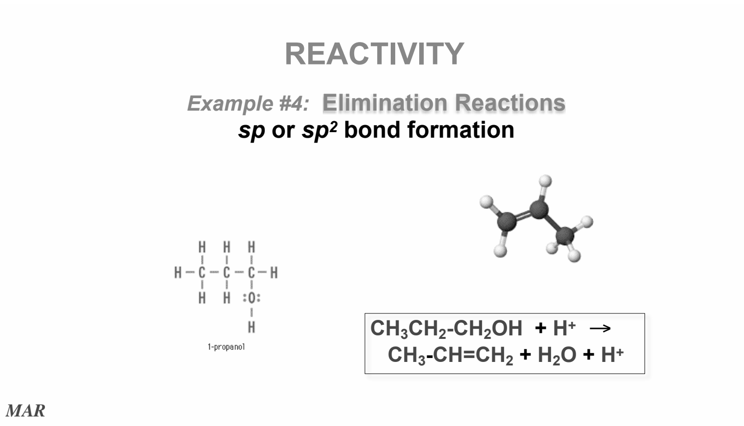
REACTIVITY

Example #3: **Substitution Reactions**

Functional groups switch places

-OH ***+NO₃***

MAR



REACTIVITY

Example #4: **Elimination Reactions**

***sp* or *sp²* bond formation**

$CH_3CH_2\text{-}CH_2OH + H^+ \rightarrow CH_3\text{-}CH{=}CH_2 + H_2O + H^+$

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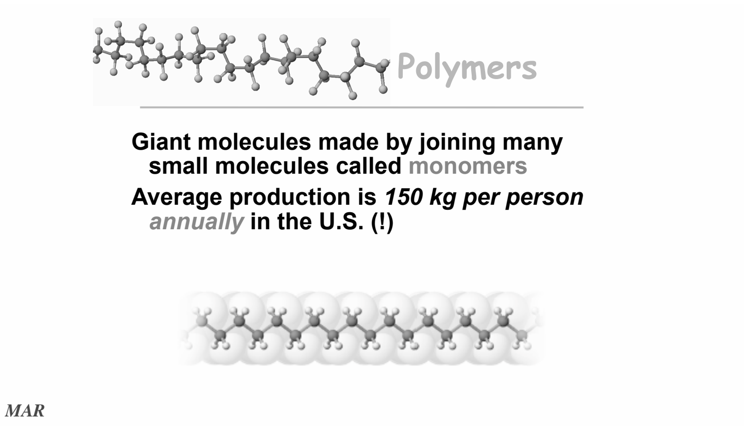
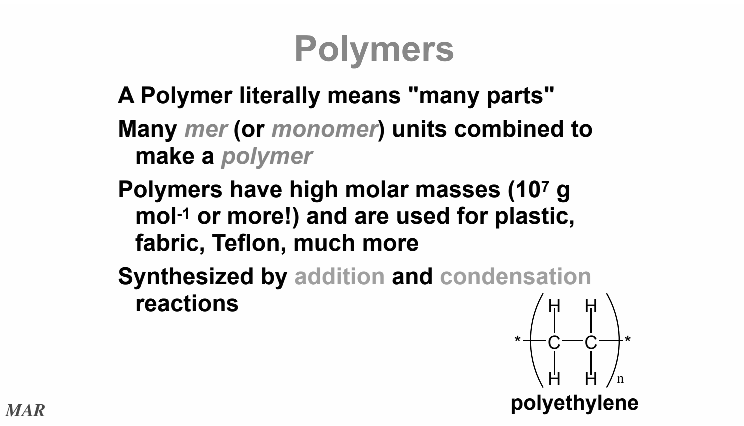


Polymers

- **Giant molecules made by joining many small molecules called monomers**
- **Average production is *150 kg per person annually* in the U.S. (!)**

MAR

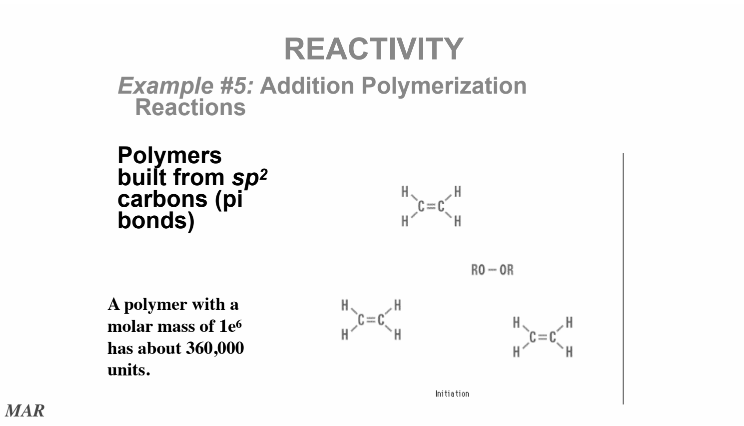


Polymers

- **A Polymer literally means "many parts"**
- **Many *mer* (or *monomer*) units combined to make a *polymer***
- **Polymers have high molar masses (10^7 g mol^{-1} or more!) and are used for plastic, fabric, Teflon, much more**
- **Synthesized by addition and condensation reactions**

polyethylene

MAR



REACTIVITY

Example #5: **Addition Polymerization Reactions**

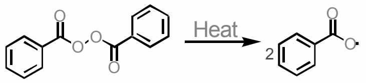
Polymers built from *sp²* carbons (pi bonds)

A polymer with a molar mass of 1e⁶ has about 360,000 units.

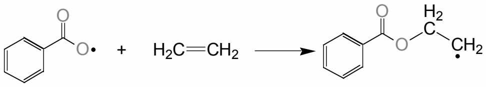
MAR

Polyethylene Synthesis

Chain initiator: **benzoyl peroxide**



Initiation Step: **Reaction of benzoyl radical**



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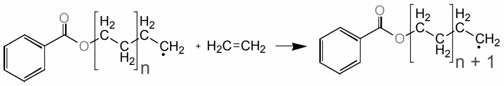
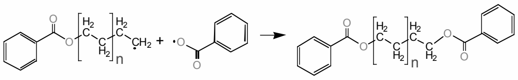
Polyethylene Synthesis

Chain Propagation: Addition of further ethylene



Chain Termination: Reaction of two radicals



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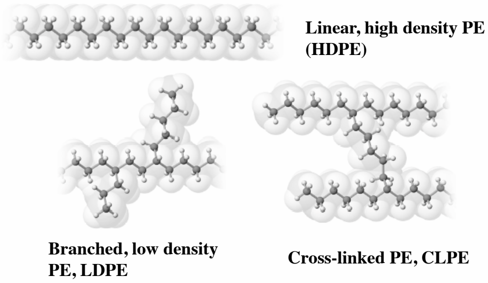
Types of Polyethylene



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Polystyrene

Polystyrene (PS) is a nonpolar material and dissolves in organic solvents.

PS foam is mostly air, and when it dissolves it collapses to a *much* smaller volume.

MAR

REACTIVITY

Example #6: Condensation Polymerization Reactions

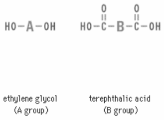
Condensation reactions combine different functional groups to make polymers with different properties

Very powerful reaction mechanism; used in contact lenses, nylon, much more

MAR

REACTIVITY

Example #6: Condensation Polymerization Reactions



$\text{HO}-\text{A}-\text{OH} \qquad \text{HO}-\overset{\text{O}}{\overset{\|}{\text{C}}}-\text{B}-\overset{\text{O}}{\overset{\|}{\text{C}}}-\text{OH}$

Polyester

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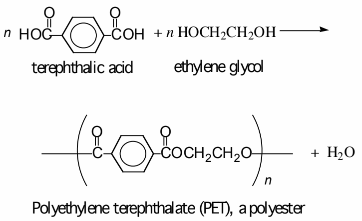
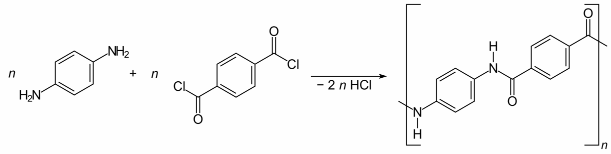
Condensation Polymerization Reactions



Formation of polyester

Formation of Kevlar



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Polyesters (PET)

Jackets made from recycled PET soda bottles

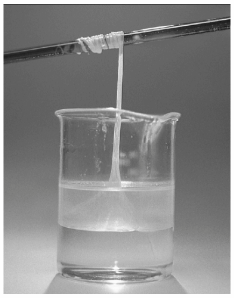
Soda bottles, mylar film.

MAR

REACTIVITY

Example #6: Condensation Polymerization Reactions

Nylon-6,6



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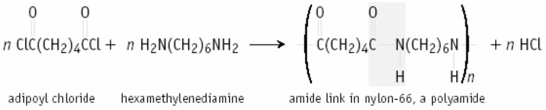
Polyamides: Nylon



$$n\,\mathrm{Cl\overset{O}{\|}{C}(CH_2)_4\overset{O}{\|}{C}Cl} + n\,\mathrm{H_2N(CH_2)_6NH_2} \longrightarrow \left(\mathrm{\overset{O}{\|}{C}(CH_2)_4\overset{O}{\|}{C}\;\underset{H}{|}{N}(CH_2)_6\underset{H}{|}{N}}\right)_n + n\,\mathrm{HCl}$$

Each monomer has 6 C atoms in its chain.

A polyamide link forms on elimination of HCl

Result = nylon-6,6

MAR

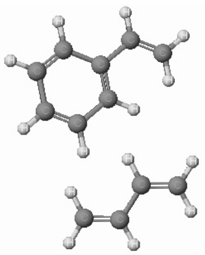
USES FOR POLYMERS

Examples of Polymers:

- Teflon - *polytetrafluoroethylene*
- Fabric - *polyester, polyacrylonitrile*
- Milk & soda bottles - *(High Density) Polyethylene*
- Styrofoam - *polystyrene*
- plastic wrap (Saran) - *poly(vinylidene chloride)*
- contact lenses - *poly(methyl methacrylate)*
- *Other uses:*

MAR

Bubble Gum! A copolymer



Styrene + butadiene

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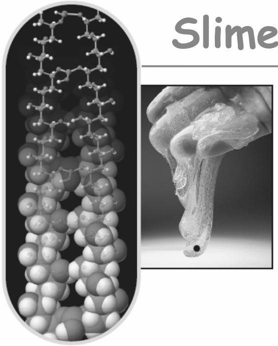
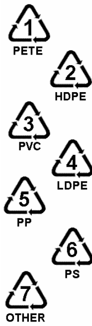
Slime!



Slime is polyvinylalcohol cross-linked with boric acid

MAR

Polymer Recyc



PETE = Polyethylene terephtha

HDPE = High density PE, 0.941

V = PVC (polyvinyl chloric

LDPE = Low density PE, 0.910-

PP = Polypropylene, 0.90-0.9

PS = Polystyrene, 1.03-1.06

Other = Other plastics (varies)

Water and soft drink bottles, salad domes, biscuit trays, salad dressing and peanut butter containers

Milk bottles, freezer bags, dip tubs, crinkly shopping bags, ice cream containers, juice bottles, shampoo, chemical and detergent bottles

Cosmetic containers, commercial cling wrap

Squeeze bottles, cling wrap, shrink wrap, rubbish bags

Microwave dishes, ice cream tubs, potato chip bags, and dip tubs

CD cases, water station cups, plastic cutlery, imitation "crystal glassware", video cases

Foamed polystyrene hot drink cups, hamburger take-away clamshells, foamed meat trays, protective packaging for fragile items

Water cooler bottles, flexible films, multi-material packaging

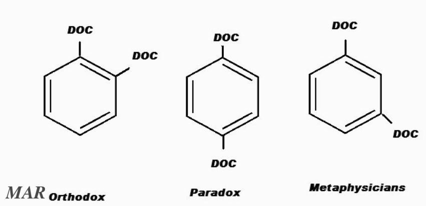
WORLD ECONOMIC FORUM, ELLEN MACARTHUR FOUNDATION, MCKINSEY & COMPANY, A NEW PLASTICS ECONOMY: RETHINKING THE FUTURE OF PLASTICS (2016) ELLENMACARTHURFOUNDATION.ORG/PUBLICATIONS

Source: Project MainStream analysis

End of Chapter 20

See:

- *Chapter Twenty Study Guide*
- *Chapter Twenty 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* **name organic compounds** *using the* **functional group** *along with the* **"longest chain, shortest number"** *concept*
- **recognize some common organic chemistry reactions**
- *see the* **Organic Chemistry Nomenclature Guide** *(handout)*

Organic Chemistry: alkyl group, alkane, cycloalkane, alkyl halide, alcohol, ether, ketones, aldehydes, alkynes, alkenes, aromatic compounds, carboxylic acids, amines, isomers

MAR

End of Chapter Problems: *Test Yourself*

Be sure to view practice problem set #3 and self quizzes for ***organic chemistry nomenclature*** *examples and practice*

1. Name a straight chain alkane with six carbon atoms.
2. Name this compound: $\mathbf{CH_3CH_2CH_2OH}$ Provide two isomers of this compound and name them.
3. Draw structural formulas for all four compounds with the formula $C_4H_{10}Br$. Give the systematic name of each.
4. Provide IUPAC numbered names for the following three compounds: *m*-dibromobenzene, *o*-dibromobenzene, *p*-dibromobenzene
5. Which of the following would exhibit *cis, trans* isomerization? 1-pentene, propene, 2-butene

MAR

End of Chapter Problems: *Answers*

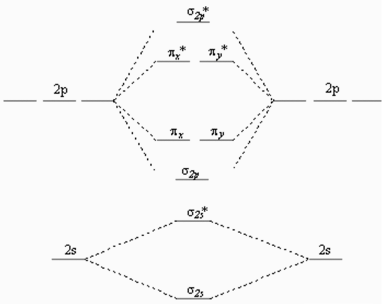
1. n-hexane
2. 1-propanol. 2-propanol and 1-methoxy propane would be isomers.
3. 1-bromobutane, 2-bromobutane, 2-bromo-2-methylpropane, 1-bromo-2-methylpropane
4. *m*-dibromobenzene = 1,3-dibromobenzene, *o*-dibromobenzene = 1,2-dibromobenzene, *p*-dibromobenzene = 1,4-dibromobenzene
5. only 2-butene would exhibit *cis, trans* isomerization.

Be sure to view practice problem set #3 and self quizzes for ***organic chemistry nomenclature*** *examples and practice*

MAR

Chemistry 222 Exam I Review
Chapters 7, 8 and 20

Chemistry 222
Professor Michael Russell



[Core electrons]

Last update: 4/29/24

Which of the following is NOT a correct Lewis dot structure?

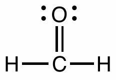
A. :N≡N:

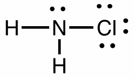
B. [:N≡O:]$^{-}$

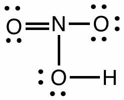
C. H—C≡N:

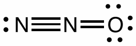
D. :C≡O:

Which of the following is NOT a correct Lewis dot structure?

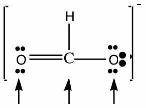
A.


C.


B.


D.


Determine the formal charges for the formate ion:



Formal charge = ___ ___ ___

A. 0 0 0
B. +1 –1 –1
C. +1 0 –1
D. 0 0 –1

Using VSEPR, which of the following corresponds to the molecular shape of SCl_2?

A. linear
B. bent
C. "T-shaped"
D. trigonal pyramid
E. orthorhombic

Based on VSEPR, which of the following corresponds to the molecular shape of the IF_2^- ion?

A. linear
B. bent (bond angle 120˚)
C. "T-shaped"
D. bent (bond angle 109.5˚)
E. trigonal pyramid

What is the approximate C-C-H angle in the acetylene, C_2H_2, molecule?

A. 90˚
B. 120˚
C. 109.5˚
D. 180˚
E. -30˚

MAR

Which of the following groups of elements is arranged correctly in order of increasing electronegativity?

A. Mg < P < N < F
B. Mg < N < P < F
C. N < Mg < P < F
D. F < P < Mg < N
E. S < U < Pr < I < Se

MAR

What is the carbon-oxygen bond order in formaldehyde, CH_2O?

A. 1
B. 2
C. 1 ½
D. 2 ½
E. 0

MAR

What is the average carbon-oxygen bond order in the formate ion, HCO_2^-?

A. 1
B. 2
C. 1 ½
D. 2 ½
E. -1

MAR

Given the bond dissociation enthalpies below, calculate the standard molar enthalpy of formation of NF_3:

$$\tfrac{1}{2} N_2(g) + \tfrac{3}{2} F_2(g) \rightarrow NF_3(g)$$

Bond	Dissociation Enthalpy (kJ/mol)
N≡N	946
F—F	159
N—F	272

A. 833 kJ/mol
B. -105 kJ/mol
C. 440. kJ/mol
D. -578 kJ/mol
E. -220. kJ/mol

MAR

Which of the bonds below is least polar?

A. B—O
B. B—N
C. B—F
D. B—C
E. AC-DC

MAR

Which of the following molecules is polar?

A. BCl_3
B. CO_2
C. N_2
D. ClF
E. Ne

MAR

Which of the following molecules is most likely to have a dipole moment?

A. CH_4
B. SF_6
C. BeF_2
D. NF_3
E. Rn

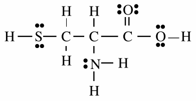
MAR

What hybrid orbital set is used by the boron atom in the BCl_4^- ion?

A. sp
B. sp^3
C. sp^2
D. sp^3d^2
E. sp^3d

MAR

Cysteine is one of the natural amino acids.

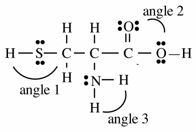


The molecule has

A. 13 sigma (σ) bonds and 1 pi (π) bond
B. 14 sigma (σ) bonds and 1 pi (π) bond
C. 20 sigma (σ) bonds and 1 pi (π) bond
D. 12 sigma (σ) bonds and 1 pi (π) bond

MAR

Cysteine is one of the natural amino acids.



Estimate the values of the indicated angles:

A. Angle 1 = 180° Angle 2 = 120° Angle 3 = 109°
B. Angle 1 = 109° Angle 2 = 120° Angle 3 = 109°
C. Angle 1 = 109° Angle 2 = 109° Angle 3 = 109°
D. Angle 1 = 180° Angle 2 = 90° Angle 3 = 90°

MAR

Nitrogen can lose an electron to form N_2^+. What is the bond order of N_2^+?

A. 1
B. 1.5
C. 2
D. 2.5
E. 3

MAR

Is N_2^+ diamagnetic or paramagnetic?

A. diamagnetic
B. paramagnetic
C. 42

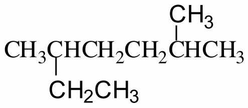
MAR

What is the bond order of the superoxide ion, O_2^-?

A. 1
B. 1.5
C. 2
D. 2.5
E. 3

MAR

What is the systematic name for this alkane?

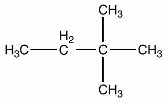


$$\underset{\displaystyle CH_2CH_3}{CH_3CHCH_2CH_2\overset{\displaystyle CH_3}{CH}CH_3}$$

A. nonane
B. 2-ethyl-5-methylhexane
C. 2,5-dimethylheptane
D. 2,3-dimethyloctane
E. 3,6-dimethylheptane

MAR

This compound:



A. The compound is an isomer of pentane and is named 2,3-dimethylbutane.
B. The compound is an isomer of octane and is named 2,2-dimethylbutane.
C. The compound is an isomer of hexane and is named 2,2-dimethylbutane.
D. The compound is an isomer of hexane and is named 3,3-dimethylbutane.

MAR

Given the following compounds:

1. C_3H_6 2. C_7H_{14} 3. $C_{11}H_{24}$ 4. C_7H_8

Which one(s) could be an alkane?

A. only 1
B. only 2
C. only 3
D. 3 and 4
E. none of the above

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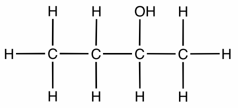
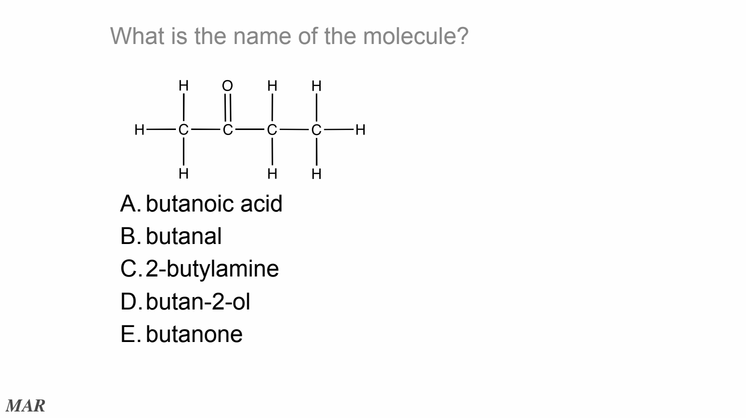
What is the name of the molecule?



A. butanoic acid
B. butanal
C. butanol
D. butan-2-ol
E. butane

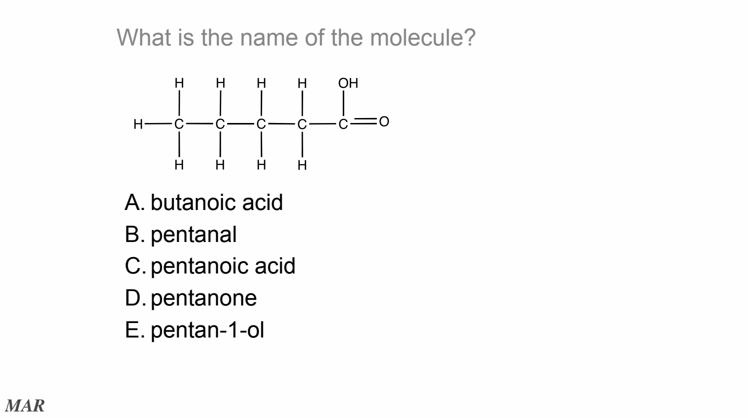
MAR



What is the name of the molecule?

A. butanoic acid
B. butanal
C. 2-butylamine
D. butan-2-ol
E. butanone

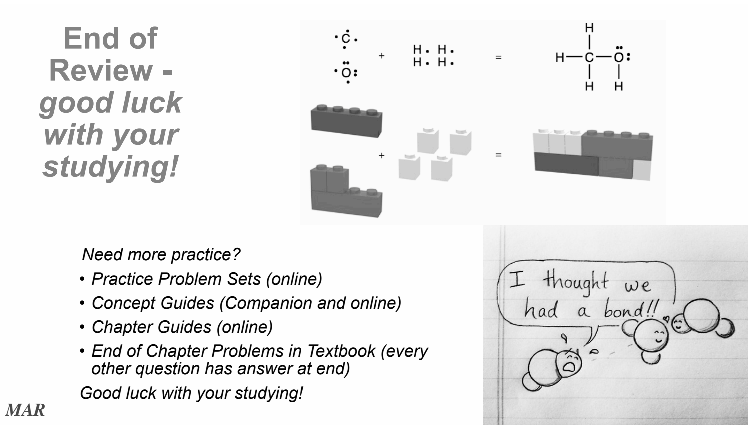
MAR



What is the name of the molecule?

A. butanoic acid
B. pentanal
C. pentanoic acid
D. pentanone
E. pentan-1-ol

MAR



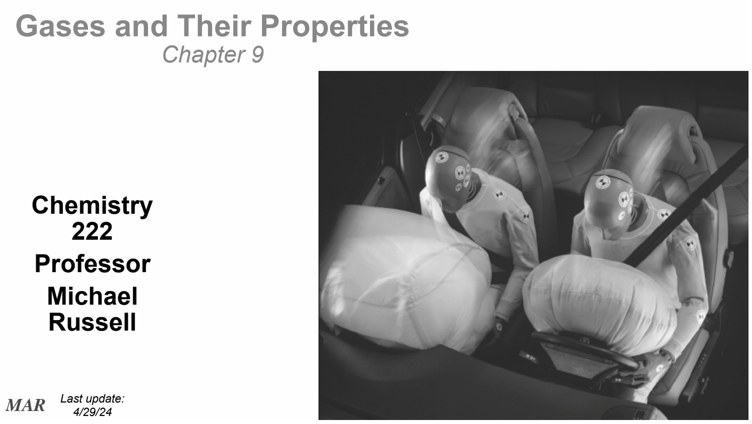
End of Review - *good luck with your studying!*

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



Gases and Their Properties

Chapter 9

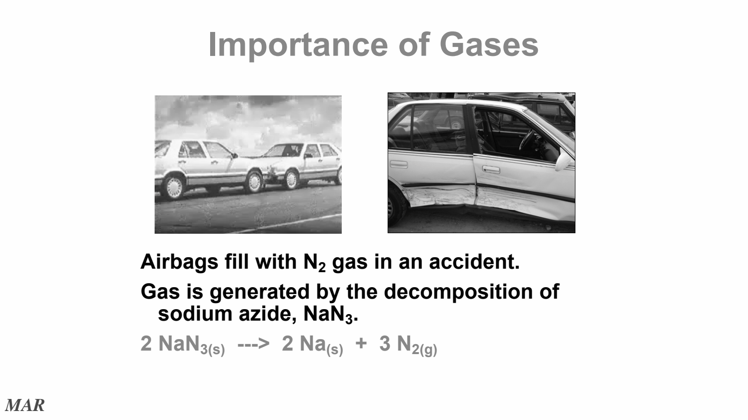
Chemistry 222

Professor

Michael Russell

MAR

Last update: 4/29/24



Importance of Gases

Airbags fill with N_2 gas in an accident.

Gas is generated by the decomposition of sodium azide, NaN_3.

$2\ NaN_{3(s)}$ ---> $2\ Na_{(s)}$ + $3\ N_{2(g)}$

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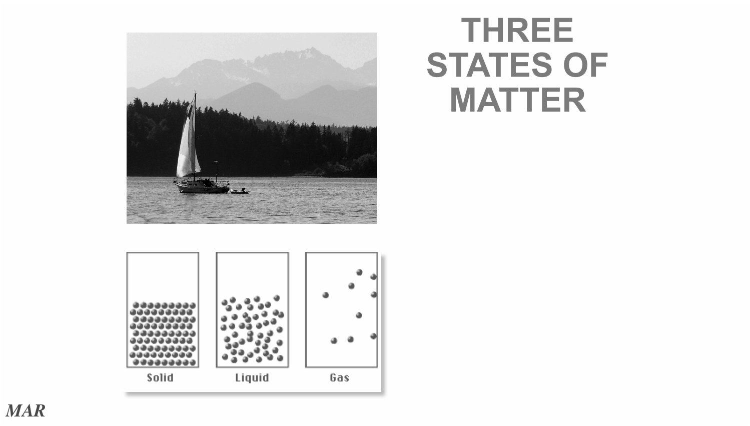

THREE STATES OF MATTER

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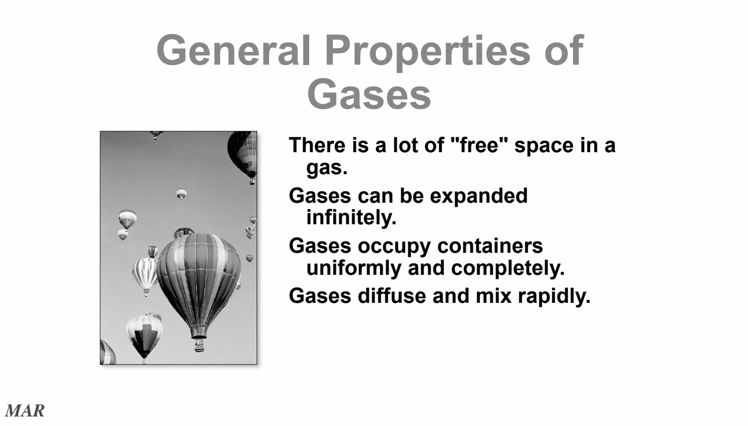


General Properties of Gases

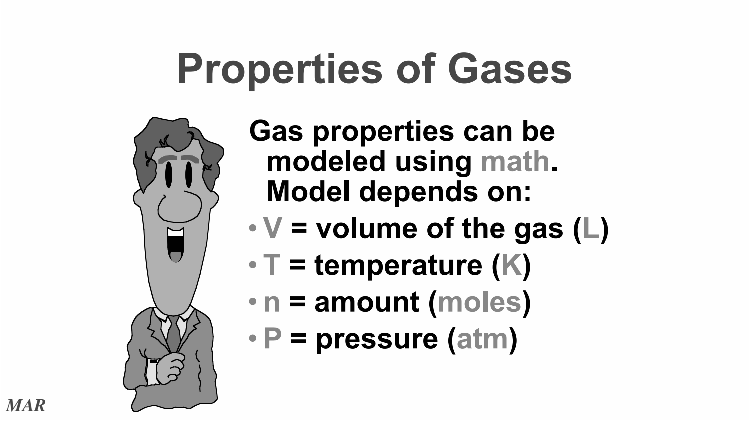
There is a lot of "free" space in a gas.

Gases can be expanded infinitely.

Gases occupy containers uniformly and completely.

Gases diffuse and mix rapidly.

MAR



Properties of Gases

Gas properties can be modeled using math. Model depends on:

- **V = volume of the gas (L)**
- **T = temperature (K)**
- **n = amount (moles)**
- **P = pressure (atm)**

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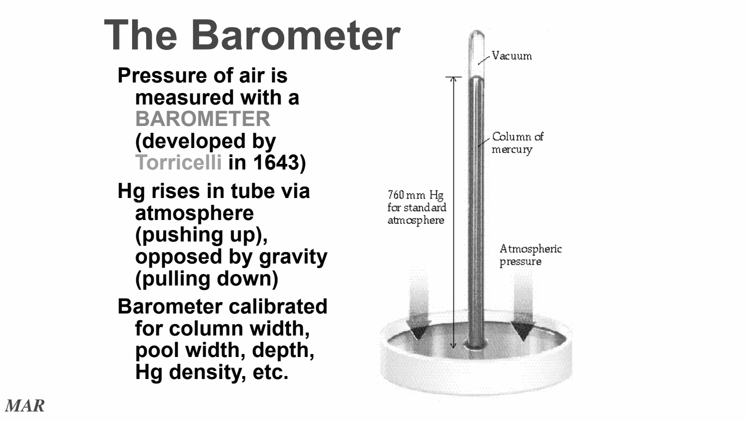


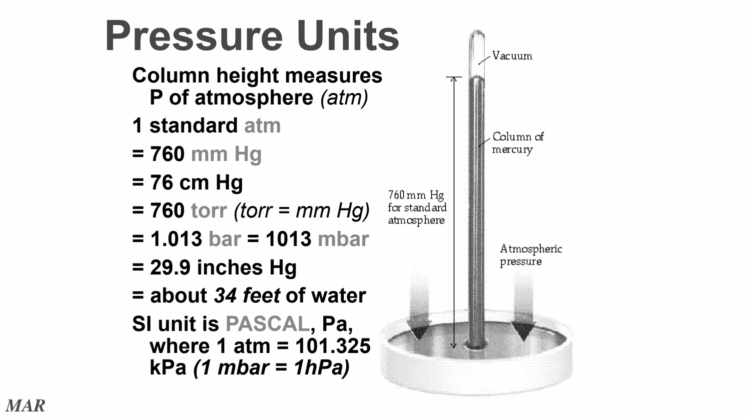
The Barometer

Pressure of air is measured with a BAROMETER (developed by Torricelli in 1643)

Hg rises in tube via atmosphere (pushing up), opposed by gravity (pulling down)

Barometer calibrated for column width, pool width, depth, Hg density, etc.

MAR



Pressure Units

Column height measures P of atmosphere *(atm)*

1 standard atm

= 760 mm Hg

= 76 cm Hg

= 760 torr *(torr = mm Hg)*

= 1.013 bar = 1013 mbar

= 29.9 inches Hg

= about *34 feet* of water

SI unit is PASCAL, Pa, where 1 atm = 101.325 kPa *(1 mbar = 1hPa)*

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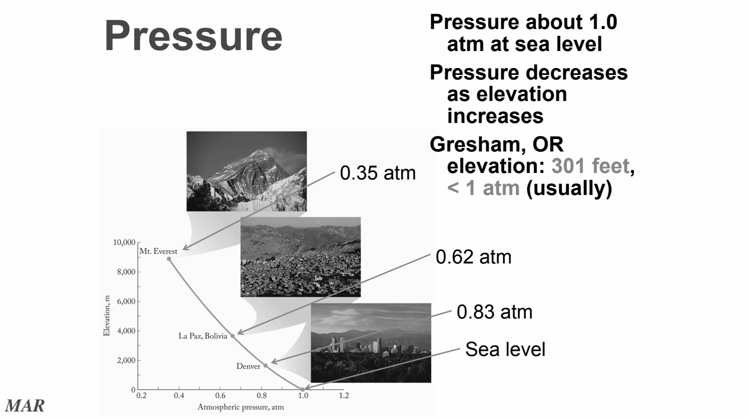

Pressure

Pressure about 1.0 atm at sea level

Pressure decreases as elevation increases

Gresham, OR elevation: 301 feet, < 1 atm (usually)

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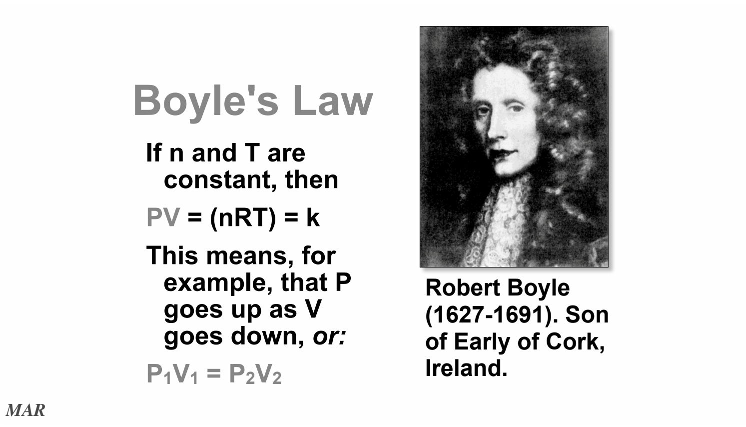

Boyle's Law

If n and T are constant, then

PV = (nRT) = k

This means, for example, that P goes up as V goes down, *or:*

$P_1V_1 = P_2V_2$

Robert Boyle (1627-1691). Son of Early of Cork, Ireland.

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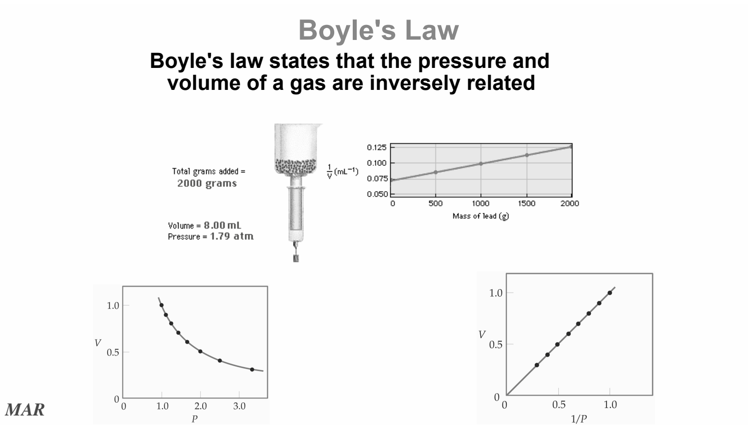

Boyle's Law

Boyle's law states that the pressure and volume of a gas are inversely related

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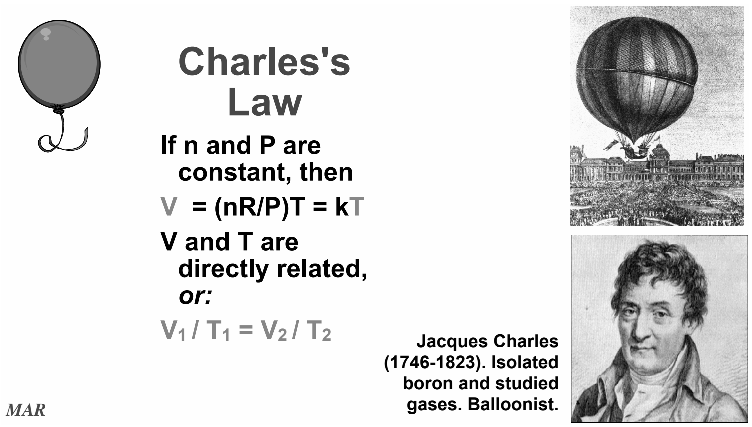


Charles's Law

If n and P are constant, then

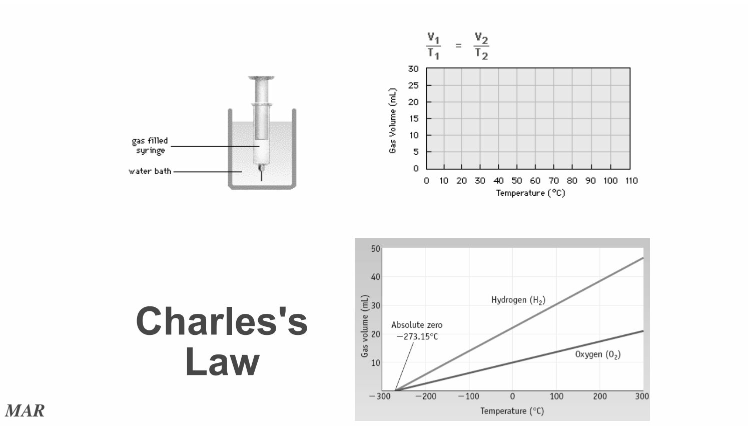
V = (nR/P)T = kT

V and T are directly related, *or:*

$V_1 / T_1 = V_2 / T_2$

Jacques Charles (1746-1823). Isolated boron and studied gases. Balloonist.

MAR



Charles's Law

MAR

Charles’s Law

Balloons immersed in liquid N_2 (at -196 ˚C) will shrink as the air cools (and is liquefied).

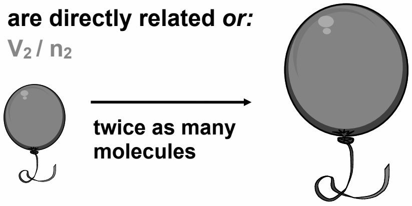
Avogadro's Hypothesis

Equal volumes of gases at the same T and P have the same number of molecules.

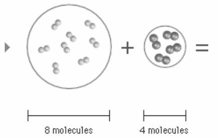
V = (RT/P)n = kn

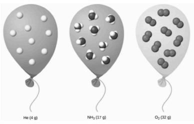
V and n are directly related *or:*

$V_1 / n_1 = V_2 / n_2$



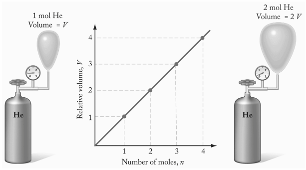
Avogadro's Hypothesis





$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$

The gases in this experiment are all measured at the same T and P.



IDEAL GAS LAW

$$P\,V = n\,R\,T$$

Brings together gas properties.

Can be derived from experiment and theory.

IDEAL GAS LAW

$$P\,V = n\,R\,T$$

The constant of proportionality is known as *R*, the gas constant.

Memorize* R! *Always use 0.082057!

We will also use 8.3145 later...

Units	Numerical Value
L-atm/mol-K	0.082057
J/mol-K*	8.3145
cal/mol-K	1.987
m^3-Pa/mol-K*	8.3145
L-torr/mol-K	62.36

*SI unit

Using PV = nRT

How much N_2 is req'd to fill a small room with a volume of 960. cubic feet ($2.70 * 10^4$ L) to P = 745 mm Hg at 25 ^{o}C?

R = 0.082057 L•atm/K•mol

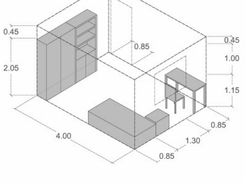
Solution

1. Get all data into proper units

V = $2.70 * 10^4$ L

T = 25 ^{o}C + 273 = 298 K

P = 745 mm Hg (1 atm/760 mm Hg)
= 0.980 atm



Using PV = nRT

How much N_2 is req'd to fill a small room with a volume of 960. cubic feet ($2.70 * 10^4$ L) to P = 745 mm Hg at 25 °C?

R = 0.082057 L•atm/K•mol

Solution

2. Now calc. n = PV / RT

$$n = \frac{(0.980\text{ atm})(2.70 \text{ x } 10^4\text{ L})}{(0.082057\text{ L} \bullet \text{atm/K} \bullet \text{mol})(298\text{ K})}$$

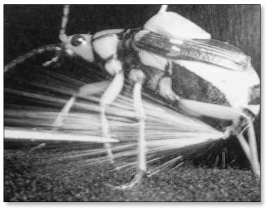
n = 1.08×10^3 mol (30.3 kg of N_2)

MAR

Gases and Stoichiometry

$2\ H_2O_2$(liq) ---> $2\ H_2O$(g) + O_2(g)

Decompose 1.1 g of H_2O_2 in a flask with a volume of 2.50 L. What is the pressure of O_2 at 25 °C? Of H_2O?



Bombardier beetle uses decomposition of hydrogen peroxide to defend itself.

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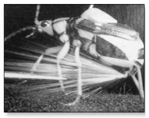

Gases and Stoichiometry

$2\ H_2O_2$(liq) ---> $2\ H_2O$(g) + O_2(g)

Decompose 1.1 g of H_2O_2 in a flask with a volume of 2.50 L. What is the pressure of O_2 at 25 °C? Of H_2O?

Solution

Strategy:

- **Calculate moles of H_2O_2 and then moles of O_2 and H_2O.**
- **Finally, calc. P from n, R, T, and V.**

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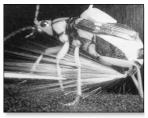

Gases and Stoichiometry

$2\ H_2O_2$(liq) ---> $2\ H_2O$(g) + O_2(g)

Decompose 1.1 g of H_2O_2 in a flask with a volume of 2.50 L. What is the pressure of O_2 at 25 °C? Of H_2O?

Solution

$$1.1\text{ g } H_2O_2 \bullet \frac{1\text{ mol}}{34.0\text{ g}} = 0.032\text{ mol}$$

$$0.032\text{ mol } H_2O_2 \bullet \frac{1\text{ mol } O_2}{2\text{ mol } H_2O_2} = 0.016\text{ mol } O_2$$

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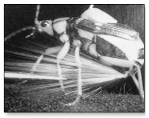

Gases and Stoichiometry

$2\ H_2O_2$(liq) ---> $2\ H_2O$(g) + O_2(g)

Decompose 1.1 g of H_2O_2 in a flask with a volume of 2.50 L. What is the pressure of O_2 at 25 °C? Of H_2O?

Solution

P of O_2 = nRT/V

$$= \frac{(0.016\text{ mol})(0.082057\text{ L} \bullet \text{atm/K} \bullet \text{mol})(298\text{ K})}{2.50\text{ L}}$$

P of O_2 = 0.16 atm

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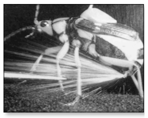


Gases and Stoichiometry

$2\ H_2O_2$(liq) ---> $2\ H_2O$(g) + O_2(g)

Solution

What is P of H_2O? Could calculate as above. But *recall* Avogadro's hypothesis.

V α n at same T and P, *and*

P α n at same T and V

There are 2 times as many moles of H_2O as moles of O_2. P is proportional to n. Therefore, P of H_2O is twice that of O_2.

P of H_2O = 0.32 atm

MAR

Dalton's Law of Partial Pressures

$2\,H_2O_2(liq) \longrightarrow 2\,H_2O(g) + O_2(g)$

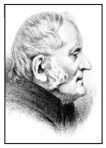
0.32 atm   0.16 atm

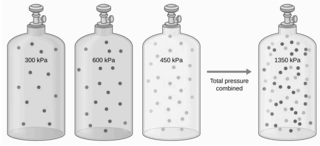
P_{total} in gas mixture = $P_A + P_B + \ldots$

So:

$P_{total} = P(H_2O) + P(O_2) = 0.48$ atm

Dalton's Law: total P equals sum of PARTIAL pressures.





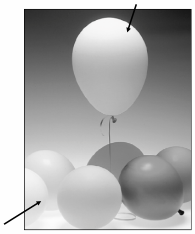
MAR

GAS DENSITY

Density of a gas is proportional to the molar mass

$SF_6(g)$, 146.1 g mol^{-1}

$Br_2(g)$, 159.8 g mol^{-1}



Low density helium

Higher Density air

MAR

GAS DENSITY

PV = nRT

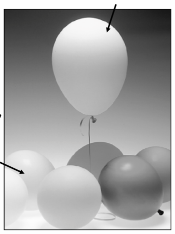
$$\frac{n}{V} = \frac{P}{RT}$$

$$\frac{m}{M \bullet V} = \frac{P}{RT}$$

where M = molar mass

$$d = \frac{m}{V} = \frac{PM}{RT}$$

or **PM = dRT** *(evening dirt equation)*



Low density helium

Higher Density air

MAR

Standard Temperature and Pressure (STP)

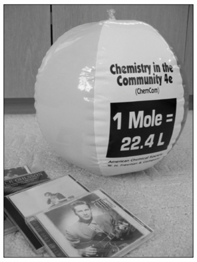
A common reference point used in applications using gases

- **Standard Temperature = 273.15 K**
- **Standard Pressure = 1.000 atm**

and if 1.00 mol of gas used,

- **Standard Volume = 22.4 L**

1.00 mol of an ideal gas occupies 22.4 L at 273 K and 1.00 atm of pressure!



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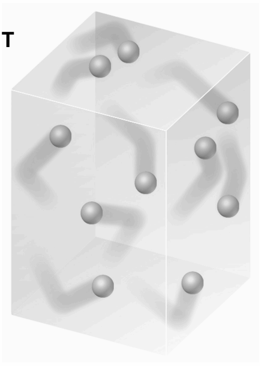
KINETIC MOLECULAR THEORY (KMT)

Theory used to explain gas laws. KMT assumptions are

- **Gases consist of molecules in constant, random motion.**
- **P arises from collisions with container walls.**
- **No attractive or repulsive forces between molecules. Collisions elastic.**
- **Volume of molecules is negligible.**

see Principal Assumptions of KMT Handout



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Kinetic Molecular Theory

We assume molecules of mass **(m, kg/mol)** are in motion **(velocity, v, m/s)**, so they have kinetic energy **(KE, J).**

Molecules at the same temperature **(T, K)** also have the same kinetic energy, so:

$$KE = {}^1/_2 mv^2 = {}^3/_2 RT$$

*Note: this R = 8.3145 J/mol*K ("energy R")*

At the same T, all gases have the same average KE.

As T goes up**,** KE also increases - and so does speed.

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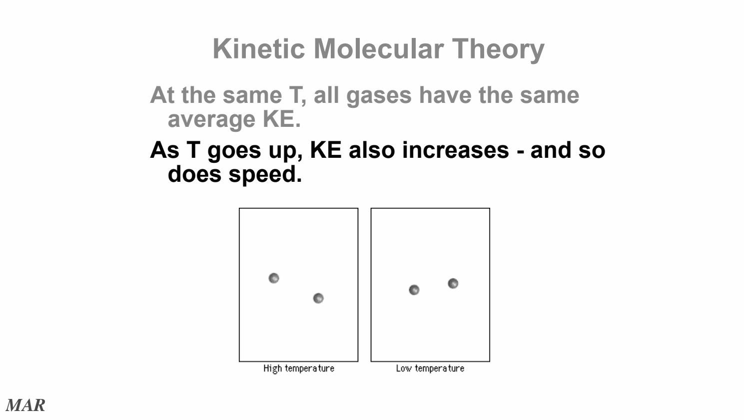

Kinetic Molecular Theory

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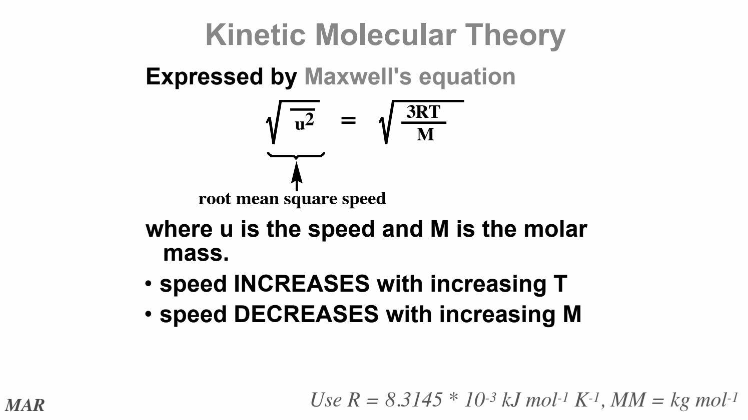
MAR



Kinetic Molecular Theory

Expressed by Maxwell's equation

$$\underbrace{\sqrt{u^2}}_{\uparrow\ \text{root mean square speed}} = \sqrt{\frac{3RT}{M}}$$

where u is the speed and M is the molar mass.

- **speed INCREASES with increasing T**
- **speed DECREASES with increasing M**

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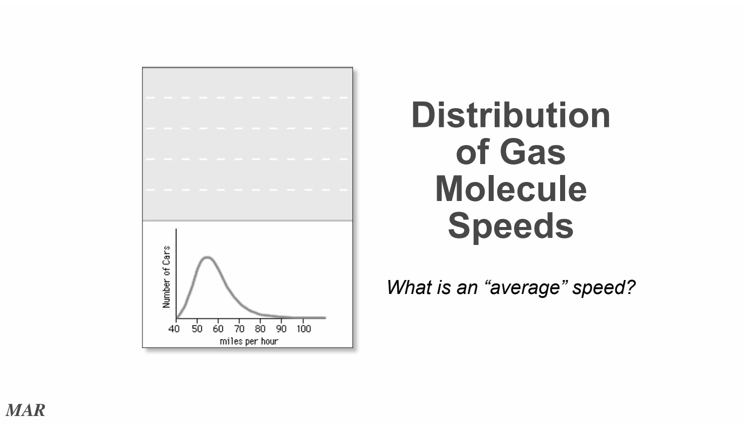
Use $R = 8.3145 * 10^{-3}$ kJ mol^{-1} K^{-1}, MM = kg mol^{-1}



Distribution of Gas Molecule Speeds

What is an "average" speed?

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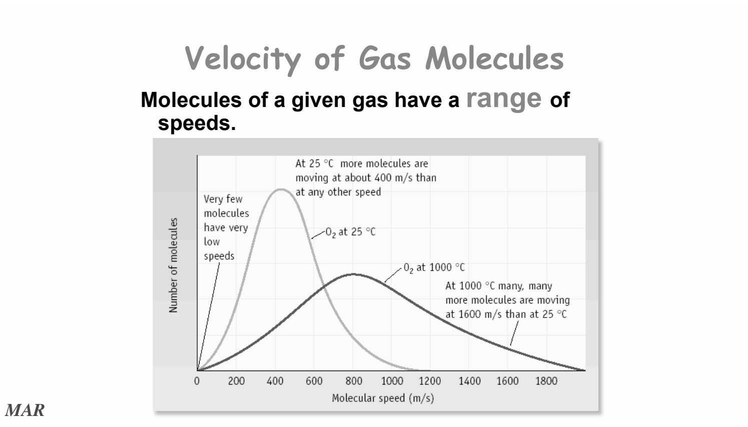

Velocity of Gas Molecules

Molecules of a given gas have a range of speeds.

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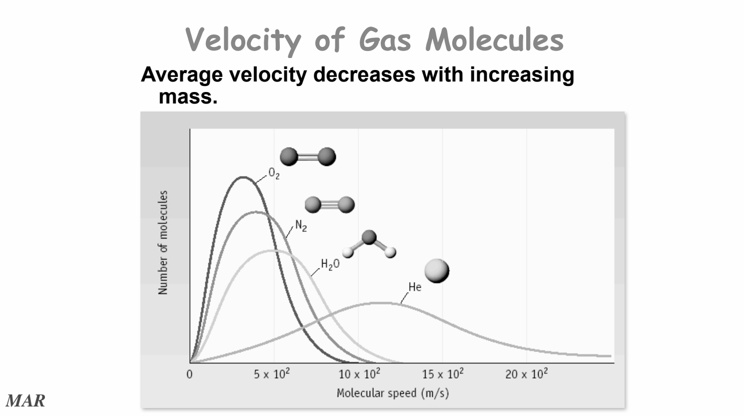

Velocity of Gas Molecules

Average velocity decreases with increasing mass.

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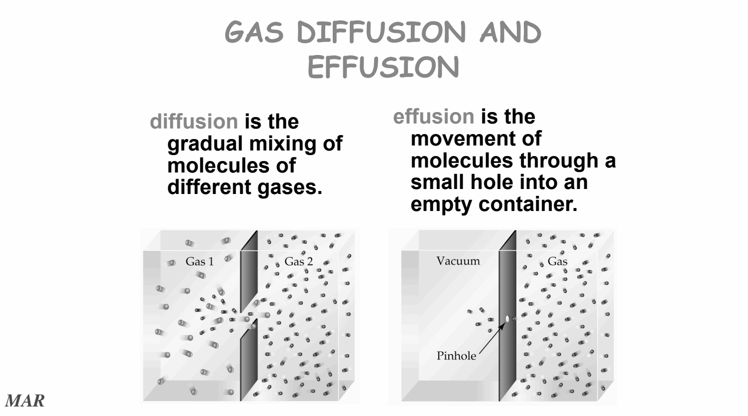


GAS DIFFUSION AND EFFUSION

diffusion is the gradual mixing of molecules of different gases.

effusion is the movement of molecules through a small hole into an empty container.

MAR

Molecules effuse through holes in a rubber balloon, for example, at a rate (= moles/time) that is

- **proportional to T**
- **inversely proportional to M.**

Therefore, He effuses more rapidly than O_2 at same T.

GAS DIFFUSION AND EFFUSION

GAS DIFFUSION AND EFFUSION

Graham's law governs effusion and diffusion of gas molecules.

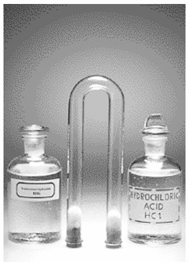
$$\frac{\text{Rate for A}}{\text{Rate for B}} = \sqrt{\frac{\text{M of B}}{\text{M of A}}}$$

Rate of effusion is inversely proportional to its molar mass.

Thomas Graham, 1805-1869. Professor in Glasgow and London.

Gas Diffusion

relation of mass to rate of diffusion



Gaseous diffusion of $NH_3(g)$ and $HCl(g)$

- **HCl and NH_3 diffuse from opposite ends of tube.**
- **Gases meet to form NH_4Cl**
- **HCl heavier than NH_3**
- **Therefore, NH_4Cl forms closer to HCl end of tube.**

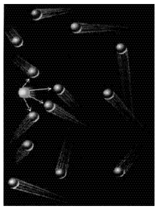
Deviations from Ideal Gas Law

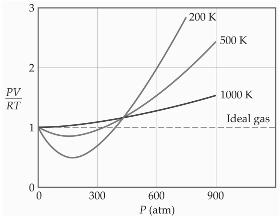
Real molecules have volume.

There are intermolecular forces.

Otherwise a gas could not become a liquid.

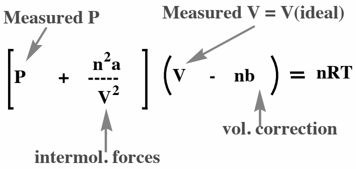
High Pressure and Low Temperature conditions show greatest deviation



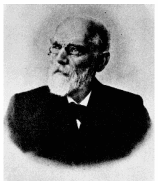


Deviations from Ideal Gas Law

Account for volume of molecules and intermolecular forces with VAN DER WAAL'S EQUATION.



$$\left[P + \frac{n^2a}{V^2}\right]\left(V - nb\right) = nRT$$



J. van der Waals, 1837-1923, Professor of Physics, Amsterdam. Nobel Prize 1910.

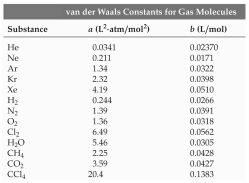
Deviations from Ideal Gas Law

Cl_2 gas has a = 6.49, b = 0.0562

For 8.0 mol Cl_2 in a 4.0 L tank at 27 °C.

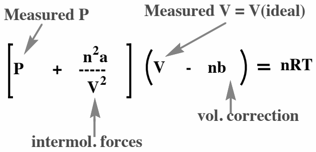
P (ideal) = nRT/V = 49.3 atm

P (van der Waals) = 29.5 atm



van der Waals Constants for Gas Molecules

Substance	a (L^2-atm/mol^2)	b (L/mol)
He	0.0341	0.02370
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0510
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
H_2O	5.46	0.0305
CH_4	2.25	0.0428
CO_2	3.59	0.0427
CCl_4	20.4	0.1383



$$\left[P + \frac{n^2a}{V^2}\right]\left(V - nb\right) = nRT$$

End of Chapter 9

"Well OF COURSE there are greenhouse gasses in here. It's a freakin' GREENHOUSE!"

See:

- *Chapter Nine Study Guide*
- *Chapter Nine 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 to use the gas laws, desired units for the gas law, STP uses, Dalton's Law of Partial Pressure, etc.**
- **understand pressure**
- **know how to use gases in stoichiometry problems**
- **know how the KMT (Kinetic Molecular Theory) describes gases**

- **PV = nRT**
- **PM = dRT**
- **mole = 6.022×10^{23}**
- **760 mm Hg = 1 atm**
- **1013 mbar = 1 atm**
- *metric prefixes* **(m, k,** *etc.***)**
- **STP = 1 atm, 273.15 K**

R = 0.082057 L atm mol^{-1} K^{-1} *(the "gas R")*
R = 8.3145 J mol^{-1} K^{-1} *(the "energy R")*

KE = $^1/_2mv^2 = {}^3/_2RT$

MAR

End of Chapter Problems: ***Test Yourself***

1. A sample of nitrogen gas has a pressure of 67.5 mm Hg in a 500. mL flask. What is the pressure of this gas sample when it is transferred to a 125 mL flask at the same temperature?
2. You have 3.5 L of NO at a temperature of 22.0 °C. What volume would the NO occupy at 37 °C? (Assume the pressure is constant.)
3. An automobile cylinder has a volume of 400. cm^3. The engine takes in air at a pressure of 1.00 atm and a temperature of 15 °C and compresses the air to a volume of 50.0 cm^3 at 77 °C. What is the final pressure of the gas in the cylinder?
4. A 1.25 g sample of CO_2 is contained in a 750. mL flask at 22.5 °C. What is the pressure of the gas?
5. A gaseous organofluorine compound has a density of 0.355 g/L at 17 °C and 189 mm Hg. What is the molar mass of the compound?
6. Sodium azide, the explosive compound in automobile air bags, decomposes according to the following equation:
   $2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$
   What mass of sodium azide is required to provide the nitrogen needed to inflate a 75.0 L bag to a pressure of 1.3 atm at 25 °C?

MAR

End of Chapter Problems: ***Answers***

1. 270. mm Hg
2. 3.7 L
3. 9.72 atm
4. 0.919 atm
5. 34.0 g/mol
6. 170 g

Be sure to view practice problem set #3 and self quizzes for ***nomenclature*** *examples and practice*

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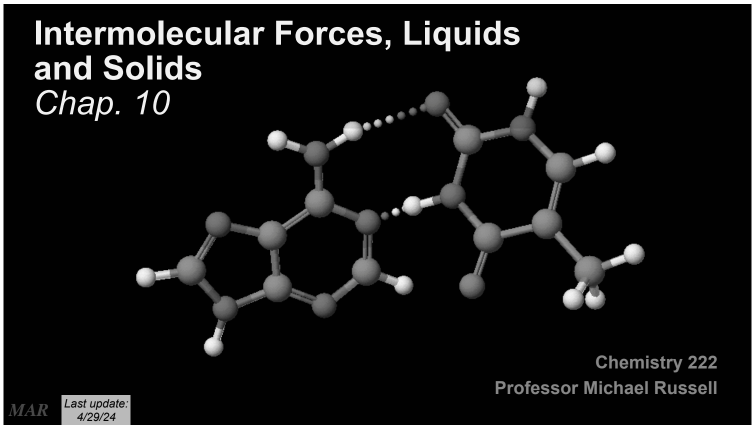


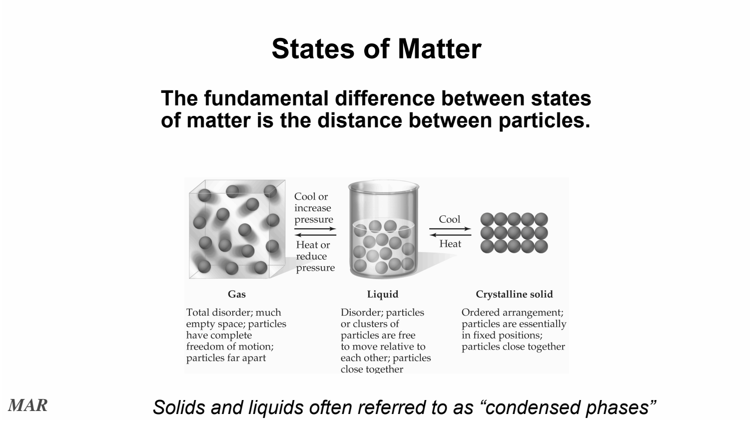
Intermolecular Forces, Liquids and Solids

Chap. 10

Chemistry 222

Professor Michael Russell

Last update: 4/29/24

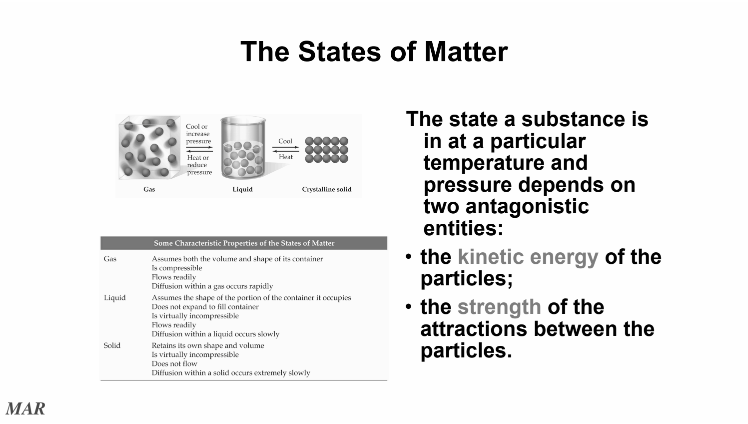


States of Matter

The fundamental difference between states of matter is the distance between particles.

Gas	Liquid	Crystalline solid
Total disorder; much empty space; particles have complete freedom of motion; particles far apart	Disorder; particles or clusters of particles are free to move relative to each other; particles close together	Ordered arrangement; particles are essentially in fixed positions; particles close together

Solids and liquids often referred to as "condensed phases"

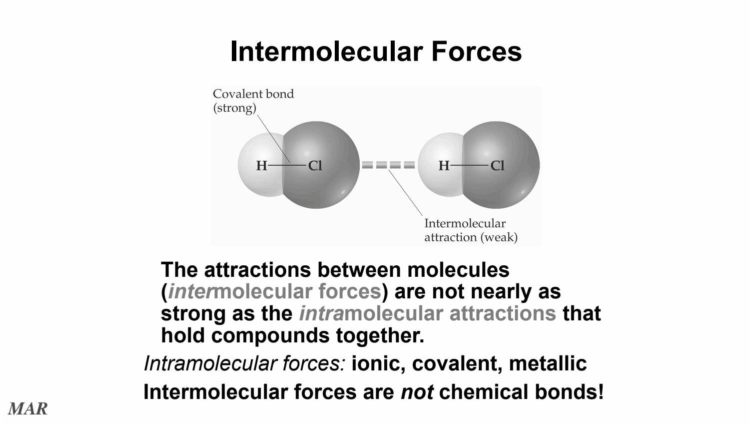


The States of Matter

Some Characteristic Properties of the States of Matter	
Gas	Assumes both the volume and shape of its container Is compressible Flows readily Diffusion within a gas occurs rapidly
Liquid	Assumes the shape of the portion of the container it occupies Does not expand to fill container Is virtually incompressible Flows readily Diffusion within a liquid occurs slowly
Solid	Retains its own shape and volume Is virtually incompressible Does not flow Diffusion within a solid occurs extremely slowly

The state a substance is in at a particular temperature and pressure depends on two antagonistic entities:

- **the kinetic energy of the particles;**
- **the strength of the attractions between the particles.**

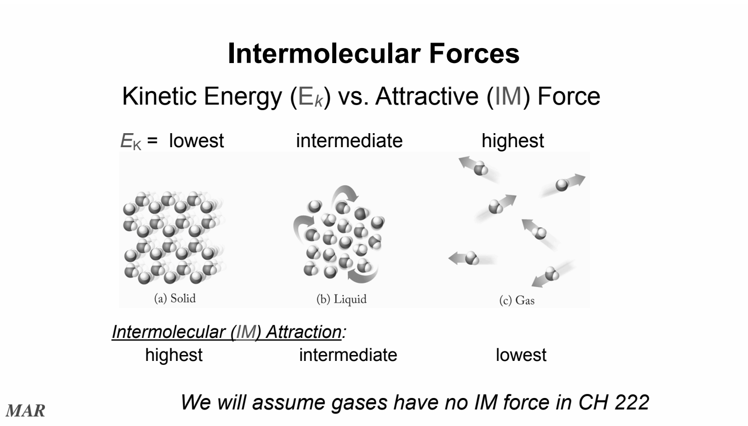


Intermolecular Forces

The attractions between molecules (*inter*molecular forces) are not nearly as strong as the *intra*molecular attractions that hold compounds together.

Intramolecular forces: **ionic, covalent, metallic**

Intermolecular forces are *not* chemical bonds!



Intermolecular Forces

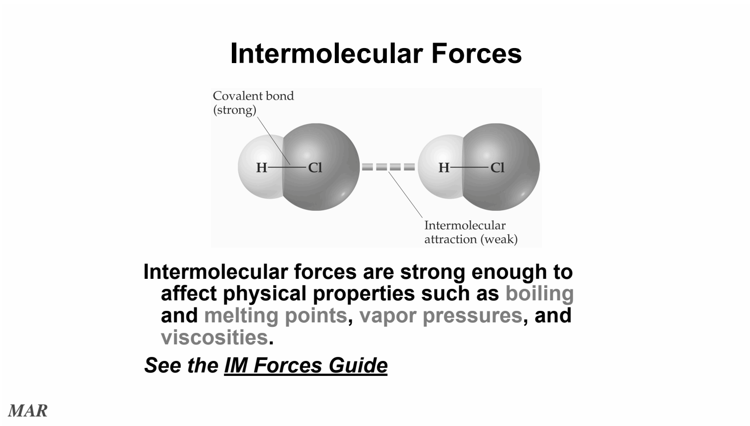
Kinetic Energy (E_k) vs. Attractive (IM) Force

E_K = lowest intermediate highest

<u>*Intermolecular (IM) Attraction*</u>*:*

highest intermediate lowest

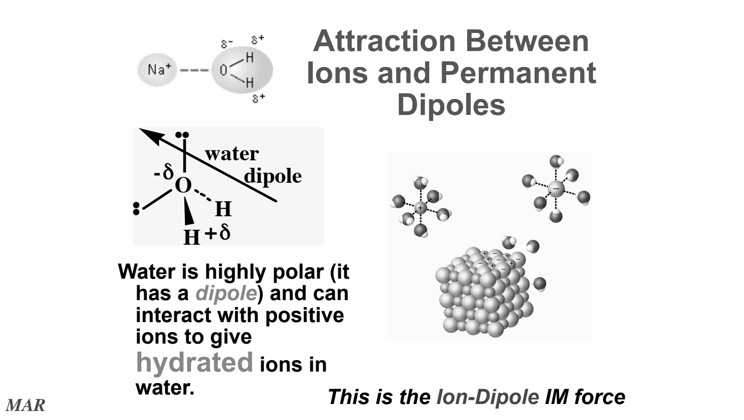
We will assume gases have no IM force in CH 222



Intermolecular Forces

Intermolecular forces are strong enough to affect physical properties such as boiling and melting points, vapor pressures, and viscosities.

See the IM Forces Guide

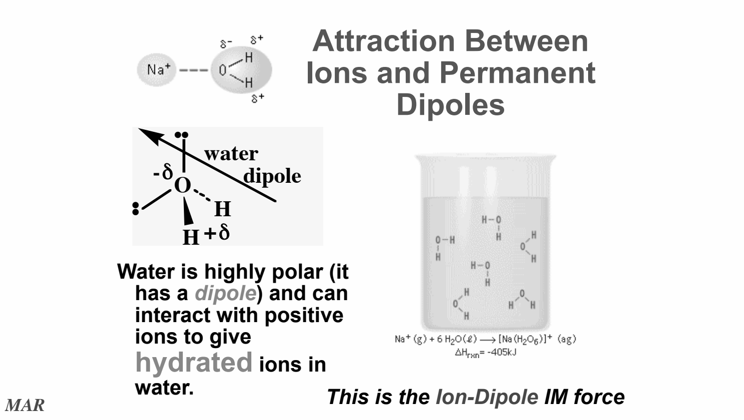


Attraction Between Ions and Permanent Dipoles

Water is highly polar (it has a *dipole*) and can interact with positive ions to give hydrated ions in water.

This is the Ion-Dipole IM force

MAR



Attraction Between Ions and Permanent Dipoles

$Na^+ (g) + 6\ H_2O(\ell) \longrightarrow [Na(H_2O)_6]^+ \ (aq)$

$\Delta H_{rxn} = -405 kJ$

Water is highly polar (it has a *dipole*) and can interact with positive ions to give hydrated ions in water.

This is the Ion-Dipole IM force

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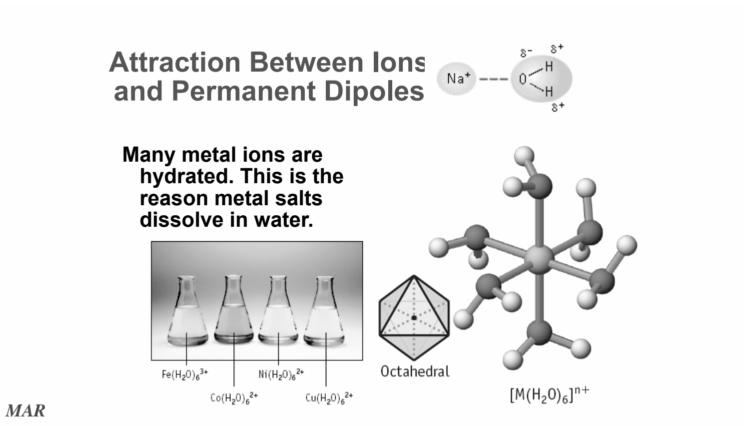

Attraction Between Ions and Permanent Dipoles

Many metal ions are hydrated. This is the reason metal salts dissolve in water.

$Fe(H_2O)_6^{3+}$ $Ni(H_2O)_6^{2+}$ $Co(H_2O)_6^{2+}$ $Cu(H_2O)_6^{2+}$

$[M(H_2O)_6]^{n+}$

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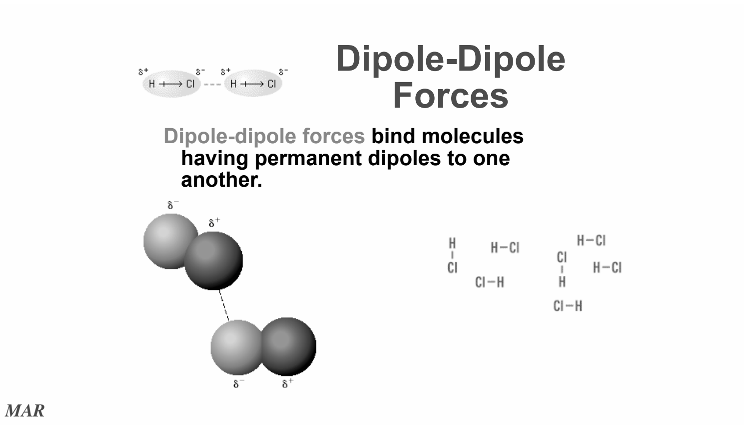

Dipole-Dipole Forces

Dipole-dipole forces bind molecules having permanent dipoles to one another.

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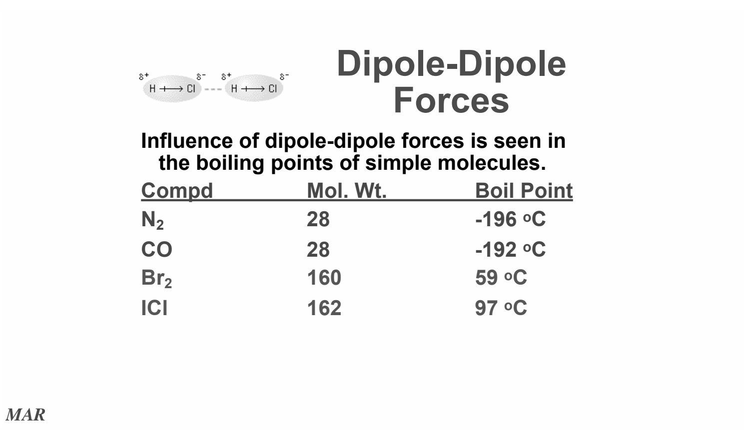

Dipole-Dipole Forces

Influence of dipole-dipole forces is seen in the boiling points of simple molecules.

Compd	Mol. Wt.	Boil Point
N_2	28	-196 °C
CO	28	-192 °C
Br_2	160	59 °C
ICl	162	97 °C

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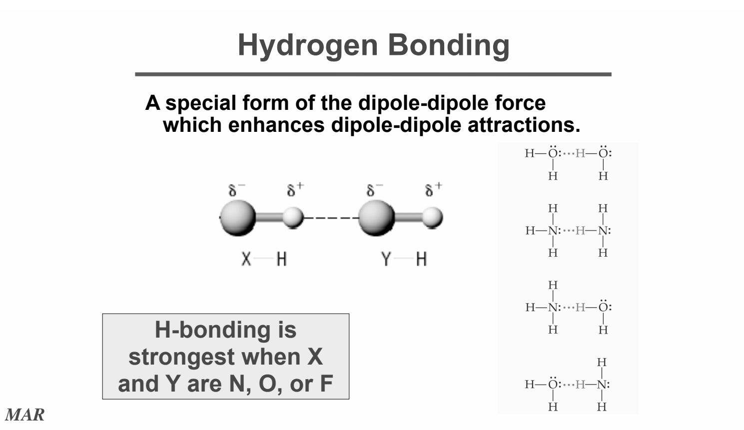

Hydrogen Bonding

A special form of the dipole-dipole force which enhances dipole-dipole attractions.

H-bonding is strongest when X and Y are N, O, or F

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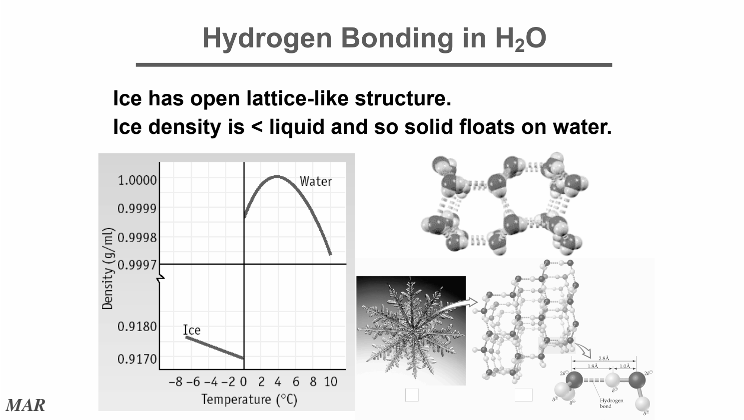

Hydrogen Bonding in H_2O

Ice has open lattice-like structure.
Ice density is < liquid and so solid floats on water.

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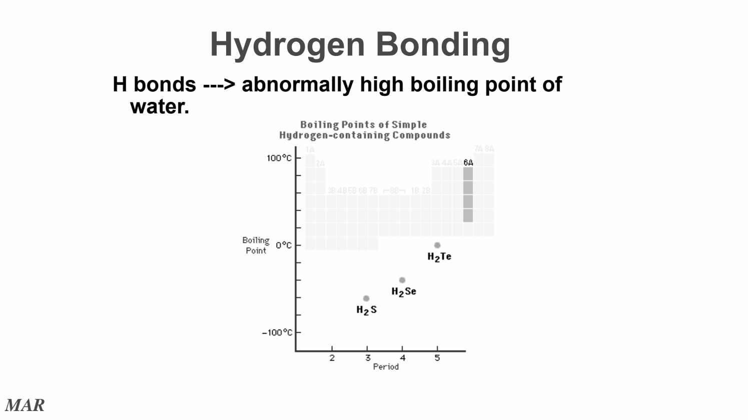

Hydrogen Bonding

H bonds ---> abnormally high boiling point of water.

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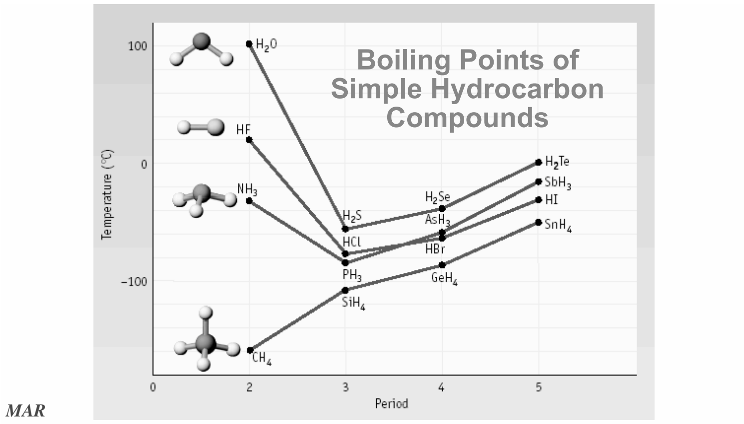

Boiling Points of Simple Hydrocarbon Compounds

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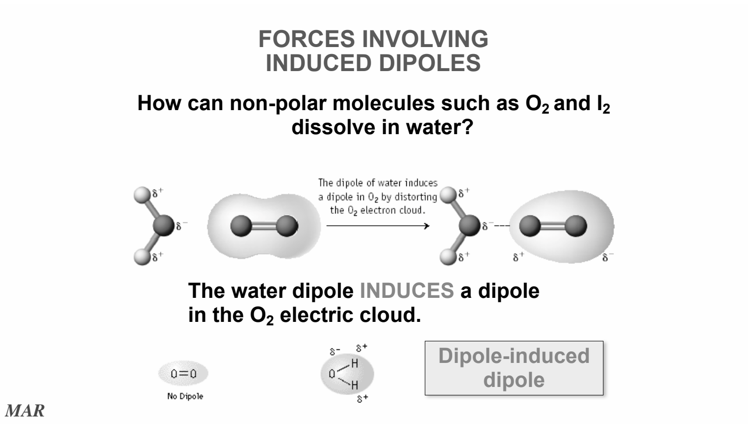

FORCES INVOLVING INDUCED DIPOLES

How can non-polar molecules such as O_2 and I_2 dissolve in water?

The dipole of water induces a dipole in O_2 by distorting the O_2 electron cloud.

The water dipole INDUCES a dipole in the O_2 electric cloud.

Dipole-induced dipole

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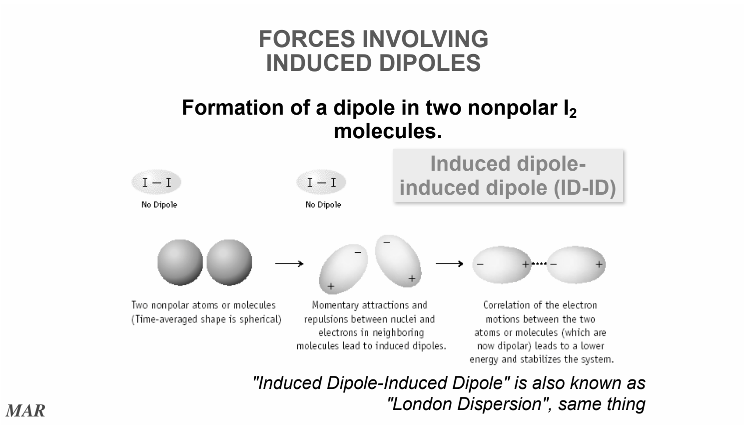


FORCES INVOLVING INDUCED DIPOLES

Formation of a dipole in two nonpolar I_2 molecules.

Induced dipole-induced dipole (ID-ID)

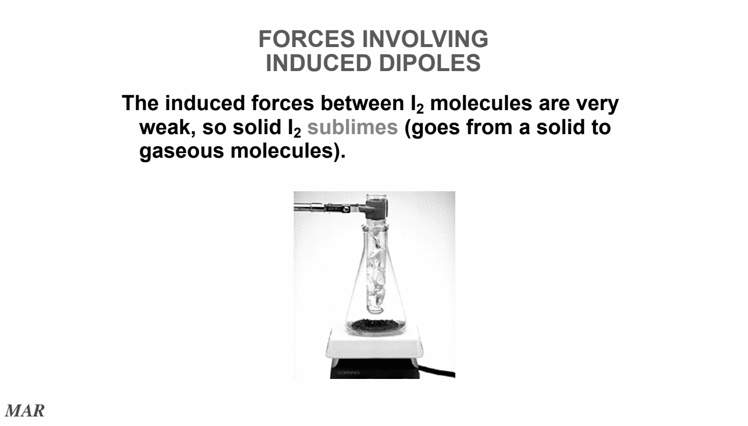
Two nonpolar atoms or molecules (Time-averaged shape is spherical)

Momentary attractions and repulsions between nuclei and electrons in neighboring molecules lead to induced dipoles.

Correlation of the electron motions between the two atoms or molecules (which are now dipolar) leads to a lower energy and stabilizes the system.

"Induced Dipole-Induced Dipole" is also known as "London Dispersion", same thing

MAR



FORCES INVOLVING INDUCED DIPOLES

The induced forces between I_2 molecules are very weak, so solid I_2 sublimes (goes from a solid to gaseous molecules).

MAR

FORCES INVOLVING INDUCED DIPOLES

The magnitude of the induced dipole depends on the tendency to be distorted.

Higher molar mass ---> larger induced dipoles.

Larger atoms have larger electron clouds which are easier to polarize

Halogen	Molecular Weight (amu)	Boiling Point (K)	Noble Gas	Molecular Weight (amu)	Boiling Point (K)
F_2	38.0	85.1	He	4.0	4.6
Cl_2	71.0	238.6	Ne	20.2	27.3
Br_2	159.8	332.0	Ar	39.9	87.5
I_2	253.8	457.6	Kr	83.8	120.9
			Xe	131.3	166.1

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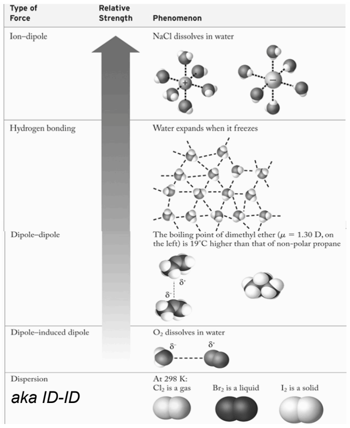

Type of Force	Relative Strength	Phenomenon
Ion–dipole		NaCl dissolves in water
Hydrogen bonding		Water expands when it freezes
Dipole–dipole		The boiling point of dimethyl ether (μ = 1.30 D, on the left) is 19°C higher than that of non-polar propane
Dipole–induced dipole		O_2 dissolves in water
Dispersion *aka ID-ID*		At 298 K: Cl_2 is a gas Br_2 is a liquid I_2 is a solid

Intermolecular Forces Strength Summary

Ion-ion / metallic force strongest of all

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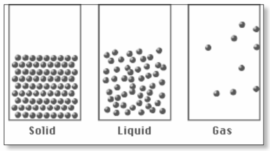
Liquids



In a liquid

- **molecules are in constant motion**
- **there are appreciable intermolecular forces**
- **molecules close together**
- **Liquids are almost incompressible**
- **Liquids do not fill the container**

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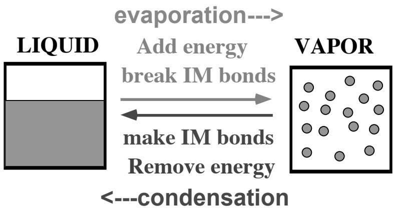
Liquids

The two key properties we need to describe are EVAPORATION and its opposite- CONDENSATION



evaporation--->

LIQUID

Add energy

break IM bonds

VAPOR

make IM bonds

Remove energy

<---condensation

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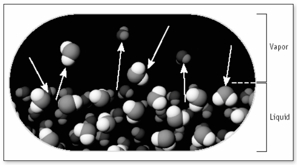
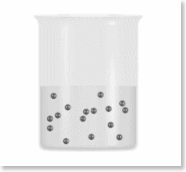
Liquids

To evaporate, molecules must have sufficient energy to break IM forces.





Breaking IM forces requires energy. The process of evaporation is endothermic.

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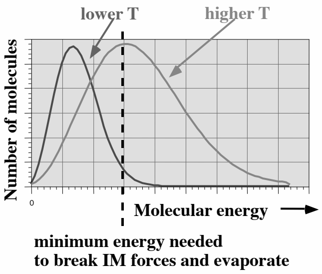
Liquids



Distribution of molecular energies in a liquid.

Kinetic Energy proportional to Temperature

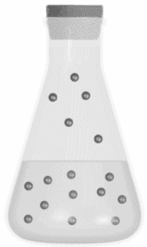
At higher T a much larger number of molecules has high enough energy to break IM forces and move from liquid to vapor state.

MAR

Liquids

When molecules of liquid are in the vapor state, they exert a VAPOR PRESSURE

EQUILIBRIUM VAPOR PRESSURE is the pressure exerted by a vapor over a liquid in a closed container when the rate of evaporation = the rate of condensation.



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Vapor Pressure

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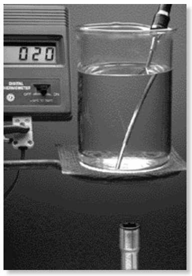
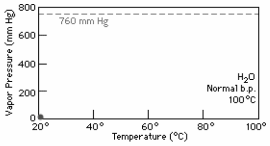
Boiling Liquids





Liquid boils when its vapor pressure equals atmospheric pressure.

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Boiling Point at Lower Pressure

When pressure is lowered, the vapor pressure can equal the external pressure at a lower temperature.

MAR

Consequences of Vapor Pressure Changes

When can cools, vp of water drops. Pressure in the can is less than that of atmosphere, so can is crushed.

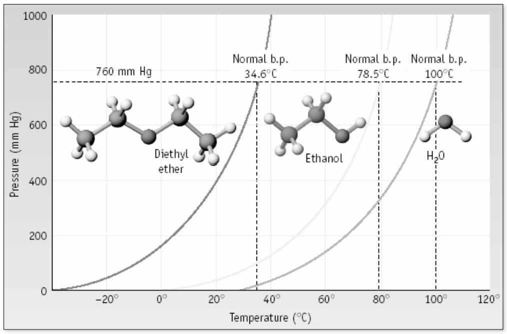
MAR

Consequences of Vapor Pressure Changes - *Whoops!*

When car cools on hot day (i.e. cleaning with cold water), vp of fumes inside drops. Pressure in the car is less than that of atmosphere, so car is crushed!

MAR

Equilibrium Vapor Pressure



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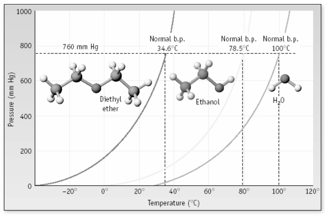
Liquids



The curves show all conditions of P and T where LIQ and VAP are in EQUILIBRIUM.

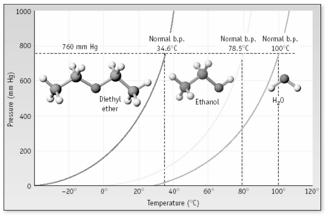
The VP rises with T.

When VP = external P, the liquid boils.

This means that BPs of liquids change with altitude.

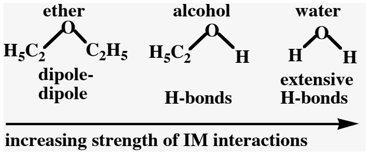
MAR

Liquids



If external P = 760 mm Hg, T of boiling is the NORMAL BOILING POINT

VP of a given molecule at a given T depends on IM forces. Here the VPs are in the order:



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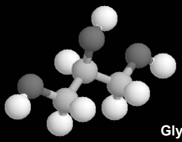
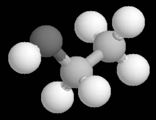
Viscosity

VISCOSITY is the tendency or resistance of liquids to flow.

Liquids "flow" differently due to the strength of their intermolecular bonds



Glycerol



Ethanol

Viscosity results from several factors, including IM interactions, molecular shape and size

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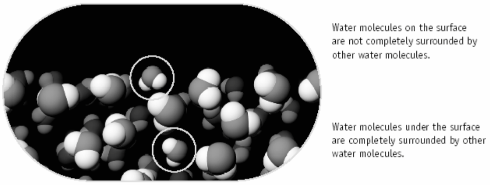
Liquids

Molecules at surface behave differently than those in the interior.



Water molecules on the surface are not completely surrounded by other water molecules.

Water molecules under the surface are completely surrounded by other water molecules.

Molecules at surface experience net INWARD force of attraction.
This leads to SURFACE TENSION - the energy required to break the surface.

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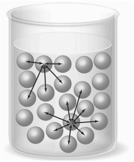
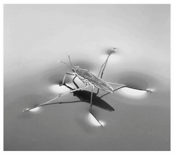
Surface Tension





SURFACE TENSION also leads to spherical liquid droplets.

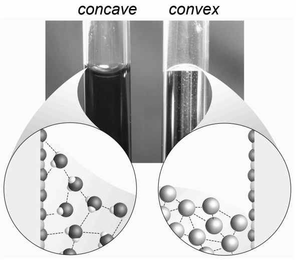
MAR

Liquids

IM forces also lead to CAPILLARY action

Cohesive forces: interactions between like particles.

Adhesive forces: interactions between unlike particles.



"Concave up is like a cup, concave down is like a frown"

Concave *(concave up)* Meniscus: adhesive forces ≥ cohesive forces (H_2O on glass)

Convex *(concave down)* Meniscus: Cohesive forces > adhesive forces (Hg on glass).

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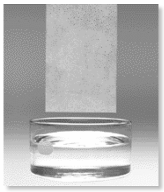
Capillary Action



Movement of water up a piece of paper depends on H-bonds between H_2O and the OH groups of the cellulose in the paper.

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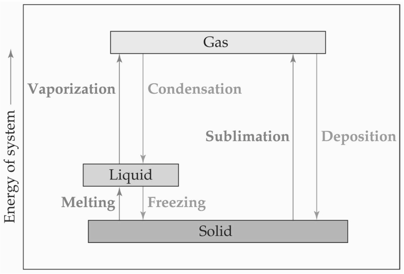
Heat Transfer with Phase Change

Overall patterns:

solid → liquid → gas = *endothermic reaction*

gas → liquid → solid = exothermic reaction



Chapter 5: Heat Transfer with no Phase Change ($q = mC\Delta T$)

MAR

Heat Transfer with Phase Change

HEAT OF VAPORIZATION is the heat required (at constant P) to *vaporize* a liquid.

LIQ + heat ---> *VAP*

Compd.	ΔH_{vap} (kJ/mol)	IM Force
H_2O	**40.7 (100 °C)**	**H-bonds**
SO_2	**26.8 (-47 °C)**	**dipole**
Xe	**12.6 (-107 °C)**	**induced dipole**

HEAT OF FUSION is the heat required (at constant P) to *melt* a solid.

SOL + heat ---> LIQ

Temperature constant during phase change

MAR

Clausius-Clapeyron

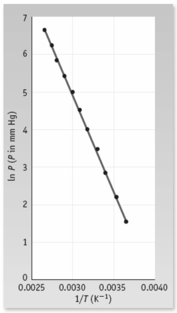
The **Clausius–Clapeyron Equation** provides a link between **vapor pressure (*P*), temperature (in *K*),** and **molar heat of vaporization (ΔH_{vap}):**

$$\ln P = -\frac{\Delta H^\circ_{vap}}{RT} + C$$

Perform a ***linear regression*** (**ln P** *vs.* **1/T(K)**) to get best values, or, with only two temps:

$$\ln \frac{P_1}{P_2} = \frac{\Delta H^\circ_{vap}}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

R = 8.3145 J mol^{-1} K^{-1}



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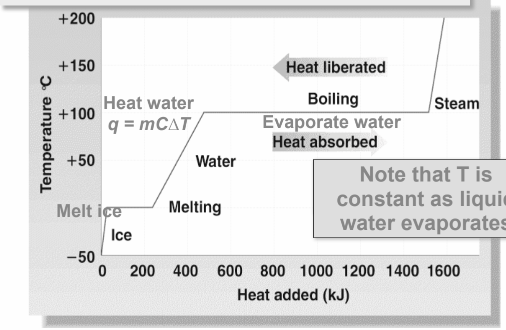
Heating/Cooling Curve for Water



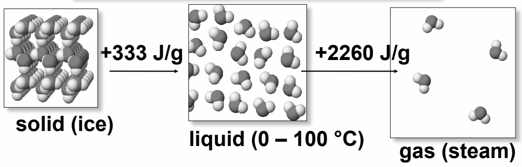
Note that T is constant as liquid water evaporates

MAR

Heat & Changes of State

What quantity of heat is required to melt 500. g of ice at 0.0 ^{o}C and heat the water to steam at 100. ^{o}C?

Heat of fusion of ice = 333 J/g
Specific heat of water = 4.184 J/g•K
Heat of vaporization = 2260 J/g



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Heat & Changes of State

What quantity of heat is required to melt 500. g of ice at 0.0 ^{o}C and heat the water to steam at 100. ^{o}C?

1. **To melt ice**

   $q = (500.\ g)(333\ J/g) = 1.67 \times 10^5\ J$

2. **To raise water from 0.0 ^{o}C to 100. ^{o}C**

   $q = (500.\ g)(4.184\ J/g\bullet K)(100. - 0)K = 2.09 \times 10^5\ J$

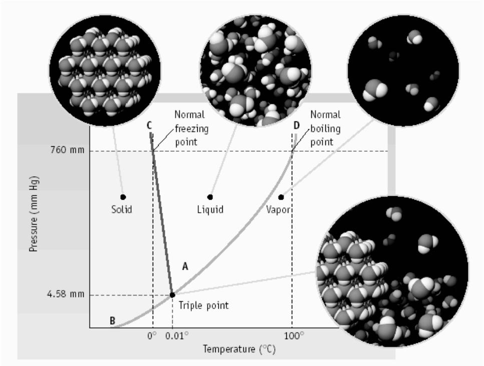
3. **To evaporate water at 100. ^{o}C**

   $q = (500.\ g)(2260\ J/g) = 1.13 \times 10^6\ J$

4. **Total heat energy = 1.51×10^6 J = 1510 kJ**

MAR

Phase Diagrams



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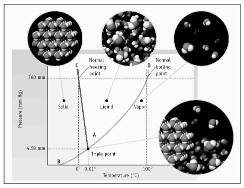

TRANSITIONS BETWEEN PHASES

See the phase diagram for water in text

Lines connect all conditions of T and P where EQUILIBRIUM exists between the phases on either side of the line.

At equilibrium particles move from liquid to gas as fast as they move from gas to liquid.

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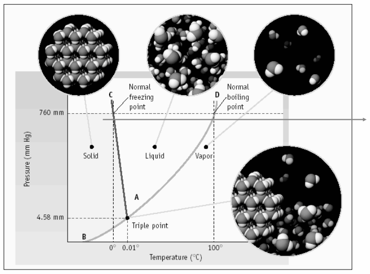
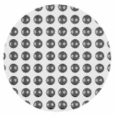
Phase Diagram for Water

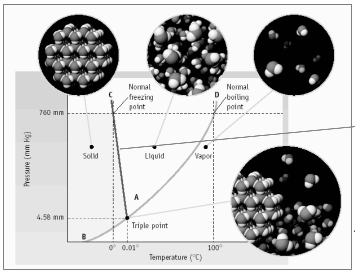


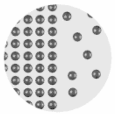


Animation of solid phase.

MAR

Phase Diagram for Water





Animation of equilibrium between solid and liquid phases.

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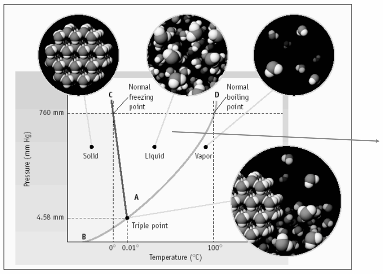
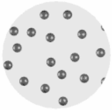
Phase Diagram for Water





Animation of liquid phase.

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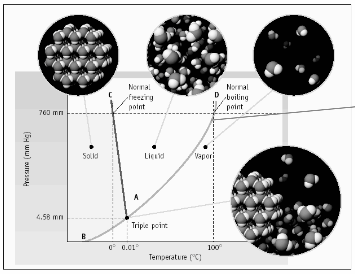
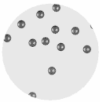
Phase Diagram for Water





Animation of equilibrium between liquid and gas phases.

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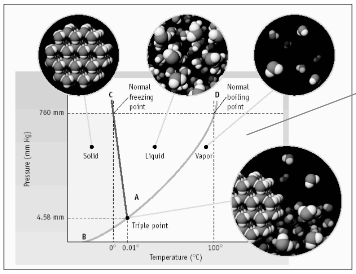
Phase Diagram for Water



Animation of gas phase.

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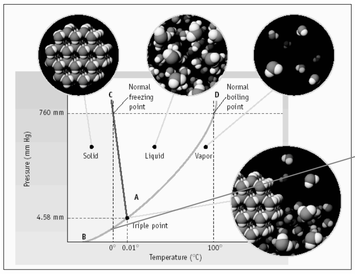
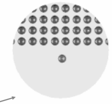
Phase Diagram for Water





Animation of equilibrium between solid and gas phases.

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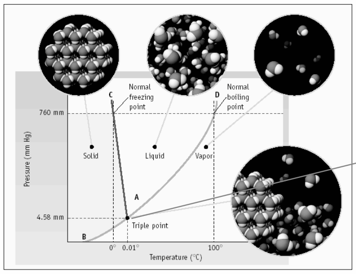
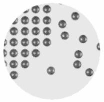
Phase Diagram for Water



Animation of triple point.



At the TRIPLE POINT all three phases are in equilibrium.

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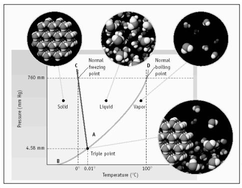


Phases Diagrams-

Important Points for Water

	T(˚C)	P(mm Hg)
Normal boil point	100	760
Normal freeze point	0	760
Triple point	**0.0098**	**4.58**

Water at the Triple Point

MAR

An Introduction to Solids

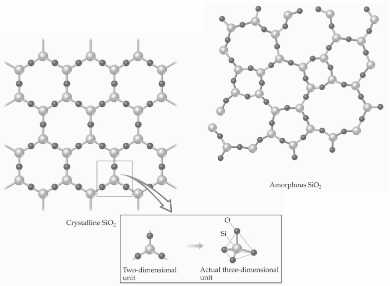
We can think of solids as falling into two groups:

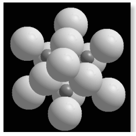
- crystalline: particles in highly ordered arrangements
- amorphous: no particular order in arrangement of particles.

We will focus on crystalline solids in CH 222.

Molecules, atoms or ions locked into a CRYSTAL LATTICE

Particles are close together with very strong IM forces. They are highly ordered, rigid, incompressible





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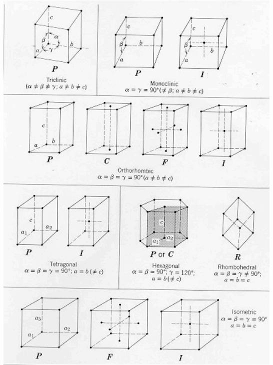
A UNIT CELL is the smallest repeating unit in a crystal lattice



Crystal Lattices

7 Brevais lattice (unit cell) types:

- Triclinic
- Monoclinic
- Orthorhombic
- Tetragonal
- Hexagonal
- Rhombohedral
- Cubic (Isometric)

We will use just the cubic system in CH 222

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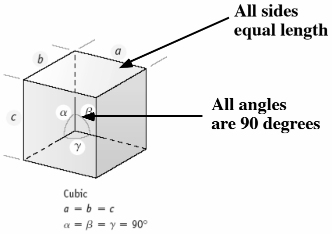
Cubic Unit Cells

There are 7 basic crystal systems, but we are only concerned with CUBIC (*isometric*) in CH 222.



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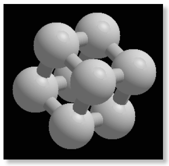
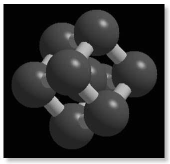
Cubic Unit Cells

Three types of Cubic Unit Cells:

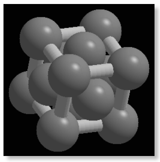
Simple cubic (SC)



Body-centered cubic (BCC)



Face-centered cubic (FCC)



See Cubic Unit Cells Guide

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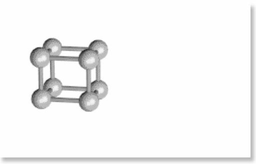
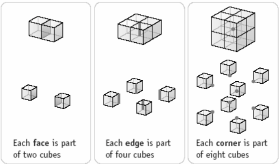
Simple Cubic Unit Cell





Simple cubic unit cell.

Note that each atom is at a corner of a unit cell and is shared among 8 unit cells.

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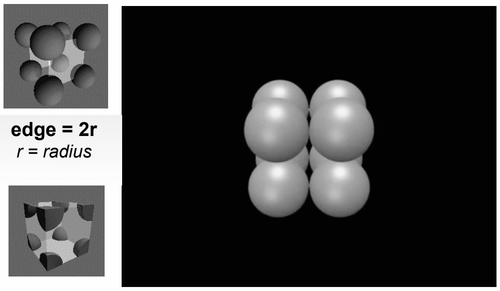
The Simple Cubic Unit Cell



edge = 2r
r = radius

Atom at each corner,
Only 1 net atom per simple cubic cell

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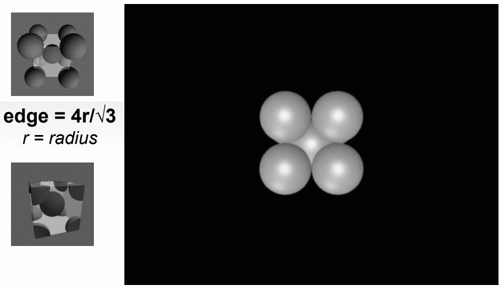
Body-Centered Cubic Unit Cell



edge = 4r/√3
r = radius

Atom at each cube corner plus one in center
Two net atoms per bcc unit cell

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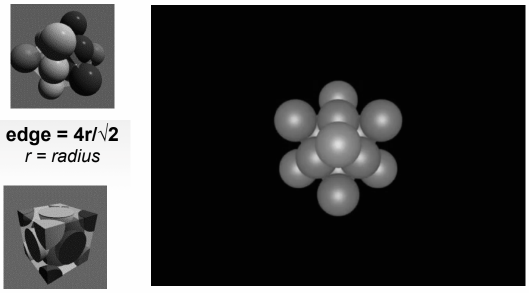
Face-Centered Cubic Unit Cell

also known as cubic close packing



edge = 4r/√2
r = radius

Atom in each cube corner plus atom in each cube face, four net atoms per fcc unit cell

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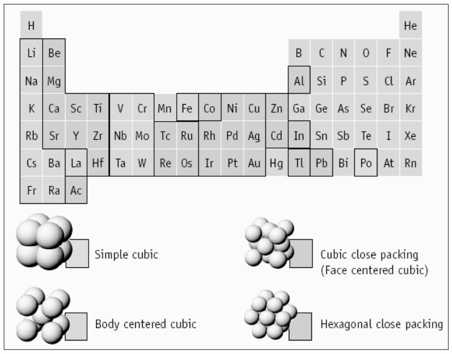
Unit Cells for Metals



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Finding the Lattice Type

To find out if a metal is SC, BCC or FCC, use the known radius and density of an atom to calc. no. of atoms per unit cell.

PROBLEM **Al has density = 2.699 g/cm^3 and Al radius = 143 pm. Verify that Al is FCC.**

SOLUTION

1. **Calc. unit cell** edge (cm)
   see ***handout:*** edge = 4 * radius / √2
   edge = 4 * 143 pm / √2 = 404 pm
   404 pm * (10^{-10} cm / pm) = 4.04 * 10^{-8} cm

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Finding the Lattice Type

PROBLEM **Al has density = 2.699 g/cm^3 and Al radius = 143 pm. Verify that Al is FCC.**

SOLUTION

2. **Calc. unit cell** volume
   edge = 4.04 x 10^{-8} cm ***(previous slide)***
   V = (cell edge)^3 = (4.04 x 10^{-8} cm)^3
   V = 6.62 x 10^{-23} cm^3
3. **Now use density to find mass**
   mass = (6.62 x 10^{-23} cm^3)(2.699 g/cm^3)
   = 1.79 x 10^{-22} g/unit cell

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Finding the Lattice Type

PROBLEM **Al has density = 2.699 g/cm^3 and Al radius = 143 pm. Verify that Al is FCC.**

SOLUTION

4. **Calculate number of Al per unit cell from mass of unit cell.**

$$\text{Mass 1 Al atom} = \frac{26.98\ \text{g}}{\text{mol}} \bullet \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{atoms}}$$

1 atom = 4.480 x 10^{-23} g, ***so***

$$\frac{1.79 \times 10^{-22}\ \text{g}}{\text{unit cell}} \bullet \frac{1\ \text{atom}}{4.480 \times 10^{-23}\ \text{g}} = 3.99\ \text{Al atoms/unit cell}$$

MAR

...more in lab and problem set

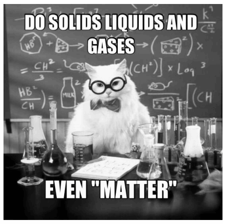
Type of Solid	Form of Unit Particles	Forces Between Particles	Properties	Examples
Molecular	Atoms or molecules	London dispersion forces, dipole–dipole forces, hydrogen bonds	Fairly soft, low to moderately high melting point, poor thermal and electrical conduction	Argon, Ar; methane, CH_4; sucrose, $C_{12}H_{22}O_{11}$; Dry Ice, CO_2
Covalent-network	Atoms connected in a network of covalent bonds	Covalent bonds	Very hard, very high melting point, variable thermal and electrical conduction	Diamond, C; quartz, SiO_2
Ionic	Positive and negative ions	Electrostatic attractions	Hard and brittle, high melting point, poor thermal and electrical conduction	Typical salts—for example, NaCl, $Ca(NO_3)_2$
Metallic	Atoms	Metallic bonds	Soft to very hard, low to very high melting point, excellent thermal and electrical conduction, malleable and ductile	All metallic elements—for example, Cu, Fe, Al, Pt

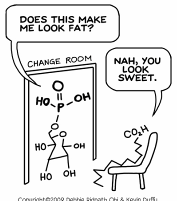
Other Types of Crystalline Solids

End of Chapter 10

See:

- Chapter Ten Study Guide
- Chapter Ten 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:

Intermolecular (IM) Forces: *know* **when** *they apply,* **strength** *(strongest to weakest)*:

- **metallic/ion-ion**
- **ion-dipole**
- **dipole-dipole (***with* Hydrogen bonding *for* **O, F** *and* **N** *to* **H)**
- **dipole-induced dipole**
- **induced dipole-induced dipole (ID-ID)**

States of Matter: **solids, liquids, gases, phase diagrams, triple point, "normal" boiling and freezing points, the slope of the solid-liquid line in a phase diagram, q = mC∆T and q = "mass*heat", vapor pressure**

Solids: *unit cell type:*

- **simple cubic (SC)**
- **body centered cubic (BCC)**
- **face centered cubic (FCC)**

sc: 1 atom, $d_0 = 2r$
bcc: 2 atoms, $d_0 = 4r/\sqrt{3}$
fcc: 4 atoms, $d_0 = 4r/\sqrt{2}$
mole = 6.022×10^{23}

r ⇔ **d_0** ⇔ **Volume** < *density* > **mass** < *molar mass* > **mols** < *avogadro's number* > **atoms/molecules**

End of Chapter Problems: *Test Yourself*

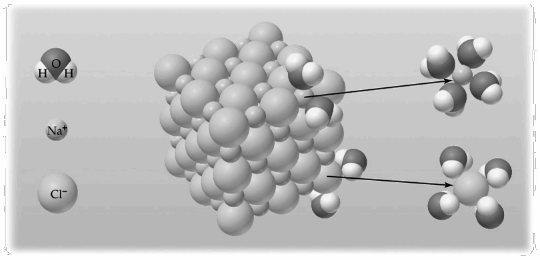
1. What type of intermolecular force must be overcome in converting each of the following from a liquid to a gas? **liquid O_2, H_2O, CH_3I, CH_3CH_2OH**
2. Ethanol, CH_3CH_2OH, has a vapor pressure of 59 mm Hg at 25 °C. What quantity of heat energy is required to evaporate 125 mL of the alcohol at 25 °C? The enthalpy of vaporization of the alcohol at 25 °C is 42.32 kJ/mol. The density of the liquid is 0.7849 g/mL.
3. Liquid ammonia, NH_3(l), was once used in home refrigerators as the heat transfer fluid. The specific heat of the liquid is 4.7 J/g • K and that of the vapor is 2.2 J/g • K. The enthalpy of vaporization is 23.33 kJ/mol at the boiling point. If you heat 12 kg of liquid ammonia from -50.0 °C to its boiling point of -33.3 °C, allow it to evaporate, and then continue warming to 0.0 °C, how much heat energy must you supply?
4. Aluminum has a face centered cubic crystal lattice and a density of 2.699 g/cm^3. What is the radius of an aluminum atom?
5. Iron has a body centered cubic unit cell and a radius of 126 pm. Find the density of iron.

End of Chapter Problems: *Answers*

1. **liquid O_2**: ID-ID, **H_2O**: Hydrogen bonding, **CH_3I**: Dipole-dipole, **CH_3CH_2OH**: Hydrogen bonding
2. 90.1 kJ
3. The total heat required is the sum of the heat required to (1) heat the liquid from –50.0 ºC to its boiling point, (2) vaporize the gas, and (3) heat the vapor to 0.0 ºC. Answer (1) = 940 kJ, Answer (2) = 16000 kJ, Answer (3) = 880 kJ, and total energy = 18000 kJ
4. 143.2 pm
5. 7.8740 g/cm^3

Solutions and Their Behavior

Chapter 11



Chemistry 222
Professor Michael Russell

Last update: 4/29/24

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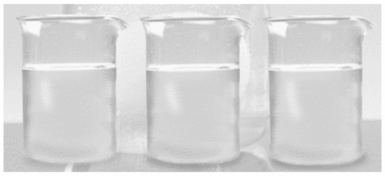
Solution Behavior



Why does a raw egg swell or shrink when placed in different solutions?

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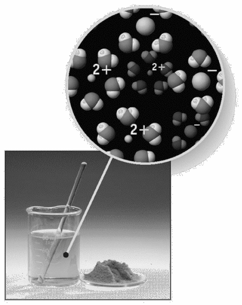
Some Definitions

- **A solution is a HOMOGENEOUS mixture of 2 or more substances in a single phase.**
- **Solutions, Colloids and Suspensions are all examples of homogenous mixtures**
- **Solutions most important to chemists**



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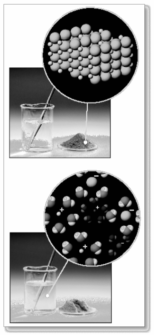
Some Definitions

- **A SOLUTION has at least two constituents**
- **One constituent is usually regarded as the SOLVENT and the others as SOLUTES**
- **So a SOLUTION = SOLUTE + SOLVENT**



State of Solution	State of Solvent	State of Solute	Example
Gas	Gas	Gas	Air
Liquid	Liquid	Gas	Oxygen in water
Liquid	Liquid	Liquid	Alcohol in water
Liquid	Liquid	Solid	Salt in water
Solid	Solid	Gas	Hydrogen in palladium
Solid	Solid	Liquid	Mercury in silver
Solid	Solid	Solid	Silver in gold

MAR

Definitions

Solutions can be classified as unsaturated or saturated.

A saturated solution contains the maximum quantity of solute that dissolves at that temperature.

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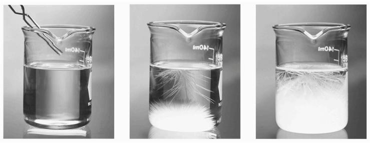
Definitions

SUPERSATURATED SOLUTIONS contain more than is possible and are unstable.

Unstable solutions; crystallization can usually be stimulated by adding a "seed crystal" or scratching the side of the flask.



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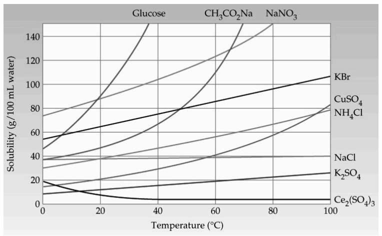
Solution Definitions

Miscible: Two (or more) liquids that are completely soluble in each other

Solubility: A measure of how much solute will dissolve in a solvent at a specific temperature



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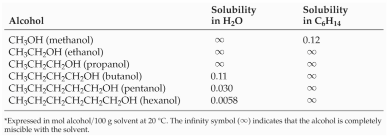
Factors Affecting Solubility

Chemists use the axiom "like dissolves like"

- **The closer the intermolecular forces of two phases of matter, the more likely they will dissolve to make a solution**
- **Polar substances tend to dissolve in polar solvents.**
- **Nonpolar substances tend to dissolve in nonpolar solvents.**



Alcohol	Solubility in H_2O	Solubility in C_6H_{14}
CH_3OH (methanol)	∞	0.12
CH_3CH_2OH (ethanol)	∞	∞
$CH_3CH_2CH_2OH$ (propanol)	∞	∞
$CH_3CH_2CH_2CH_2OH$ (butanol)	0.11	∞
$CH_3CH_2CH_2CH_2CH_2OH$ (pentanol)	0.030	∞
$CH_3CH_2CH_2CH_2CH_2CH_2OH$ (hexanol)	0.0058	∞

*Expressed in mol alcohol/100 g solvent at 20 °C. The infinity symbol (∞) indicates that the alcohol is completely miscible with the solvent.

MAR

Factors Affecting Solubility

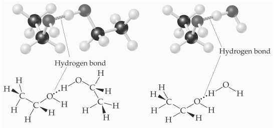
"Like dissolves like"

Example: **water (polar, H-bonds) dissolves in acetone (polar)**

Example: **benzene (ID-ID) dissolves in toluene (ID-ID)**

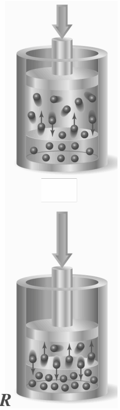
Example: **benzene (ID-ID) does not dissolve very well in water (polar, H-bonds)**

See *Solubility Handout*



We will discuss solubility more in CH 223

MAR



The solubility of liquids and solids does not change appreciably with pressure.

The solubility of a gas in a liquid is directly proportional to its pressure.

Henry's Law: $S_g = kP_g$

- **S_g is the solubility of the gas (M)**
- **k is the Henry's Law constant for that gas in that solvent, and**
- **P_g is the partial pressure of the gas above the liquid.**

Henry's Law

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Dissolving Gases & Henry's Law

Gas solubility (M) = S_g = $k \cdot P_g$

k for O_2 = 1.66×10^{-6} M/mm Hg

When P_g drops, solubility drops.

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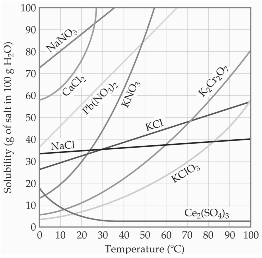
Temperature and Solubility



Generally, the solubility of solid solutes in liquid solvents *increases* with increasing temperature.

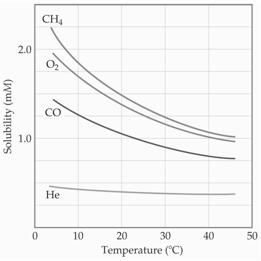
MAR

Temperature and Solubility

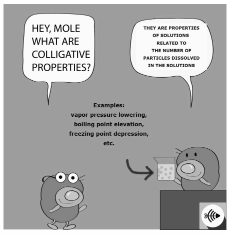
The opposite is true for gases.

Carbonated soft drinks are more “bubbly” if stored in the refrigerator.

Warm lakes have less O_2 dissolved in them than cool lakes.



Colligative Properties



On adding a **solute to a solvent,** the ***solvent*** properties are modified.

- **Vapor pressure** **decreases**
- **Melting point** **decreases**
- **Boiling point** **increases**
- **Osmosis is possible (osmotic pressure)**

These changes are called **COLLIGATIVE PROPERTIES.**

They depend only on the **NUMBER of solute particles relative to solvent particles,** not on the **KIND** of solute particles. We need **new concentration units**! (more than molarity, M)

Concentration Units

- **MOLE FRACTION, X**

For a mixture of A, B, and C

$$X_A = \text{ mol fraction A} = \frac{\text{mol A}}{\text{mol A} + \text{mol B} + \text{mol C}}$$

- **MOLALITY, m**

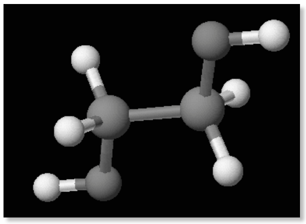
$$\text{m of solute} = \frac{\text{mol solute}}{\text{kilograms solvent}}$$

- **WEIGHT % = grams solute per total g in solution**

see Concentration Units Handout

Calculating Concentrations

Dissolve 62.1 g (1.00 mol) of ethylene glycol in 250. g of H_2O. Calculate mol fraction, molality, and weight % of glycol.



Calculating Concentrations

Dissolve 62.1 g (1.00 mol) of ethylene glycol in 250. g of H_2O. Calculate X, m, and % of glycol.

250. g H_2O = 13.9 mol

$$X_{\text{glycol}} = \frac{1.00 \text{ mol glycol}}{1.00 \text{ mol glycol} + 13.9 \text{ mol } H_2O}$$

X_{glycol} = 0.0672

X_{water} = 1 - 0.0672 = 0.9328

Calculating Concentrations

Dissolve 62.1 g (1.00 mol) of ethylene glycol in 250. g of H_2O. Calculate X, m, and % of glycol.

Calculate molality

$$\text{molality} = \frac{1.00 \text{ mol glycol}}{0.250 \text{ kg } H_2O} = 4.00 \text{ m}$$

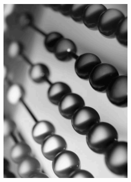
Calculate weight %

$$\%\text{glycol} = \frac{62.1 \text{ g}}{62.1 \text{ g} + 250. \text{ g}} \times 100\% = 19.9\%$$

%water = 100% - 19.9% = 80.1%

see *Concentration Units Handout*

M *vs* m, χ and wt%



Converting between molality (m), mole fraction (χ) and weight percent (wt%) relatively straightforward

Molarity (M) = mol solute per *Liter of solution*

Converting from m, χ or wt% to molarity (M) requires *solution density* in units of g/mL *or* g/cm^3

Recall: **solution = solute + solvent**

MAR

M *vs* m, χ and wt%

***Example:* Convert a 10.7 m aqueous NaOH solution to χ, wt% and M if the solution density = 1.33 g/cm^3.**

Solution:

***Assume* 1 kg solvent and 10.7 mol solute**

1 kg water = 1000 g H_2O = 55.5 mol H_2O

10.7 mol NaOH * 40.0 g mol^{-1} = 428 g NaOH

MAR

M *vs* m, χ and wt%

***Example:* Convert a 10.7 m aqueous NaOH solution to χ, wt% and M if the solution density = 1.33 g/cm^3.**

χ_{NaOH} = 10.7 / (10.7 + 55.5) = 0.162

wt%$_{NaOH}$ = 428 / (1000 + 428) * 100%

wt%$_{NaOH}$ = 30.0%

You do not need the density of the solution to calculate wt% and χ from molality!

MAR

M *vs* m, χ and wt%

***Example:* Convert a 10.7 m aqueous NaOH solution to χ, wt% and M if the solution density = 1.33 g/cm^3.**

If we have 1000 g of solvent (H_2O),

then we have 428 g of solute (NaOH) or

1428 g of solution (NaOH + H_2O)

***Volume of solution* = 1428 g * (cm^3 / 1.33 g)**

***Volume of solution* = 1070 mL = 1.07 L**

***Molarity* = 10.7 mol NaOH / 1.07 L solution**

Molarity NaOH = 10.0 M

MAR

see Concentration Units Handout

Parts Per Million (ppm)

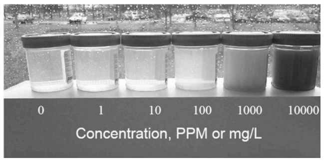
Expresses very dilute concentrations; used by environmental chemists, biologists, geologists, etc.

1.0 ppm = 1.0 g of a substance in **a sample with a total mass of** 1.0 million (10^6) g

Density of water about 1.0 g/mL, units often interconverted

Also parts per thousand (ppt), parts per billion (ppb), etc.

Different than the lab parts per thousand, which is a measure of *precision*



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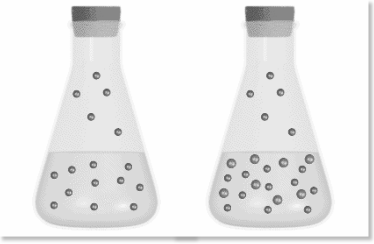
Understanding Colligative Properties

To understand colligative properties, study the LIQUID-VAPOR EQUILIBRIUM for a solution.



The right animation assumes a non-volatile solute

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Understanding Colligative Properties

Vapor Pressure of H_2O over a solution depends on the number of H_2O molecules per solute molecule.

$P_{solvent}$ *proportional to* $X_{solvent}$

or

$P_{solvent} = X_{solvent} \bullet P^o_{solvent}$

Vapor Pressure of *solvent* over *solution* = (Mol frac solvent)•(VP pure solvent)

RAOULT'S LAW: The vapor pressure of solvent over a solution is always LOWER than the pure solvent!

MAR

Raoult's Law

Non-volatile ethylene glycol (62.1 g) is placed in 250. g of water. What is the vapor pressure of water in the solution at 30 °C? (The vapor pressure of pure H_2O is 31.8 mm Hg)

Solution

$X_{glycol} = 0.0672$ and so $X_{water} = ?$

Because $X_{glycol} + X_{water} = 1$

$X_{water} = 1.000 - 0.0672 = 0.9328$

$P_{water} = X_{water} \bullet P^o_{water} = (0.9328)(31.8 \text{ mm Hg})$

$P_{water} = 29.7$ mm Hg

MAR

Raoult's Law

Previous examples assumed a non-volatile solute to calculate vapor pressure. If solute is volatile, then:

$P_{total} = X_A \bullet P^o_A + X_B \bullet P^o_B + X_C \bullet P^o_C + \ldots$

***Example:* At 25 ˚C, heptane (P^o = 31 torr) and octane (P^o = 11 torr) are mixed such that $X_{heptane} = 0.15$. Calculate the total vapor pressure of the system.**

***Solution:* This is a two component system.**

If $X_{heptane} = 0.15$, *then*

$X_{octane} = 1 - 0.15 = 0.85$

MAR

Raoult's Law

***Problem:* At 25 ˚C, heptane (P^o = 31 torr) and octane (P^o = 11 torr) are mixed such that $X_{heptane} = 0.15$ and $X_{octane} = 0.85$. Calculate the total vapor pressure of the system.**

***Solution:* In a two component system,**

$P_{total} = X_{hept} \bullet P^o_{hept} + X_{oct} \bullet P^o_{oct}$

$P_{total} = 0.15 \bullet 31 \text{ torr} + 0.85 \bullet 11 \text{ torr}$

$P_{total} = 4.7 + 9.4 = 14.1$ torr

The total vapor pressure in the mixture is 14.1 torr

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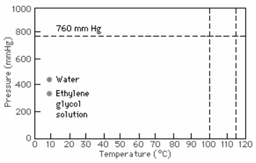
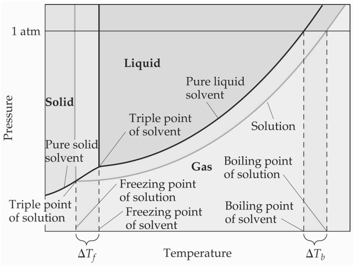
The boiling point of a solution is higher than that of the pure solvent.





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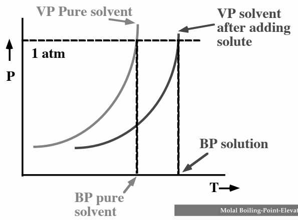
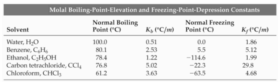
Elevation of Boiling Point

Elevation in BP = $\Delta T_{BP} = K_{BP} \bullet m$

(where K_{BP} is characteristic of solvent)





Molal Boiling-Point-Elevation and Freezing-Point-Depression Constants

Solvent	Normal Boiling Point (°C)	K_b (°C/m)	Normal Freezing Point (°C)	K_f (°C/m)
Water, H_2O	100.0	0.51	0.0	1.86
Benzene, C_6H_6	80.1	2.53	5.5	5.12
Ethanol, C_2H_5OH	78.4	1.22	−114.6	1.99
Carbon tetrachloride, CCl_4	76.8	5.02	−22.3	29.8
Chloroform, $CHCl_3$	61.2	3.63	−63.5	4.68

MAR

Change in Boiling Point

Dissolve 62.1 g of glycol (1.00 mol) in 250. g of water. What is the BP of the solution?

K_{BP} = 0.512 °C/molal for water (see Textbook)

Solution

1. **Calculate solution molality = 4.00 m**
2. **$\Delta T_{BP} = K_{BP} \cdot m$**

   ΔT_{BP} = 0.512 °C/molal (4.00 molal)

   ΔT_{BP} = 2.05 °C

   BP = 100.00 + 2.05 = 102.05 °C

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Change in Freezing Point

Pure water **Ethylene glycol/water solution**

The freezing point of a solution is LOWER than that of the pure solvent.

Notice how the animations incorrectly use "dense" ice!

MAR

The freezing point of a solution is lower than that of the pure solvent.

FP depression = $\Delta T_{FP} = K_{FP} \cdot m$

■ Molal Boiling-Point-Elevation and Freezing-Point-Depression Constants

Solvent	Normal Boiling Point (°C)	K_b (°C/m)	Normal Freezing Point (°C)	K_f (°C/m)
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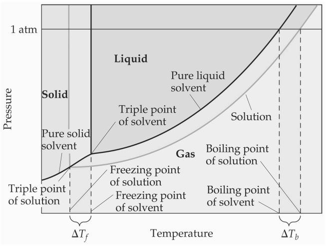

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Freezing Point Depression

Calculate the FP of a 4.00 molal glycol/water solution.

K_{FP} = 1.86 °C/molal (from *Textbook*)

Solution

$\Delta T_{FP} = K_{FP} \cdot m$

= (1.86 °C/molal)(4.00 m)

ΔT_{FP} = 7.44 °C

So the FP = **0 ˚C - 7.44 ˚C = -7.44 ˚C**

Many textbooks use *negative* K_{FP} values
If using *negative* K_{FP}, ΔT_{FP} is also negative
Concentration (molality) *must* be positive!

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Freezing Point Depression of Electrolytes

How much NaCl must be dissolved in 4.00 kg of water to lower FP to -10.00 °C?

Solution

Calculate required molality

$\Delta T_{FP} = K_{FP} \cdot m$

(0 - -10.00 °C) = (1.86 °C/molal) • m

m = 5.38 molal

Use -10 °C for ΔT if -1.86 °C/m

MAR

Freezing Point Depression of Electrolytes

How much NaCl must be dissolved in 4.00 kg of water to lower FP to -10.00 °C?

Solution

Concentration required = 5.38 molal

This means we need 5.38 mol of dissolved particles per kg of solvent.

Recognize that *m* represents the total concentration of *all* dissolved particles.

Recall that 1 mol NaCl(aq) --> 1 mol Na^+(aq) + 1 mol Cl^-(aq)... or: 2 mol of particles per 1 mol of ionic solute!

MAR

Freezing Point Depression of Electrolytes

How much NaCl must be dissolved in 4.00 kg of water to lower FP to -10.00 °C?

Solution

Concentration required = 5.38 molal

We need 5.38 mol of dissolved particles per kg of solvent.

$NaCl(aq) \longrightarrow Na^+(aq) + Cl^-(aq)$

2 particles per mol NaCl!

To get 5.38 mol/kg of particles we need

5.38 mol / 2 = 2.69 mol NaCl / kg

2.69 mol NaCl / kg ---> 157 g NaCl / kg

(157 g NaCl / kg)•(4.00 kg) = 629 g NaCl

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Boiling Point Elevation and Freezing Point Depression

$$\Delta T = k \bullet m \bullet i$$

A generally useful equation

i = van't Hoff factor = number of particles produced per formula unit.

Compound	*Theoretical* Value of i
glycol	1
NaCl	2
$CaCl_2$	3

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We will use the theoretical value of i in CH 222

Colligative Properties and Molar Mass

Can use colligative properties to find molar mass of solute

Molar mass = grams solute / moles solute

Example: **Find the molar mass of azulene if 0.640 g are dissolved in 99.0 g of benzene (k_b = 2.53 ˚C/m), and the observed boiling point is 80.230 ˚C. (normal bp = 80.100 ˚C)**

MM = (grams solute * k)/(ΔT*kg solvent)
useful for both bp and fp calculations

Solution:

$\Delta T = (80.230 - 80.100) = 0.130$ ˚C

$m = \Delta T/k_b = 0.130$ ˚C/2.53 ˚C/m $= 5.14 \times 10^{-2}$ mol / kg

0.0990 kg * 5.14×10^{-2} mol / kg solvent = 5.05×10^{-3} mol

Molar mass = grams solute / moles solute

Molar mass = 0.640 g / 5.05×10^{-3} moles = 127 g mol^{-1}

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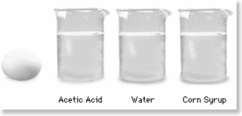
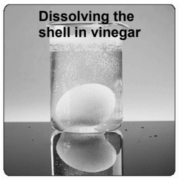
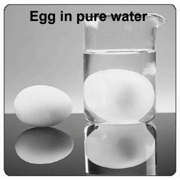
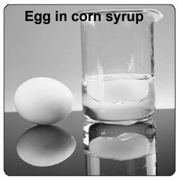
Osmosis









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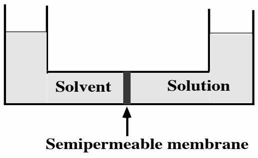
Osmosis



The semipermeable membrane allows only the movement of solvent molecules.

Solvent molecules move from pure solvent to solution in an attempt to make both have the same concentration of solute.

Driving force is entropy (CH 223)

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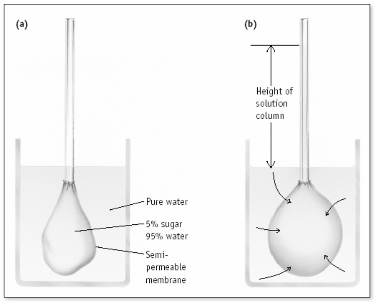
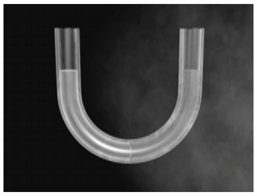
Process of Osmosis





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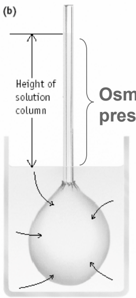
Osmotic Pressure, Π



Equilibrium is reached when pressure - the *OSMOTIC PRESSURE, Π* - produced by extra solution counterbalances pressure of solvent molecules moving through the membrane.

$$\Pi = icRT$$

c = conc. in mol/L

i = the van't Hoff factor

R = 0.082057 (L atm)/(mol K)

T = Kelvin Temp.

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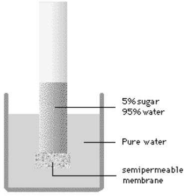

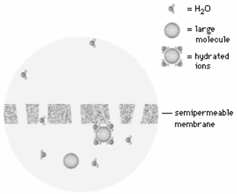

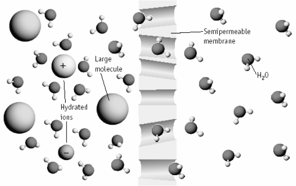

Osmosis

MAR

Osmosis
Calculating a Molar Mass

Dissolve 35.0 g of hemoglobin in enough water to make 1.00 L of solution. π measured to be 10.0 mm Hg at 25 ˚C. Calc. molar mass of hemoglobin.

Solution

(a) Calc. π in atmospheres

$$\pi = 10.0 \text{ mm Hg} \cdot (1 \text{ atm} / 760 \text{ mm Hg})$$

$$\pi = 0.0132 \text{ atm}$$

MAR

Osmosis
Calculating a Molar Mass

Dissolve 35.0 g of hemoglobin in enough water to make 1.00 L of solution. π measured to be 10.0 mm Hg at 25 ˚C. Calc. molar mass of hemoglobin.

Solution

(b) Calc. concentration from π = cRT

$$\text{Conc} = \frac{0.0132 \text{ atm}}{(0.082057 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(298\text{K})}$$

Conc = 5.40 x 10^{-4} mol/L (* 1 L)

(c) Calc. molar mass

Molar mass = 35.0 g / 5.40 x 10^{-4} mol

Molar mass = 64,800 g/mol

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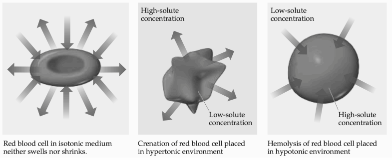
Osmosis

Osmosis of solvent from one solution to another can continue until the solutions are ISOTONIC - they have the same concentration.

Osmotic pressure important in living systems

MAR

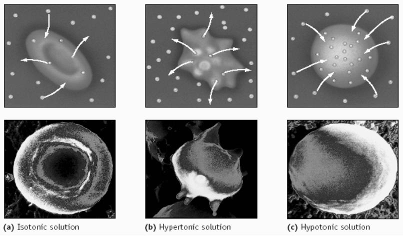
Osmosis & Living Cells



Red blood cell in isotonic medium neither swells nor shrinks.

Crenation of red blood cell placed in hypertonic environment

Hemolysis of red blood cell placed in hypotonic environment

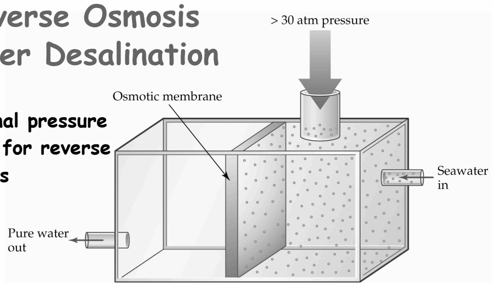


(a) Isotonic solution
(b) Hypertonic solution
(c) Hypotonic solution

MAR

Reverse Osmosis Water Desalination

External pressure allows for reverse osmosis



Water desalination plant

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Other Homogeneous Mixtures

Homogeneous mixtures: more than just solutions!

- **Solutions - most important, particles in the range 0.2 - 2 nm in diameter (salt water, acids, etc.)**
- **Colloids - a "solution" with particles in the range 2 - 1000 nm (milk, fog, etc.)**
- **Suspensions - contain particles greater than 1000 nm in diameter (blood, paint, aerosol sprays, etc.)**

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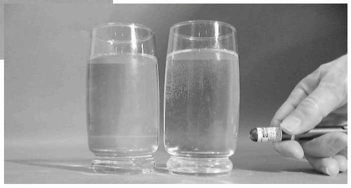
The Tyndall Effect

Colloids (and suspensions) have particles too small to settle out of solution - hence, the Tyndall effect



***Right*: Pure tap water**
***Left:* 5 ppm colloidal silver**

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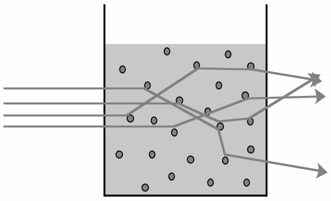
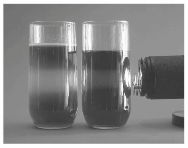
The Tyndall Effect





MAR

Colloids

Colloids have a dispersed phase ("solute") and a dispersing medium ("solvent"). Many types of colloids:

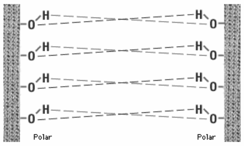
Type	D Medium	D Phase	Examples
aerosol	gas	liquid	fog, clouds
aerosol	gas	solid	smoke, viruses
foam	liquid	gas	shaving cream
emulsion	liquid	liquid	milk, mayonnaise
sol	liquid	solid	mud, paint, ink
foam	solid	gas	foam rubber, sponge
gel	solid	liquid	jelly, cheese, *yogurt*
solid sol	solid	solid	steel, bronze, pearls

MAR

Surfactant Emulsifying Agents

Surfactants (or Surfactant Emulsifying Agents) promote the formation of colloids by coating dispersed phase, preventing the formation of large particles.

Surfactants used for cleaning called detergents.

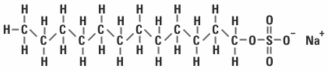


MAR

Surfactant Emulsifying Agents

Soaps and detergents: surfactants with both a hydrophobic (nonpolar) and a hydrophilic (polar) end. Can remove nonpolar oils *and* polar compounds.

Soaps often grouped in micelles with like-polarities grouped together.



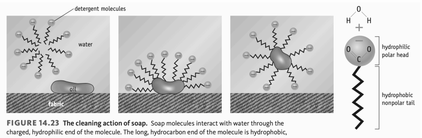
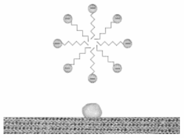
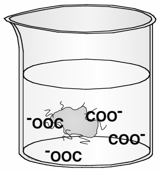


FIGURE 14.23 The cleaning action of soap. Soap molecules interact with water through the charged, hydrophilic end of the molecule. The long, hydrocarbon end of the molecule is hydrophobic, but it can bind through dispersion forces with hydrocarbons and other nonpolar substances.



Oil and Water

Oil and water do not mix due to differences in polarity; water *excludes* oil due to stronger intermolecular forces
Soap removes oil readily using nonpolar end of micelle



End of Chapter 11

See:
- *Chapter Eleven Study Guide*
- *Chapter Eleven Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*

Important Equations, Constants, and Handouts from this Chapter:

- **solution = solute + solvent**
- **see Concentration Units Handout**

Solution Concentrations:

$$Molarity\ (M) = \frac{mol\ Solute}{L\ Solution}$$

$$molality(m)\ = \frac{mol\ Solute}{kg\ Solvent}$$

$$weight\ \% = \frac{mass\ Solute}{Total\ mass}$$

$$mole\ fraction\ (\chi) = \frac{mol\ A}{total\ mol}$$

$$ppm = \frac{1.0\ g\ of\ substance}{1.0\ million\ g\ sample}$$

Henry's Law:

$$S_g = k \bullet P_g$$

Raoult's Law / Vapor Pressure Depression:

$$P_{solvent} = \chi_{solvent} \bullet P^o_{solvent}$$

Boiling Point Elevation / Freezing Point Depression:

$$\Delta T_{BP/FP} = K_{BP/FP} \bullet \left(\frac{mol\ Solute}{kg\ Solvent}\right) \bullet i$$

Osmosis:

$$\pi = i\left(\frac{mol\ Solute}{L\ Solvent}\right)RT$$

R = 0.082057 L atm mol⁻¹ K⁻¹
i = van't Hoff factor

End of Chapter Problems: *Test Yourself*

1. Fill in the blanks in the table. All solutions are aqueous.

Compound	Molality	Weight Percent	Mole Fraction
NaI	0.15	_____	_____
C_2H_5OH	_____	5	_____
$C_{12}H_{22}O_{11}$	_____	_____	0.0027

2. Hydrochloric acid is sold as a concentrated aqueous solution. If the molarity of commercial HCl is 12.0 and its density is 1.18 g/cm3, calculate the molality and the weight percent of HCl in the solution.
3. A soda can has an aqueous CO_2 concentration of 0.0506 M. What is the pressure of CO_2 gas in the can? ($k_H = 4.48 \times 10^{-5}$ M/mm Hg)
4. Pure nonvolatile iodine (105 g) is dissolved in 325 g of CCl_4 at 65 °C. Given that the vapor pressure of CCl_4 at this temperature is 531 mm Hg, what is the vapor pressure of the solution at 65 °C?
5. What is the boiling point of a solution composed of 15.0 g of $CHCl_3$ (which boils at 61.70 °C) and 0.515 g of the nonvolatile solute acenaphthene, $C_{12}H_{10}$?
6. An aqueous solution containing 1.00 g of bovine insulin (a protein, not ionized) per liter has an osmotic pressure of 3.1 mm Hg at 25 °C. Calculate the molar mass of bovine insulin.

End of Chapter Problems: *Answers*

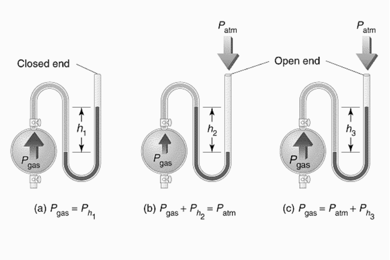
1. Completed table:

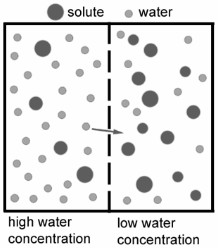
Compound	Molality	Weight percent	Mole fraction
NaI	0.15	2.20	.0027
C_2H_5OH	1.1	5.00	.020
$C_{12}H_{22}O_{11}$	0.15	4.90	.0027

2. 16.2 m, 37.1%
3. 1130 mm Hg
4. 444 mm Hg
5. 62.51 °C
6. 6.0×10^3 g/mol

Chemistry 222 Exam II Review

Chapters 9, 10 and 11





Last update: 4/29/24

Chemistry 222
Professor Michael Russell

MAR

A sample of gas has a volume of 222 mL at 695 mm Hg and 0 °C. What would be the volume of this same sample of gas if it were measured at 333 mm Hg and 0 °C?

A. 894 mL
B. 463 mL
C. 657 mL
D. 359 mL
E. -155 mL

MAR

Gas density: Which has the greatest density at 25 °C and 1.00 atm pressure?

A. O_2
B. N_2
C. H_2
D. CO_2
E. Xe

MAR

Diborane reacts with O_2 to give boric oxide and water vapor:

$$B_2H_6(g) + 3\ O_2(g) \rightarrow B_2O_3(s) + 3\ H_2O(g)$$

If 1.5 L of B_2H_6 is mixed with O_2, what volume of O_2 is required for reaction? Assume constant T and P.

A. 1.5 L
B. 4.5 L
C. 0.50 L
D. 6.0 L
E. 0.0 L

MAR

Diborane, B_2H_6, burns in air according to the equation:

$$B_2H_6(g) + 3\ O_2(g) \rightarrow B_2O_3(s) + 3\ H_2O(g)$$

There are three gases involved in the reaction above. Place them in order of increasing velocity.

A. $B_2H_6 < O_2 < H_2O$
B. $O_2 < B_2H_6 < H_2O$
C. $H_2O < B_2H_6 < O_2$
D. $O_2 < H_2O < B_2H_6$

MAR

Equal masses of helium and neon are placed in separate containers of equal V at the same T. Compare the pressures of the gases.

A. P(He) > P(Ne)
B. P(Ne) > P(He)
C. P(He) = P(Ne)
D. Too much pressure!

MAR

Using intermolecular forces, the predicted order of decreasing boiling points for the following substances is

A. $CH_3OH > CH_4 > H_2$
B. $CH_3OH > H_2 > CH_4$
C. $CH_4 > CH_3OH > H_2$
D. $CH_4 > H_2 > CH_3OH$
E. $H_2 > CH_4 > CH_3OH$

MAR

Rank the following molecules in order of increasing intermolecular forces:

SO_2, NaCl, CH_3OH, He

A. NaCl < He < SO_2 < CH_3OH
B. He < CH_3OH < SO_2 < NaCl
C. He < SO_2 < CH_3OH < NaCl
D. He < SO_2 < NaCl < CH_3OH

MAR

When KCl dissolves in water, what types of intermolecular bonds are formed?

A. ion-ion
B. ion-dipole
C. hydrogen bonds
D. ion-ion forces and H-bonds
E. matrimony bonds

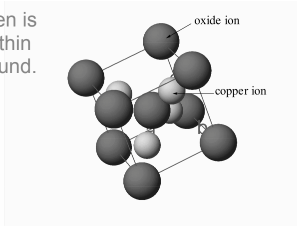
MAR

Which of the following should have the highest ΔH_{vap}?

A. F_2
B. CH_3OH
C. H_2O
D. NH_3
E. all have the same ΔH_{vap} value

MAR

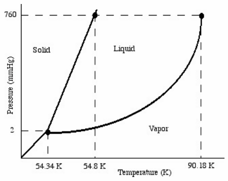
The unit cell for a compound of copper and oxygen is illustrated here. All the grey copper atoms are within the unit cell. Determine the formula of the compound.



A. CuO
B. Cu_2O
C. CuO_2
D. Cu_2O_3
E. Cu_4O_9

MAR

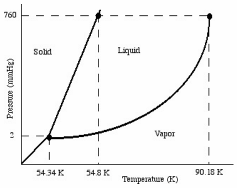
Below is a phase diagram for O_2. What is the name of the point at P = 2 mm Hg and T = 54.34 K?



A. normal freezing point
B. triple point
C. normal boiling point
D. critical point
E. freak out point

MAR

At right is a phase diagram for O_2. Which statement is correct regarding the densities of solid and liquid O_2 at the same temperature?

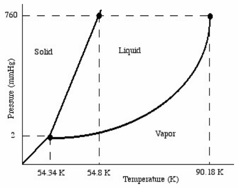


A. The density of liquid O_2 is 1.426 g/cm^3 whereas the density of solid O_2 is 1.149 g/cm^3.
B. The density of solid O_2 is 1.426 g/cm^3 whereas the density of liquid O_2 is 1.149 g/cm^3.
C. The densities of solid and liquid O_2 are the same.

MAR

Below is a phase diagram for O_2. In what phase is oxygen found at a pressure of 1 mmHg and a temperature of 55 K?

A. solid
B. liquid
C. vapor
D. plasma



MAR

Calculate the energy required to convert 1.00 L of liquid ethanol at 25.0 °C to a gas at 78.3 °C.

Constants for ethanol:

- density = 0.7849 g/mL
- molar mass = 46.08 g/mol
- boiling point = 78.3 °C
- heat capacity = 2.44 J/g·K
- heat of vaporization = 38.56 kJ/mol

A. -329 kJ
B. 329 kJ
C. -759 kJ
D. 759 kJ
E. 0 KJ

MAR

You dissolve 92.0 grams of CH_3CH_2OH, ethanol, in 270 g of water. What is the mole fraction of ethanol in the solution?

A. 2.0
B. 0.12
C. 0.067
D. 0.0083
E. 0

MAR

You dissolve 92.0 grams of CH_3CH_2OH, ethanol, in 270. g of water. What is the weight percent of ethanol in the solution?

A. 34.1 %
B. 17.0 %
C. 25.4 %
D. 74.8 %
E. 100. %

MAR

You dissolve 92.0 grams of CH_3CH_2OH, ethanol, in 270. g of water. What is the molality of ethanol in the solution?

A. 341 *m*
B. 170. *m*
C. 0.341 *m*
D. 7.41 *m*
E. 18.0 *m*

MAR

You dissolve 92.0 grams of CH_3CH_2OH, ethanol, in 270. g of water, and the density of the solution is 0.9780 g/mL. What is the molarity of ethanol in the solution?

A. 0.370 M
B. 2.00 M
C. 7.42 M
D. 3.05 M
E. 5.41 M

MAR

An aqueous solution of ethanol (CH_3CH_2OH) has a solution density of 0.9163 g/mL and a concentration of 0.801 M. What is the mole fraction of ethanol in the solution?

A. 0.911
B. 1.1111
C. 0.9839
D. 0.0161
E. 4.03

MAR

The Henry's Law constant for N_2 in water at 53 °C is 8.4×10^{-7} M/mm Hg, and the vapor pressure of water at 53 °C is 107 torr. Find the equilibrium concentration of N_2 in water if the total pressure equals 1 atm.

A. 5.5×10^{-4} M
B. 9.0×10^{-5} M
C. 6.4×10^{-4} M
D. 3.2×10^{-5} M
E. 42 M

MAR

0.0400 mol of I_2 (10.1 g) is dissolved in 1.96 mol of CCl_4 (300 g) at 65 °C. Given that the vapor pressure of pure CCl_4 is 504 mm Hg at this temperature, what is the vapor pressure of the CCl_4 over this solution?

A. 504 mm Hg
B. 514 mm Hg
C. 494 mm Hg
D. 10.1 mm Hg
E. Too much pressure!

MAR

What mass of ethylene glycol (HOC_2H_4OH, molar mass = 62.0 g/mol) must be added to 125 g of water to raise the boiling point by 1.00 °C? ($k_{bp}(H_2O) = +0.512$ °C/m)

A. 1.95 g
B. 0.244 g
C. 15.1 g
D. 31.0 g
E. 0 g

MAR

Which water-based solution is expected to have the higher boiling point?

A. 0.10 molal NaCl
B. 0.15 molal sugar
C. both the same
D. not enough information

MAR

29.3 g of NaCl (molar mass = 58.5 g/mol) is dissolved in 500. grams water. What is the boiling point of this solution? ($k_{bp}(H_2O)$ = +0.512 °C/m)

A. 100.512 °C
B. 98.976 °C
C. 101.02 °C
D. 1.024 °C
E. 10^4 °K

MAR

Erythritol occurs naturally in algae and fungi. A solution of 2.50 g of erythritol in 50.0 g of water freezes at -0.762 °C. What is the molar mass of the compound? ($k_{fp}(H_2O)$ = -1.86 °C/m)

A. 26.9 g/mol
B. 35.5 g/mol
C. 122 g/mol
D. 224 g/mol
E. 0.0100 g/mol

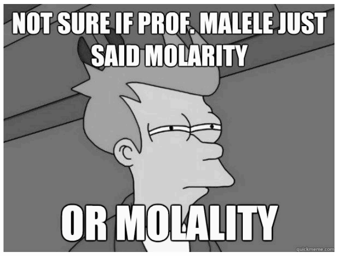
MAR

Insulin (20.0 mg) dissolved in 5.00 mL of water at 300. K provides an osmotic pressure of 12.5 mm Hg. What is the molar mass of insulin?

A. 18.0 g/mol
B. 42.0 g/mol
C. 3980 g/mol
D. 6010 g/mol
E. 12,100 g/mol

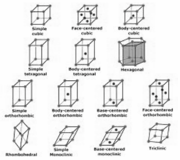
MAR

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)*

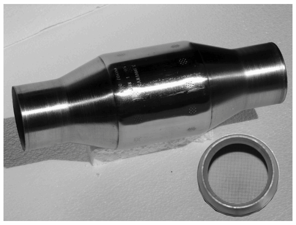
Good luck with your studying!



MAR

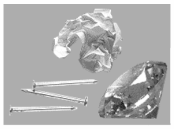
Chemical Kinetics: The Rates of Chemical Reactions

Chapter 12



An automotive catalytic muffler.

Chemistry 222
Professor Michael Russell

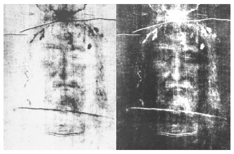


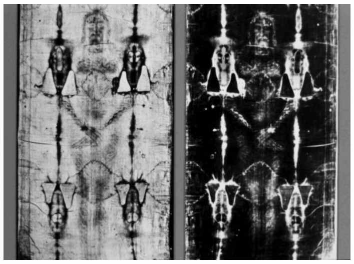
Last update: 4/29/24

MAR

Shroud of Turin

Shroud of Jesus?!?
Fake or Real?
Explored with Kinetics





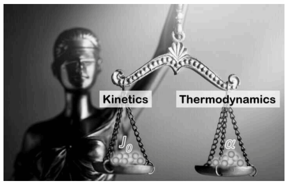
MAR

Chemical Kinetics

We can use thermodynamics to tell if a reaction is product or reactant favored.

But this gives us no info on HOW FAST reaction goes from reactants to products.

KINETICS - the study of REACTION RATES and their relation to the way the reaction proceeds, i.e., its MECHANISM.

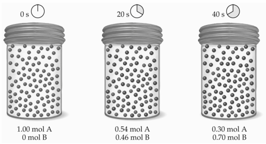


MAR

Reaction Rates

Reaction rate is the change in concentration of a reactant or product with time

Ex: for **A → B**



MAR

Reaction Rates

Reaction rate is the change in concentration of a reactant or product with time

Reactants *disappear* with time (hence, negative sign), and products *appear* with time (hence, positive sign)

Ex: for $N_2O_{4(g)} \rightarrow 2\ NO_{2(g)}$

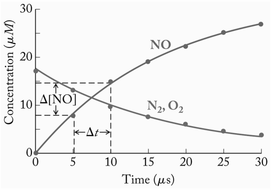
$$rate = -\frac{\Delta[N_2O_4]}{\Delta t} = +\frac{1}{2}\frac{\Delta[NO_2]}{\Delta t}$$

MAR

Rates in chemistry are usually "amount per unit time", i.e. M/s, etc.

Reaction Rates

Ex: $N_{2(g)} + O_{2(g)} \rightarrow 2\ NO_{(g)}$

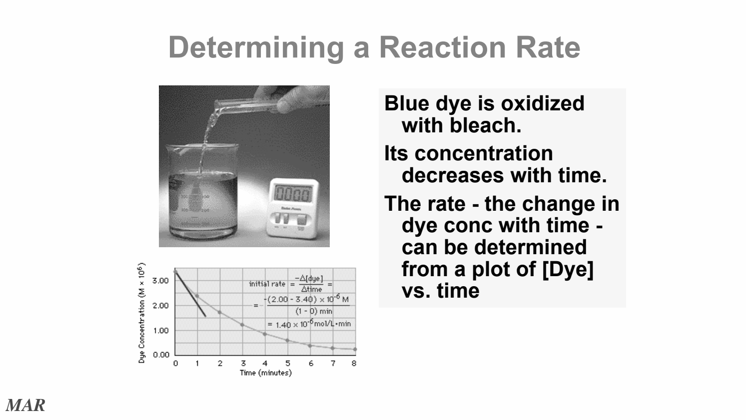


$$\text{Rate of } N_2 \text{ consumption} = -\frac{\Delta[N_2]}{\Delta t} = -\frac{[N_2]_{\text{final}} - [N_2]_{\text{initial}}}{t_{\text{final}} - t_{\text{initial}}}$$

$$\text{Rate of } O_2 \text{ consumption} = -\frac{\Delta[O_2]}{\Delta t} = -\frac{[O_2]_{\text{final}} - [O_2]_{\text{initial}}}{t_{\text{final}} - t_{\text{initial}}}$$

$$\text{Reaction Rate} = -\frac{\Delta[N_2]}{\Delta t} = -\frac{\Delta[O_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[NO]}{\Delta t}$$

MAR



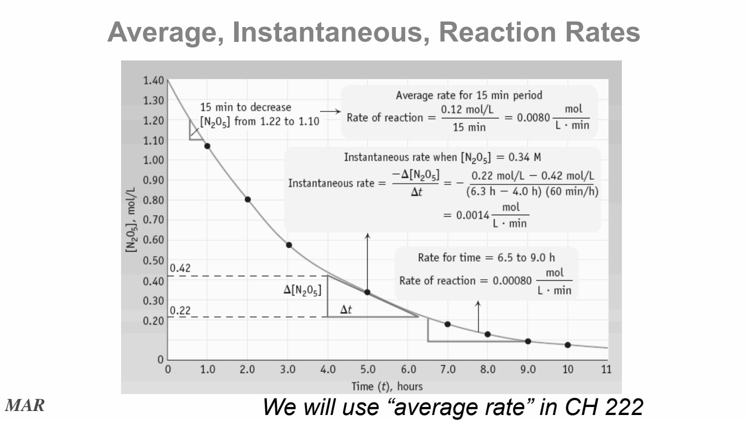
Determining a Reaction Rate

Blue dye is oxidized with bleach.

Its concentration decreases with time.

The rate - the change in dye conc with time - can be determined from a plot of [Dye] vs. time

MAR



Average, Instantaneous, Reaction Rates

Average rate for 15 min period

$$\text{Rate of reaction} = \frac{0.12 \text{ mol/L}}{15 \text{ min}} = 0.0080 \frac{\text{mol}}{\text{L} \cdot \text{min}}$$

Instantaneous rate when $[N_2O_5] = 0.34$ M

$$\text{Instantaneous rate} = \frac{-\Delta[N_2O_5]}{\Delta t} = -\frac{0.22 \text{ mol/L} - 0.42 \text{ mol/L}}{(6.3 \text{ h} - 4.0 \text{ h})(60 \text{ min/h})} = 0.0014 \frac{\text{mol}}{\text{L} \cdot \text{min}}$$

Rate for time = 6.5 to 9.0 h

$$\text{Rate of reaction} = 0.00080 \frac{\text{mol}}{\text{L} \cdot \text{min}}$$

MAR

We will use “average rate” in CH 222

Factors Affecting Rates

Concentration

Greater concentration of reactants means more collisions and faster rates

Temperature

Higher temperatures means particles collide with greater kinetic energy, increasing the rates

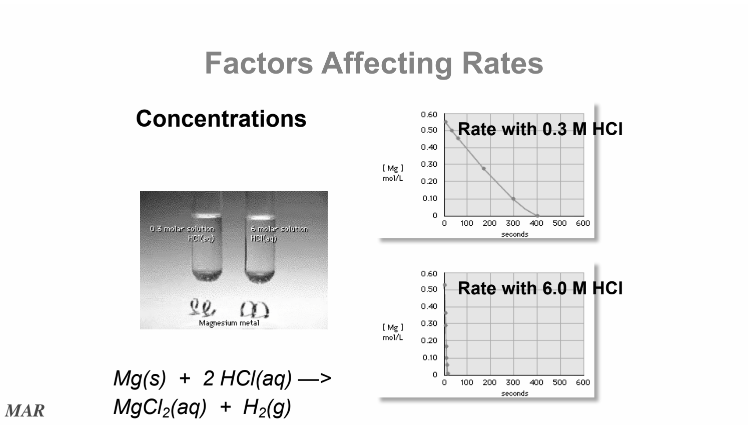
Surface Area

Increased surface area means greater chances for collisions and faster rates

Catalysts

Catalysts speed up reactions without being used up. They lower the activation energy and increase the rates

MAR



Factors Affecting Rates

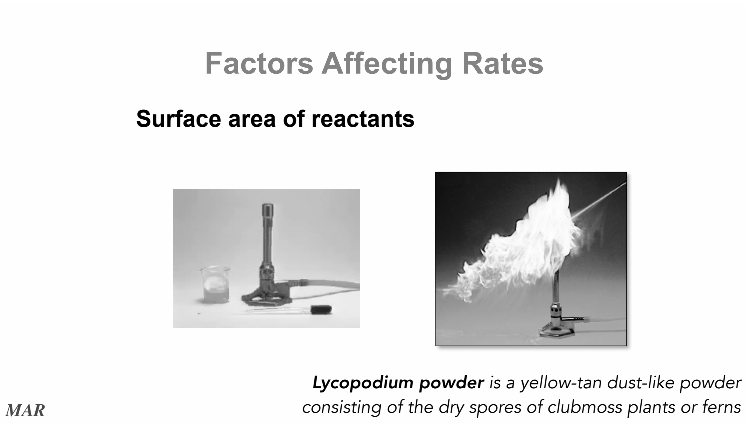
Concentrations

Rate with 0.3 M HCl

Rate with 6.0 M HCl

$Mg(s) + 2 HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$

MAR

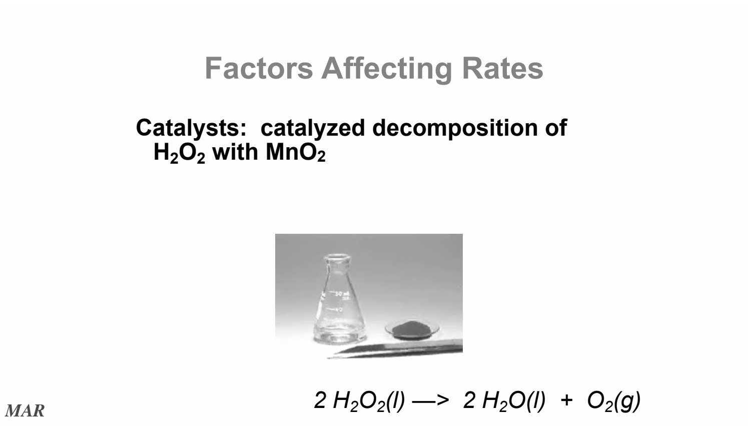


Factors Affecting Rates

Surface area of reactants

Lycopodium powder *is a yellow-tan dust-like powder consisting of the dry spores of clubmoss plants or ferns*

MAR



Factors Affecting Rates

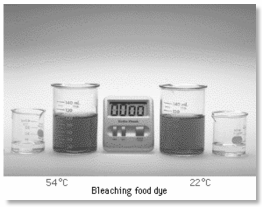
Catalysts: catalyzed decomposition of H_2O_2 with MnO_2

$2 H_2O_2(l) \longrightarrow 2 H_2O(l) + O_2(g)$

MAR

Factors Affecting Rates

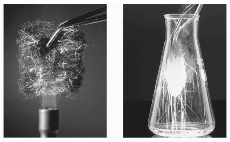
Temperature



MAR

Concentrations and Rates

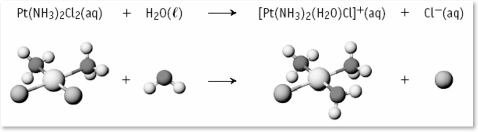
To postulate a reaction mechanism, we study its reaction rate and concentration dependence



MAR

Concentrations and Rates

Take reaction where Cl^- in cisplatin $[Pt(NH_3)_2Cl_2]$ is replaced by H_2O



Rate of change of conc of Pt compd

$$= \frac{\text{Am't of cisplatin reacting (mol/L)}}{\text{elapsed time (t)}}$$

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Concentrations and Rates

Rate of change of conc of Pt compd

$$= \frac{\text{Am't of cisplatin reacting (mol/L)}}{\text{elapsed time (t)}}$$

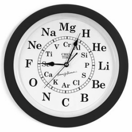
If calculated **rate =** 4.3 x 10^{-6} M s^{-1} *and* **[cisplatin] =** 0.00250 M, *then* ***approximate*** **time** *for cisplatin to react:*

rate = [cisplatin] / time, *and*

time = [cisplatin] / rate

time = 0.00250 M **/** 4.3 x 10^{-6} M s^{-1}

time = 580 s



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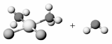

Concentrations and Rates

Rate of change of conc of Pt compd

$$= \frac{\text{Am't of cisplatin reacting (mol/L)}}{\text{elapsed time (t)}}$$

Rate of reaction is *proportional* to $[Pt(NH_3)_2Cl_2]$

We express this as a RATE LAW

Rate of reaction = k $[Pt(NH_3)_2Cl_2]$

where k = rate constant

k is independent of conc. but increases with T

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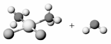

Concentrations, Rates, and Rate Laws

In general, for

a A + b B ---> x X **with a catalyst C**

Rate = k $[A]^m[B]^n[C]^p$

The exponents m, n, and p

- **are the** reaction order
- **can be** 0, 1, **or** 2 *(in CH 222 - CH 223!)*
- **must be determined by experiment!**
- ***overall order* = m + n + p**

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Interpreting Rate Laws

Example: **Rate = k [A]^m**

If m = 1, rxn. is 1st order in A

Rate = k [A]^1

If [A] doubles, then rate goes up by factor of ?

If m = 2, rxn. is 2nd order in A.

Rate = k [A]^2

Doubling [A] increases rate by ?

If m = 0, rxn. is zero order.

Rate = k [A]^0

If [A] doubles, rate ?

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Deriving Rate Laws

Example: **Derive rate law and k for**

$CH_3CHO(g)$ ---> $CH_4(g)$ + $CO(g)$

from experimental data for rate of disappearance of CH_3CHO

Expt.	$[CH_3CHO]$ (mol/L)	Disappear of CH_3CHO (mol/L•sec)
1	0.10	0.020
2	0.20	0.081
3	0.30	0.182
4	0.40	0.318

The rate law: **rate = k$[CH_3CHO]^x$**

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Deriving Rate Laws

Expt.	$[CH_3CHO]$ (mol/L)	Rate of CH_3CHO (mol/L•sec)
1	0.10	0.020
2	0.20	0.081
3	0.30	0.182
4	0.40	0.318

Let's find the order (x) with respect to CH_3CHO

Use two trials where $[CH_3CHO]$ changes - any change in rate is caused by CH_3CHO

Apply the rate law to these two trials:

$$\frac{\text{Rate 2}}{\text{Rate 1}} = \frac{0.081}{0.020} = \frac{k(0.20)^x}{k(0.10)^x}$$

$$4.1 = 2.0^x$$

$$2 = x$$

Could also:

$\log 4.1 = \log 2.0^x = x \log 2.0$

$x = (\log 4.1)/(\log 2.0)$

$x = 2$

So the order of reaction for CH_3CHO is "2". We say the reaction is second order with respect to CH_3CHO

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Deriving Rate Laws

Rate of rxn = k $[CH_3CHO]^2$

Here the rate goes up by ______ when initial conc. doubles. Therefore, we say this reaction is _________________ order.

Now determine the value of k. Use expt. #3 data on earlier slide,

0.182 mol/L•s = k (0.30 mol/L)^2

k = 2.0 (L / mol•s)

Using k you can calc. rate at other values of $[CH_3CHO]$ at same T.

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Concentration/Time Relations

Chemists need to know what the concentration of reactant is as function of time.

***Most* reactions are first order (radioactive, biological, etc.)**

Let's consider FIRST ORDER REACTIONS

For 1st order reactions (1), the rate law is:

$$-(\Delta [R] / \Delta \text{time}) = k [R]^1$$

Integration (calculus): $\int_0^t \frac{1}{[R]} dR = -k \int_0^t dt$

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Concentration/Time Relations

Integrating **$-(\Delta [R] / \Delta \text{time}) = k [R]$** *we get:*

$$\ln \frac{[R]_t}{[R]_0} = -kt$$

natural logarithm

[R] at time = 0

$[R]_t / [R]_0$ = fraction remaining after time t has elapsed.

This is the integrated first-order rate law.

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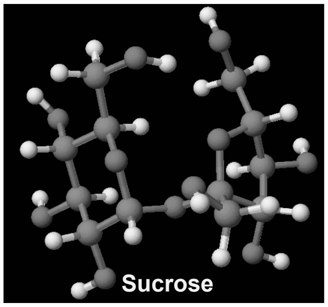
Concentration/Time Relations



Sucrose decomposes to simpler sugars

Rate of disappearance of sucrose = k [sucrose]

k = 0.21 hr^{-1}

Initial [sucrose] = 0.010 M

How long to drop 90% (to 0.0010 M)?

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Concentration/Time Relations

Rate of disappear of sucrose = k [sucrose], k = 0.21 hr^{-1}. If initial [sucrose] = 0.010 M, how long to drop 90% or to 0.0010 M?

Use the first order integrated rate law

$$\ln\left(\frac{0.0010 \text{ M}}{0.010 \text{ M}}\right) = -(0.21 \text{ hr}^{-1})\, t$$

ln (0.10) = - 2.30 = - (0.21 hr^{-1}) • time

sig figs and logarithms covered in CH 223

time = 11 hours

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Using the Integrated Rate Law

The integrated rate law suggests a way to tell if a reaction is first order based on experiment.

$2\ N_2O_5(g) \longrightarrow 4\ NO_2(g) + O_2(g)$

Rate = $k[N_2O_5]$

Time (min)	$[N_2O_5]$ (M)	ln $[N_2O_5]$
0	1.00	0
1.0	0.705	-0.35
2.0	0.497	-0.70
5.0	0.173	-1.75

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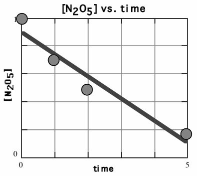
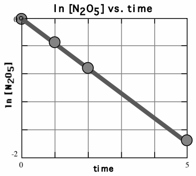
Using the Integrated Rate Law

$2\ N_2O_5(g) \longrightarrow 4\ NO_2(g) + O_2(g)$ Rate = $k[N_2O_5]$



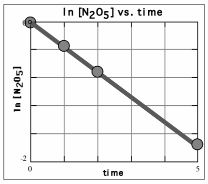
Data of conc. vs. time plot do not fit straight line.



Plot of ln $[N_2O_5]$ vs. time is a straight line!

MAR

Using the Integrated Rate Law



Plot of ln $[N_2O_5]$ vs. time is a straight line!
Eqn. for straight line:
y = ax + b

$$\ln [N_2O_5] = \underset{\substack{\Uparrow \\ \text{rate const} \\ = \text{slope}}}{-kt} + \underset{\substack{\Uparrow \\ \text{conc at} \\ \text{time = 0}}}{\ln [N_2O_5]_o}$$

$$\underset{\substack{\Uparrow \\ \text{conc at} \\ \text{time t}}}{\ln [N_2O_5]}$$

***All* 1st order reactions have straight line plot for ln [R] vs. time.**

(2nd order gives straight line for plot of 1/[R] vs. time; zero order [R] vs. time) - *see Handout*

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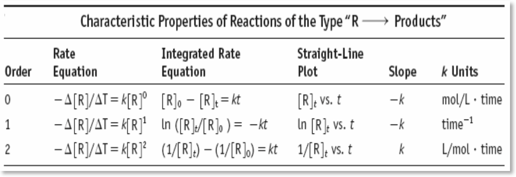
Using the Rate Laws



Characteristic Properties of Reactions of the Type "R ⟶ Products"

Order	Rate Equation	Integrated Rate Equation	Straight-Line Plot	Slope	k Units
0	$-\Delta[R]/\Delta T = k[R]^0$	$[R]_0 - [R]_t = kt$	$[R]_t$ vs. t	$-k$	mol/L · time
1	$-\Delta[R]/\Delta T = k[R]^1$	$\ln([R]_t/[R]_0) = -kt$	$\ln [R]_t$ vs. t	$-k$	time^{-1}
2	$-\Delta[R]/\Delta T = k[R]^2$	$(1/[R]_t) - (1/[R]_0) = kt$	$1/[R]_t$ vs. t	k	L/mol · time

Also see the "Kinetics Cheat Sheet" Handout

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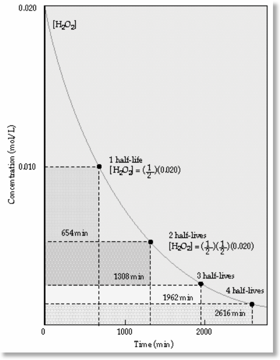
Half-Life

HALF-LIFE is the time it takes for 1/2 a sample is disappear.

$$\frac{N_t}{N_0} = 0.5^n$$

n = the number of half-lives

For 1st order reactions, the concept of HALF-LIFE is especially useful.



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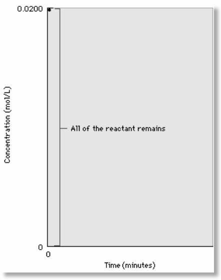

Half-Life

Reaction is 1st order decomposition of H_2O_2.

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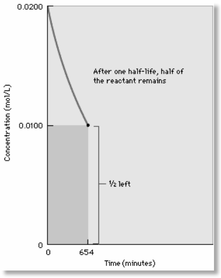

Half-Life

Reaction after 654 min, 1 half-life.

1/2 of the reactant remains.

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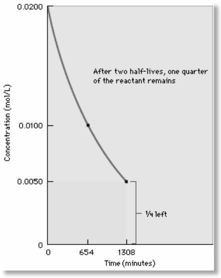

Half-Life

Reaction after 1308 min, or 2 half-lives.

1/4 of the reactant remains.

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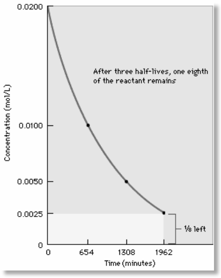

Half-Life

Reaction after 3 half-lives, or 1962 min.

1/8 of the reactant remains.

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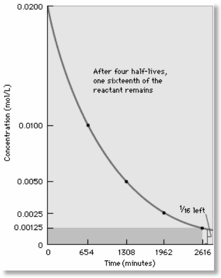

Half-Life

Reaction after 4 half-lives, or 2616 min.

1/16 of the reactant remains.

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Half-Life

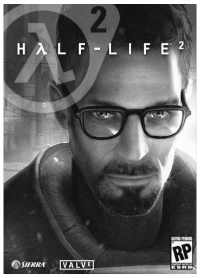
Sugar is fermented in a 1st order process (using an enzyme as a catalyst).

sugar + enzyme --> products

Rate of disappear of sugar = k[sugar]

$k = 3.3 \times 10^{-4}$ sec^{-1}

What is the half-life of this reaction?



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Half-Life

Rate = k[sugar] and $k = 3.3 \times 10^{-4}$ sec^{-1}. What is the half-life of this reaction?

Solution

$[R] / [R]_0 = 1/2$ when $t = t_{1/2}$

Therefore, $\ln (1/2) = - k \cdot t_{1/2}$

$- 0.693 = - k \cdot t_{1/2}$

$$t_{1/2} = 0.693 / k$$

So, for sugar,

$t_{1/2} = 0.693 / k = 2100$ sec = 35 min

see Handout

MAR

Half-Life

Rate = k[sugar] and $k = 3.3 \times 10^{-4}$ sec^{-1}. Half-life is 35 min. Start with 10.00 g sugar. How much is left after 2 hr and 20 min? *($t_{1/2} = 0.693 / k$)*

Two Solution pathways - *The "CH 104" method:*

2 hr and 20 min = 4 half-lives

Half-life	Time Elapsed	Mass Left
1st	**35 min**	**5.00 g**
2nd	**70**	**2.50 g**
3rd	**105**	**1.25 g**
4th	**140**	0.625 g *answer*

MAR

Half-Life

Rate = k[sugar] and $k = 3.3 \times 10^{-4}$ sec^{-1}. Half-life is 35 min. Start with 10.00 g sugar. How much is left after 2 hr and 20 min? *($t_{1/2} = 0.693 / k$)*

Two Solution Pathways - *The "CH 222" Method:*

Convert 2 hrs 20 min to 8400 seconds

$\ln (R / 10.00 \text{ g}) = - (3.3 \times 10^{-4} \text{ sec}^{-1})(8400 \text{ s})$

$\ln (R / 10.00 \text{ g}) = - 2.772$

Take *antilog*: **$(R / 10.00 \text{ g}) = e^{-2.772} = 0.0625$**

$R = 10.00 \text{ g} * 0.0625 =$ 0.625 g

CH 222 Method recommended,
not limited to whole number of half-lives, etc.
e = "Euler's number" in math = 2.71828...

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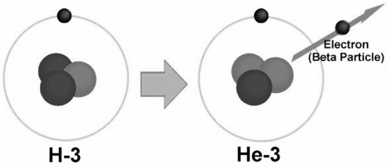
Half-Life

Radioactive decay is a first order process.

Tritium ---> electron + helium

$^{3}_{1}H$ $^{0}_{-1}e$ $^{3}_{2}He$

If you have 1.50 mg of tritium, how much is left after 49.2 years? $t_{1/2} = 12.3$ years



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Half-Life

Start with 1.50 mg of tritium, how much is left after 49.2 years? $t_{1/2} = 12.3$ years

$\ln [R] / [R]_0 = -kt$

$[R] = ?$ $[R]_0 = 1.50$ mg $t = 49.2$ yrs

Need k, so we calc k from: $k = 0.693 / t_{1/2}$

Obtain $k = 0.0564$ y^{-1}

Now $\ln [R] / [R]_0 = -kt = - (0.0564 \text{ y}^{-1}) \cdot (49.2 \text{ y})$

$= - 2.77$

Take antilog: $[R] / [R]_0 = e^{-2.77} =$ 0.0627

0.0627 is the fraction remaining

$[R]_0 = 1.50$ mg, so [R] = 1.50*0.0627 = 0.0941 mg

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Half-Life

Half-Life for first order reactions:

$$t_{1/2} = 0.693 / k$$

All biological, radioactive processes first order, most useful

Half-Life for second order reactions:

$$t_{1/2} = 1 / (k[A]_0)$$

Half-Life for zero order reactions:

$$t_{1/2} = [A]_0 / 2k$$

$[A]_0$ is the initial concentration of the reactant

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Half-Lives of Radioactive Elements

Rate of decay of radioactive isotopes given in terms of half-life ($t_{1/2}$).

$^{238}U \rightarrow\ ^{234}Th + He$	**4.5 x 10^9 y**
$^{14}C \rightarrow\ ^{14}N$ + beta	**5730 y**
$^{131}I \rightarrow\ ^{131}Xe$ + beta	**8.05 d**
Element 106 - seaborgium	
$^{263}Sg \rightarrow\ ^{259}Rf + He$	**0.8 s**
Element 111 - roentgenium	
$^{272}Rg \rightarrow\ ^{268}Mt + He$	**0.0015 s**

Half-life also used in medicine;
$t_{1/2}$ caffeine = 3.5 hours in the body

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MECHANISMS
A Microscopic View of Reactions

How are reactants converted to products at the molecular level?

We want to connect the

RATE LAW to the MECHANISM

the *experiment* to the *theory*

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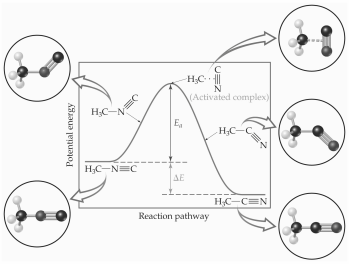
Reaction Coordinate Diagrams

It is helpful to visualize energy changes throughout a process on a reaction coordinate diagram.

This diagrams shows the rearrangement of methyl isocyanide to methyl cyanide (acetonitrile).



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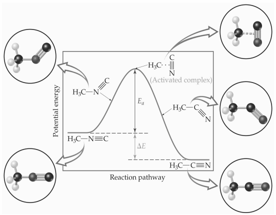
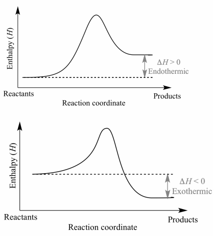
Reaction Coordinate Diagrams

The diagram shows the energy of the reactants and products (and, therefore, ΔE); *exo-* or *endo-*thermic!

The high point on the diagram is the transition state.





The species present at the transition state is called the activated complex.

The energy gap, E_a between the reactants and the activated complex is the activation energy barrier.

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Activation Energy

There is a minimum amount of energy required for a reaction: the activation energy, E_a.

In general, differences in activation energy are the reason reactions vary from fast to slow.

An analogy to chemical activation energy. In order for the volley ball to go over the net, the player must give it sufficient energy.

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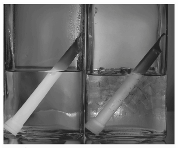
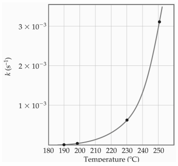
Temperature and Rate



Higher temperature Lower temperature



Generally, as temperature increases, so does the reaction rate.

This is because *k* is temperature dependent.

Find activation energy through changes in temperature.

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Effect of Temperature

Reactions generally occur slower at lower T.

Iodine clock reaction

$H_2O_2 + 2\ I^- + 2\ H^+ \text{-->}$
$2\ H_2O + I_2$

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Collision Theory

To break and make bonds, reactions require
(a) activation energy and
(b) correct geometry.

$O_3(g)\ +\ NO(g)\ \text{--->}\ O_2(g)\ +\ NO_2(g)$

Insufficient activation energy

Incorrect geometry

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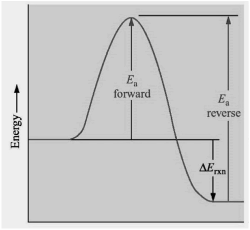
Collision Theory

To break and make bonds, reactions require
(a) activation energy and
(b) correct geometry.

$O_3(g)\ +\ NO(g)\ \text{--->}\ O_2(g)\ +\ NO_2(g)$

Correct activation energy and geometry



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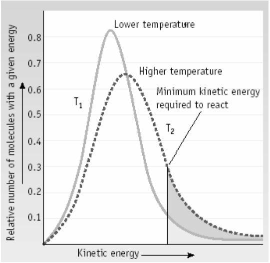
Activation Energy and Temperature

Reactions are slower at lower T because a smaller fraction of reactant molecules have enough energy to convert to product molecules.



In general, differences in activation energy cause reactions to vary from fast to slow.

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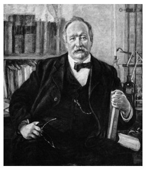

Svante Arrhenius

The Arrhenius Equation

Arrhenius equation -

$$k = Ae^{-E_a/RT}$$

Rate constant → k

Frequency factor → A

Activation energy → E_a

Temp (K) → T

8.3145 J/K•mol → R

Frequency factor = frequency of collisions with correct geometry.

$$\ln k = -\left(\frac{E_a}{R}\right)\left(\frac{1}{T}\right) + \ln A$$

Plot ln k vs. 1/T ---> *straight line.*
***slope* = -E_a/R,**
best way to find E_a

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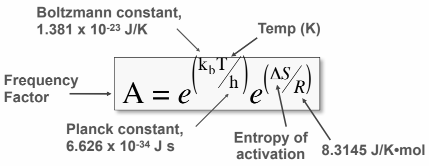
Always use 8.3145 for "energy" R!

More About the Frequency Factor, A

Frequency Factor often expressed as:



$$A = e^{\left(k_b T / h\right)} e^{\left(\Delta S / R\right)}$$

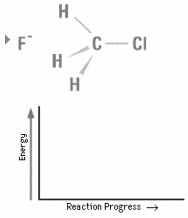
We will be seeing entropy (∆S) again in CH 223.

REACTION MECHANISMS

A reaction mechanism is the sequence of events that describes the actual process by which reactants become products.

Reactions may occur all at once ("single step") or through several discrete steps.

Each of these steps known as an elementary step



A *bimolecular* reaction

Exo- or endothermic?

MOLECULARITY

The **molecularity** of a process tells how many molecules are involved in the elementary step.

UNIMOLECULAR - only one reactant is involved.

BIMOLECULAR - two different molecules must collide

TERMOLECULAR - three different molecules collide

Elementary Reactions and Their Rate Laws

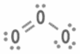
Molecularity	Elementary Reaction	Rate Law
*Uni*molecular	$A \longrightarrow$ products	Rate = $k[A]$
*Bi*molecular	$A + A \longrightarrow$ products	Rate = $k[A]^2$
*Bi*molecular	$A + B \longrightarrow$ products	Rate = $k[A][B]$
*Ter*molecular	$A + A + A \longrightarrow$ products	Rate = $k[A]^3$
*Ter*molecular	$A + A + B \longrightarrow$ products	Rate = $k[A]^2[B]$
*Ter*molecular	$A + B + C \longrightarrow$ products	Rate = $k[A][B][C]$

MECHANISMS

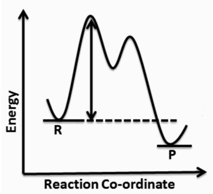
Some reactions occur in a single elementary step. *Most* reactions involve a sequence of elementary steps.

Adding elementary steps gives NET reaction.

Step 1	Unimolecular	$O_3(g) \longrightarrow O_2(g) + O(g)$
Step 2	Bimolecular	$O_3(g) + O(g) \longrightarrow 2\,O_2(g)$
Overall reaction		$2\,O_3(g) \longrightarrow 3\,O_2(g)$



Every elementary step gets a "hump" in the diagram

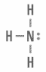


MECHANISMS

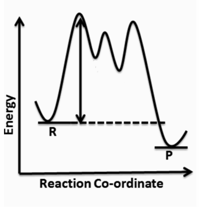
Most reactions involve a *sequence* of elementary steps.

Adding elementary steps gives NET reaction.

Step 1	Bimolecular	$NH_3 + OCl^- \longrightarrow NH_2Cl + OH^-$
Step 2	Bimolecular	$NH_2Cl + NH_3 \longrightarrow N_2H_5^+ + Cl^-$
Step 3	Bimolecular	$N_2H_5^+ + OH^- \longrightarrow N_2H_4 + H_2O$
Overall reaction		$2\,NH_3 + OCl^- \longrightarrow N_2H_4 + H_2O + Cl^-$



Every elementary step gets a "hump" in the diagram



Mechanisms

In multistep mechanisms, one step will be slower than all others - this is the rate determining step (rds).

Most reactions involve a sequence of elementary steps. *Example:*

$$2\,I^- + HOOH + 2\,H^+ \text{ ---> } I_2 + 2\,H_2O$$

$$\text{Rate} = k\,[I^-]\,[H_2O_2]$$

Note that the **rate law comes from experiment**.

Also note that *order of reactants* in rate law *not necessarily the same as stoichiometric coefficients*!

MECHANISMS

$2\ I^- + HOOH + 2\ H^+ \text{ ---> } I_2 + 2\ H_2O$

Rate = k $[I^-]$ [HOOH]

A scientist proposes the following mechanism for this reaction:

Step 1 - slow $HOOH + I^- \text{ --> } HOI + OH^-$

Step 2 - fast $HOI + I^- \text{ --> } I_2 + OH^-$

Step 3 - fast $2\ OH^- + 2\ H^+ \text{ --> } 2\ H_2O$

Note that H^+ not involved in slow step, so $[H^+]$ is not in the rate law (a zero order reactant.)

Bimolecular slow step reflected in the rate law; this is a legitimate mechanism.

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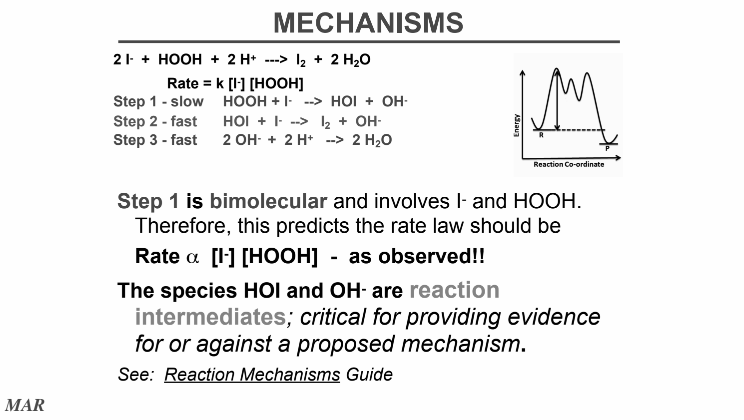


MECHANISMS

$2\ I^- + HOOH + 2\ H^+ \text{ ---> } I_2 + 2\ H_2O$

Rate = k $[I^-]$ [HOOH]

Step 1 - slow $HOOH + I^- \text{ --> } HOI + OH^-$

Step 2 - fast $HOI + I^- \text{ --> } I_2 + OH^-$

Step 3 - fast $2\ OH^- + 2\ H^+ \text{ --> } 2\ H_2O$

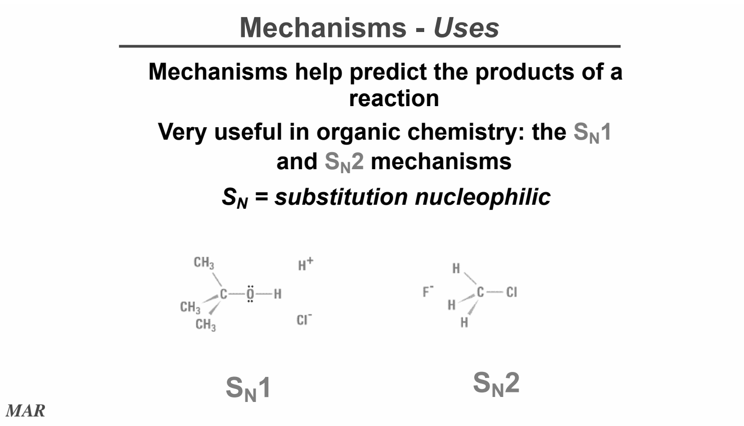
Step 1 is bimolecular and involves I^- and HOOH. Therefore, this predicts the rate law should be

Rate α $[I^-]$ [HOOH] - as observed!!

The species HOI and OH^- are reaction intermediates; *critical for providing evidence for or against a proposed mechanism*.

See: Reaction Mechanisms Guide

MAR



Mechanisms - *Uses*

Mechanisms help predict the products of a reaction

Very useful in organic chemistry: the S_N1 and S_N2 mechanisms

S_N = substitution nucleophilic

S_N1 S_N2

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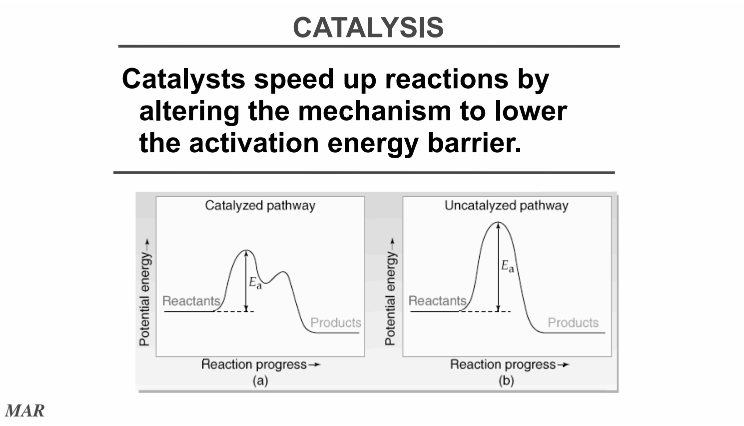

CATALYSIS

Catalysts speed up reactions by altering the mechanism to lower the activation energy barrier.

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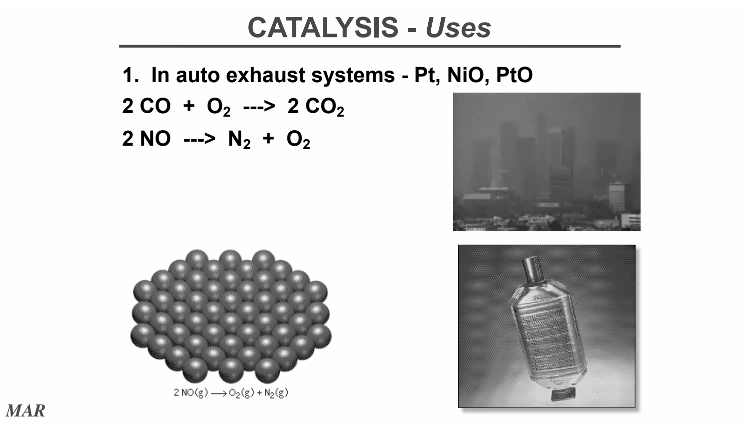


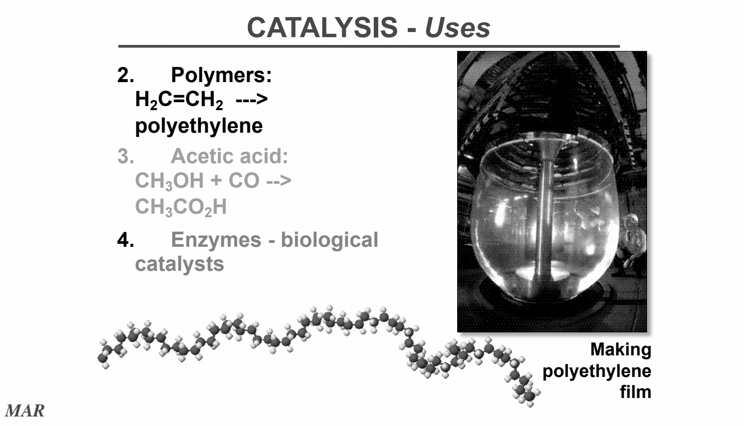
CATALYSIS - *Uses*

1. In auto exhaust systems - Pt, NiO, PtO

$2\ CO + O_2 \text{ ---> } 2\ CO_2$

$2\ NO \text{ ---> } N_2 + O_2$

MAR



CATALYSIS - *Uses*

2. Polymers: $H_2C{=}CH_2$ ---> polyethylene

3. Acetic acid: $CH_3OH + CO \text{ --> } CH_3CO_2H$

4. Enzymes - biological catalysts

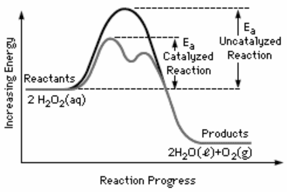
Making polyethylene film

MAR

CATALYSIS

Catalysis and activation energy

MnO_2 catalyzes decomposition of H_2O_2
$2\ H_2O_2$ ---> $2\ H_2O + O_2$



— **Uncatalyzed reaction**

— **Catalyzed reaction**

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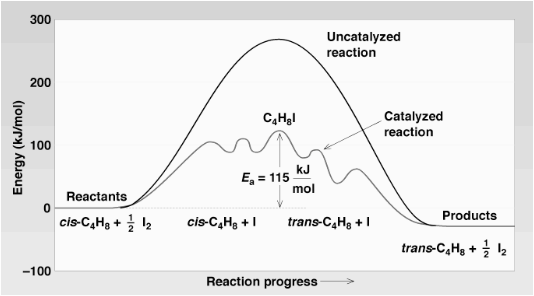
Iodine-Catalyzed Isomerization of cis-2-Butene



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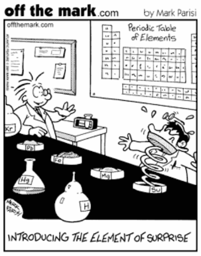

End of Chapter 12

See:

- *Chapter Twelve Study Guide*
- *Chapter Twelve 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:

The Rate Law:

$$Rate = k[A]^p[B]^m[C]^n \ldots$$

m, n, p = 0, 1 or 2 only (in our classes)

1st Order Integrated Rate Law;

$$ln\frac{[R]}{[R_0]} = -kt$$

$$t_{\frac{1}{2}} = \frac{0.693}{k}$$

The Arrhenius Equation:

$$ln(k) = -\left(\frac{E_a}{R}\right)\left(\frac{1}{T}\right) + ln(A)$$

- **R = 8.3145 J/mol·K**
- **"Kinetics Cheat Sheet"** *handout*
- **"Reactions Mechanisms"** *handout*

Kinetics: rate, rate law, orders of reaction, the rate constant (k), 1st vs. 2nd. vs. zero order, half life, mechanism, elementary reaction, bimolecular (and uni- *and* ter-molecular), Arrhenius equation, activation energy, frequency factor, mechanism, intermediate, catalyst, rds (rate determining step)

MAR

End of Chapter Problems: ***Test Yourself***

1. The reaction between ozone and nitrogen dioxide at 231 K is first order in both NO_2 and O_3: $2\ NO_2(g) + O_3(g) \rightarrow N_2O_5(s) + O_2(g)$ Write the rate equation for the reaction. If the concentration of NO_2 is tripled, what is the change in the reaction rate?
2. After 2.57 h at 27 °C, a first order sucrose concentration decreased from 0.0146 M to 0.0132 M. Find the rate constant, *k*.
3. The compound $Xe(CF_3)_2$ decomposes in a first-order reaction to elemental Xe with a half-life of 30. min. If you place 7.50 mg of $Xe(CF_3)_2$ in a flask, how long must you wait until only 0.25 mg of $Xe(CF_3)_2$ remains?
4. Gaseous NO_2 decomposes at 573 K: $2\ NO_2(g) \rightarrow 2\ NO(g) + O_2(g)$ The concentration of NO_2 was measured as a function of time. A graph of $1/[NO_2]$ versus time gives a straight line with a slope of 1.1 L/mol.s. What is the rate law for this reaction? What is the rate constant k?
5. What is the rate law for the following *elementary* reaction: $NO(g) + NO_3(g) \rightarrow 2\ NO_2(g)$
6. For a reaction, ln k versus 1/T(K) is plotted, and the linear regression line is: y = -6373.3x + 18.19, r = -0.997 What is the activation energy for this reaction?

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End of Chapter Problems: ***Answers***

1. Rate = k[NO2][O3]; tripling NO triples the rate.
2. 0.0392 h^{-1}
3. 150 min
4. Rate = $k[NO_2]^2$, k = 1.1 L/mol·s
5. Rate = $k[NO][NO_3]$
6. 53.0 kJ/mol

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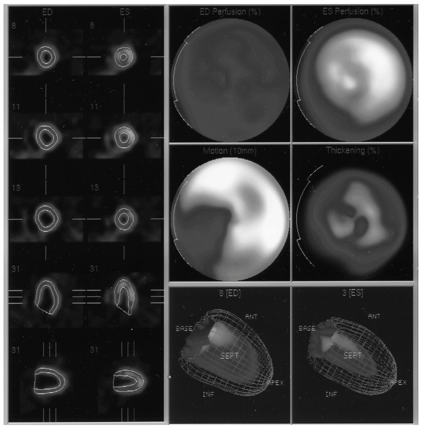
Nuclear Chemistry

Chapter 21

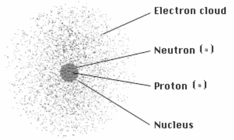


Chemistry 222
Professor Michael Russell

MAR

Last update: 4/29/24

Nuclear Chemistry: *the study of the nucleus*



Nucleus = neutrons and protons
Differs from "normal" electron chemistry

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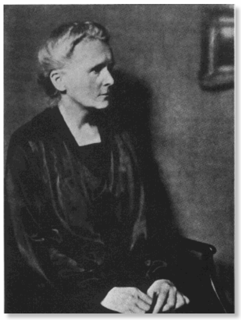
Early Nuclear Chemistry Pioneers:

Marie Curie



Marie Curie
Pierre Curie
Henri Becquerel

Noticed beams of light on photographic plates

Danger (and potential) of nuclear chemistry poorly understood

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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

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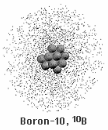
Mass Number, A

Mass Number
= # protons + # neutrons

A boron atom can have
A = 5 p + 5 n = 10 *and*
Z = 5 protons = 5

$$\begin{matrix} A \rightarrow & 10 \\ Z \rightarrow & 5 \end{matrix}\mathrm{B}$$



Isotopes of boron have different # of neutrons but same # of protons *(Boron-10, Boron-11)*

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Radioactive Particles

ALPHA	**helium nucleus**	$^{4}_{2}\mathrm{He}$ ***or*** $^{4}_{2}\alpha$
BETA	**electron**	$^{0}_{-1}\mathrm{e}$ ***or*** $^{0}_{-1}\beta$
GAMMA	***energy (massless)***	γ
PROTON	**proton**	$^{1}_{1}\mathrm{p}$ ***or*** $^{1}_{1}\mathrm{H}$
NEUTRON	**neutron**	$^{1}_{0}\mathrm{n}$
POSITRON	**antielectron**	$^{0}_{+1}\mathrm{e}$ ***or*** $^{0}_{+1}\beta$

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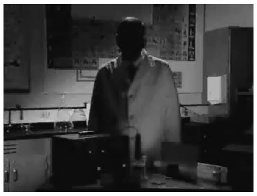
see: Nuclear Chemistry Guide

Radioactive Particles

Each particle has different properties



Alpha **- stopped by clothes**
Beta **- stopped by skin**
Gamma **- stopped by lead**

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Balancing Nuclear Reactions

*** Number of *reactant* protons must equal the number of *product* protons**

*** Number of *reactant* neutrons must equal the number of *product* neutrons**

Example:

$$^{226}_{88}\text{Ra} \longrightarrow {}^{4}_{2}\text{He} + {}^{222}_{86}\text{Rn}$$

A: $226 = 4 + 222$
Z: $88 = 2 + 86$

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Balancing Nuclear Reactions

Problem: **Uranium-235 decomposes through** beta decay **to a new product. Find the identity of the new product.**

Solution:

Uranium has 92 protons, so: $^{235}_{92}\text{U}$

Beta decay **means "losing a beta particle", or generating a beta particle as a product, so:**

$$^{235}_{92}\text{U} \longrightarrow {}^{0}_{-1}\text{e} + {}^{A}_{Z}\text{X}$$

MAR

Balancing Nuclear Reactions

Problem: **Uranium-235 decomposes through beta decay to a new product. Find the identity of the new product.**

$$^{235}_{92}\text{U} \longrightarrow {}^{0}_{-1}\text{e} + {}^{A}_{Z}\text{X}$$

Solution:

235 = 0 + A, therefore A = 235
92 = -1 + Z, therefore Z = 93
If Z = 93, X = Neptunium (Np**), and**

$$\boxed{^{235}_{92}\text{U} \longrightarrow {}^{0}_{-1}\text{e} + {}^{235}_{93}\text{Np}}$$

The "inner reaction":

$$^{1}_{0}\text{n} \longrightarrow {}^{0}_{-1}\text{e} + {}^{1}_{1}\text{p}$$

MAR

Emission, Decay and Capture

Emission **and** Decay **imply a *product* particle**

An example of positron emission**:**

$$\boxed{^{207}_{84}\text{Po} \longrightarrow {}^{0}_{+1}\beta + {}^{207}_{83}\text{Bi}}$$

Capture **implies a *reactant* particle**

An example of neutron capture**:**

$$\boxed{^{19}_{9}\text{F} + {}^{1}_{0}\text{n} \longrightarrow {}^{20}_{9}\text{F}}$$

Many types of "particles" and "decays"

MAR

Positrons and Antimatter

Paul Dirac **first predicted antimatter in 1928, identical to "regular" matter except for opposite charge**

Positrons **are** antielectrons**; combining with electrons leads to** annihilation ***(gamma)***

$$\boxed{^{0}_{-1}\text{e} + {}^{0}_{+1}\text{e} \longrightarrow 2\gamma}$$

electron + positron ---> gamma radiation
Used in Positron Emission Tomography (PET)

Many examples of annihilation reactions **are known:**

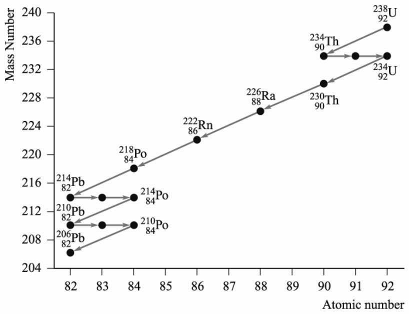
$$\boxed{\nu + \bar{\nu} \longrightarrow 2\gamma}$$

neutrino + antineutrino ---> gamma radiation

Neutrino first postulated by Pauli (CH 221)

MAR

Decay Scheme for Uranium-238



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Metastable Nuclei

When *electrons* are excited in an atom, they relax and emit UV, IR and visible radiation

A metastable nuclei is created when a *nucleus* is excited. Relaxation results in high energy gamma ray emission

$$\boxed{^{99m}_{43}\text{Tc} \longrightarrow {}^{99}_{43}\text{Tc} + \gamma}$$

metastable Tc ---> "regular" Tc + gamma ray
Note the "m" for "metastable" and no change in atomic or mass numbers

Metastable nuclei (*especially Technetium-99m*) used extensively in medical imaging

MAR

The Four Forces of Nature

Gravitation: weak, long range force

Weak nuclear: short range, gives beta decay

Electromagnetic (E/M): long range, keeps electrons around nucleus, 10^{36} times more powerful than gravity (*aka* electrostatic, Coulombic, etc.)

Strong nuclear: 10^{6} times more powerful than E/M, *very* short range (10^{-15} m); overpowers E/M repulsion between protons. Strong keeps (protons + neutrons) and (neutrons + neutrons) together. "Glue" that keeps nucleus together

MAR

Nuclear Stability

Why are nuclei unstable? E/M Force (long range) begins overpowering strong force (short range)

Magnetic dilution* (more neutral neutrons than positive protons) helps stabilize nuclei - *to a point

<u>Up to calcium</u> (Z = 20), most stable nuclei occur when # protons = # neutrons

Exceptions: **helium-3 and hydrogen-1**

<u>Up to lead</u> (Z = 82), most stable nuclei occur when # protons < # neutrons

<u>Beyond lead</u>, all isotopes unstable and radioactive

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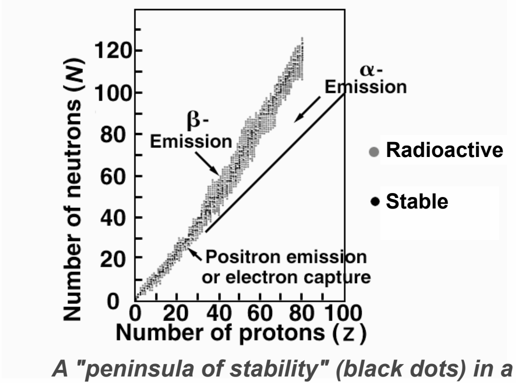

A "peninsula of stability" (black dots) in a "sea of instability" (red dots)

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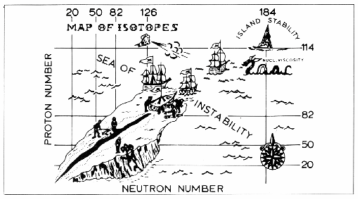
An "Island of Stability"?



Glenn Seaborg and Ken Moody believe that heavier elements can be made - *exciting!*

Interested? See: http://www.pbs.org/wgbh/nova/sciencenow/3313/02.html

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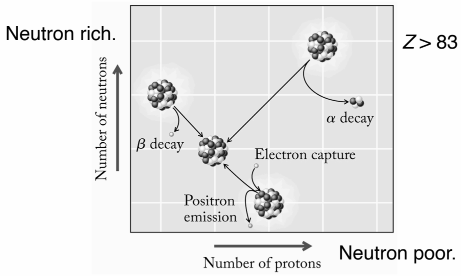
Nuclear Stability

Isotopes *often* decay based on their number of neutrons:



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Nuclear Stability

β-decay occurs in isotopes with a high neutron to proton ratio. *Result:* higher atomic number, same mass number

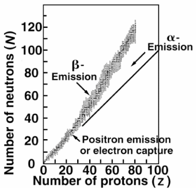
$$^{235}_{92}\mathrm{U} \longrightarrow {}^{0}_{-1}\mathrm{e} + {}^{235}_{93}\mathrm{Np}$$

Low neutron to proton ratio: **positron emission** or **electron capture**. *Result:* lower atomic number, same mass number

$$^{13}_{7}\mathrm{N} \longrightarrow {}^{0}_{+1}\mathrm{e} + {}^{13}_{6}\mathrm{C}$$

All elements beyond Bi decay, usually by **α-decay**

$$^{226}_{88}\mathrm{Ra} \longrightarrow {}^{4}_{2}\mathrm{He} + {}^{222}_{86}\mathrm{Rn}$$



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Nuclear Stability

Energy required to overcome positive-positive repulsion of protons is substantial - use strong *(and weak)* forces

- *Strong and weak forces much stronger than electromagnetic force or gravity*

Nuclear binding energy, E_b, used to estimate force contribution

- **E_b is the negative of the energy change if nucleus formed directly from individual protons and neutrons**

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Nuclear Stability

Problem: Calculate the binding energy, E_b, for deuterium (hydrogen-2).

Solution:

The process: $^{1}_{1}\mathrm{p} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{2}_{1}\mathrm{H}$

Note that: $\mathrm{mass_p} + \mathrm{mass_n} \neq \mathrm{mass_D}$

$1.007825 + 1.00865 \neq 2.01410$

$2.016475 \neq 2.01410$

$\Delta m = -0.00239$ g/mol

mass is *not* conserved!

MAR

Nuclear Stability

Problem: Calculate the binding energy, E_b, for deuterium (hydrogen-2). $\Delta m = -0.00239$ g/mol

Solution:

Now use <u>Einstein's equation</u>: $\Delta E = \Delta mc^2$

$\Delta E = (-2.39 \times 10^{-6}\ \mathrm{kg\ mol^{-1}})(2.998 \times 10^{8}\ \mathrm{m\ s^{-1}})^2$

$\Delta E = -2.15 \times 10^{11}\ \mathrm{J\ mol^{-1}}$, ***and***

$E_b = -\Delta E$, **so**

$E_b = 2.15 \times 10^{8}\ \mathrm{kJ\ mol^{-1}}$

E very negative (exothermic) - lots of energy produced. E_b very positive - lots of energy saved through stabilization

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Nuclear Stability

Problem: Calculate the binding energy, E_b, for deuterium (hydrogen-2). $\Delta m = -0.00239$ g/mol

Solution:

Now use <u>Einstein's equation</u>: $\Delta E = \Delta mc^2$

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$\Delta E = -2.15 \times 10^{11}\ \mathrm{J\ mol^{-1}}$, ***and***

$E_b = -\Delta E$, **so**

$E_b = 2.15 \times 10^{8}\ \mathrm{kJ\ mol^{-1}}$

<u>E_b per nucleon</u> $= 2.15 \times 10^{8} / (1\ \mathrm{p} + 1\ \mathrm{n})$ ***nucleons***

$= 1.08 \times 10^{8}\ \mathrm{kJ\ mol^{-1}\ nucleon^{-1}}$

This is the strong force contribution in the nucleus!

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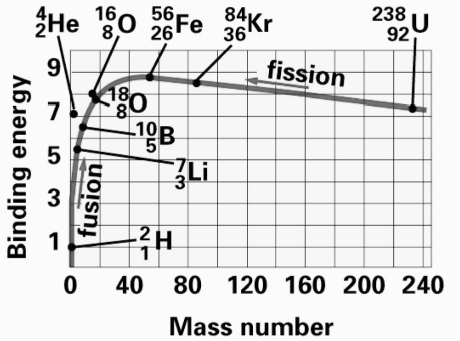
Nuclear Stability

Can use Binding energy (E_b) to calculate stability of nuclei:



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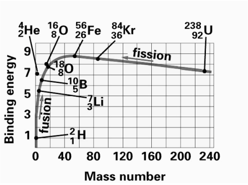
Nuclear Stability

Note that iron-56 is the most stable element *thermodynamically* - end of solar cycles, etc.

Elements with Z < 26 can use fusion to become more like iron-56

Elements with Z > 26 can use fission to become more like iron-56

Kinetics of nuclear reactions important - half life, rate of decay, etc.



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Tremendous amounts of energy are generated when light nuclei combine to form heavier nuclei - nuclear fusion

"Clean" energy, used in stars and bombs, requires *plasma* and/or high temperatures - high activation energy barrier

No "meltdown" - reaction just stops, no waste

$$^2_1\mathrm{H} + ^3_1\mathrm{H} \longrightarrow ^4_2\mathrm{He} + ^1_0\mathrm{n}$$

E = $-1.7*10^9$ kJ/mol

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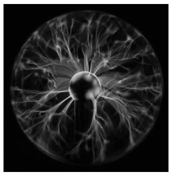
Fusion

Plasma

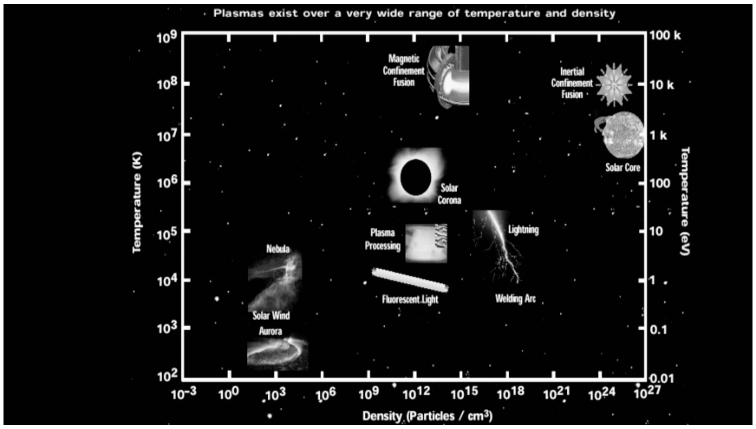


Plasma is the fourth state of matter - along with solids, liquids and gases

Plasma is an electrically conducting fluid composed of freely roaming electrons and ions; strong magnetic and electric fields; hot!

Plasmas comprise the vast majority of the apparent universe, and only in occasional "islands" (like the planet Earth) is matter found in condensed forms (solids, liquids, gases)!

MAR



Plasma - *in your microwave?*

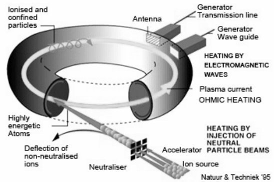
Video by Sergiu and Ben Todor 2010

Grape microwave "plasma" - ?!?

MAR

Fusion - *Tokamak Reactor*

Tokamak reactor uses magnetic fields to constrict plasma for fusion in "donut" shape; most promising "future" magnetic fusion device



Natuur & Techniek '95
D.A.Gorissen

Figure 2: Various tokamak plasma heating techniques

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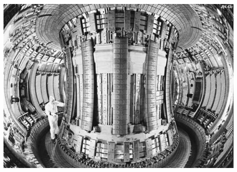
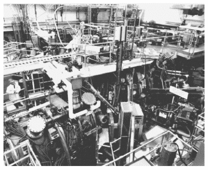
Fusion - *Tokamak Reactor*





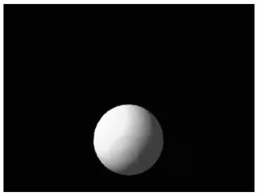
Tokamak reactors - *inside* and *outside* views

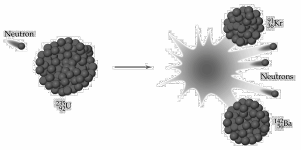
MAR

Fission

Tremendous amounts of energy *(electricity)* generated when heavy nuclei split to form lighter nuclei - *nuclear fission*

Generally requires a "neutron trigger"





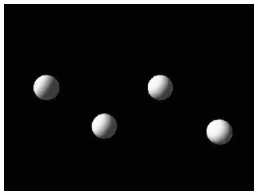
MAR

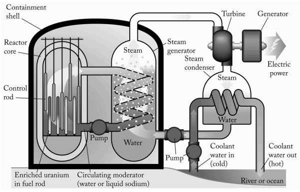
Fission

Excess neutrons *must* be controlled! *Supercritical chain reactions* can result without proper caution

Waste products from fission messy and virtually perpetual

More than 400 nuclear fission plants in 30 different countries!

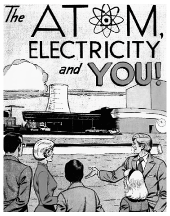




MAR

Uses for Nuclear Chemistry

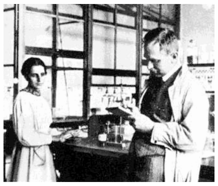
Energy: Fission (commercial) *and* Fusion (coming!)



Nuclear fission "problem events":

- SL-1, Idaho Falls (1961)
- Three Mile Island, Pennsylvania (1979)
- Chernobyl, Ukraine (1986)
- Fukushima Daiichi, Japan (2011)
- Hanford, Washington (ongoing!)

LISE MEITNER - *unsung hero*



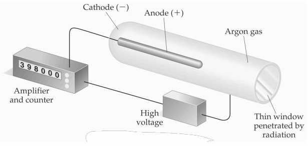
With Otto Hahn

- **Meitner and colleagues Otto Hahn, Fritz Strassmann & Otto Frisch explained the process of fission (1938)**
- **Meitnerium (Mt, #109) named after her**
- **Forced to work in basement, never got Nobel Prize (but Hahn did!)**
- **Pioneering woman in a male-dominated field; deserves more credit for her work**

MAR

Radiation and Health

Some radiation is around us all of the time



Geiger counters measure the amount of radioactive activity present in a sample. The radiation creates ions which conduct a detectable current.

MAR

Radiation and Health

Many units for radiation exposure - röntgen (R), curie (Ci), rad, rem, *others*

The röntgen is a measure of radiation exposure proportional to air ionization (*X*- and *γ*-rays)

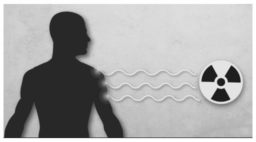
***Ex:* chest X-ray = 0.1 R**

rem = röntgen equivalent man

One curie = quantity of any isotope that undergoes 3.7 * 10^{10} *dps* (disintegrations per second)

One rad = 1.00 * 10^{-5} J absorbed per gram of material.

A whole-body dose of 450 rad = LD_{50} for humans

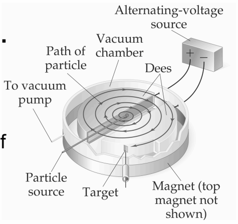


MAR

Nuclear Transformations

Nuclear transformations can be induced by accelerating a particle and colliding it with the nuclide.

- **Linear accelerator**—uses alternating electric fields to accelerate particle.
- **Cyclotron**—uses combination of electric and magnetic fields to accelerate particles in spiral pathway.



Cyclotron

MAR

Particle Accelerators

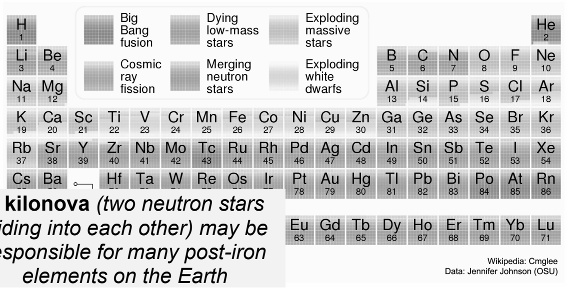
These particle accelerators are *enormous*, having circular tracks with radii that are miles long. New elements created by "smashing" smaller nuclei together.

CERN Particle Accelerator, Geneva, Switzerland

MAR

Uses for Nuclear Chemistry

Astronomers and chemists work together to define the origin of elements. Many made through fusion, some via fission, much not understood



Wikipedia: Cmglee
Data: Jennifer Johnson (OSU)

*A **kilonova** (two neutron stars sliding into each other) may be responsible for many post-iron elements on the Earth*

MAR

NSCL = *National Superconducting Cyclotron Laboratory at Michigan State University (http://www.nscl.msu.edu/)*

Full video *available here:* **https://youtu.be/677ZmPEFIXE**

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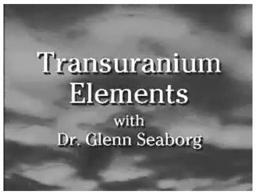
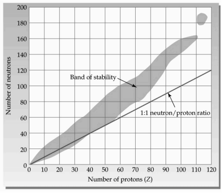
Uses for Nuclear Chemistry

Uses for Nuclear Chemistry

Expanding the Frontiers of Science





Dr. Glenn Seaborg (1912 - 1999)

MAR

Uses for Nuclear Chemistry

Radiocarbon Dating

$t_{1/2}$ = 5730 years for carbon-14
Accurate up to 60,000 years old!

MAR

Uses for Nuclear Chemistry

Medicine: PET *and* MRI

Tc (in PET) not found in nature, created via Mo:

$$ {}^{99}_{42}\text{Mo} \longrightarrow {}^{0}_{-1}\text{e} + {}^{99\text{m}}_{43}\text{Tc} $$

- Tc-99*m*: metastable; $t_{1/2}$ = 6.0 hrs.
- Can be incorporated into variety of compounds to target specific organs (heart, etc.).

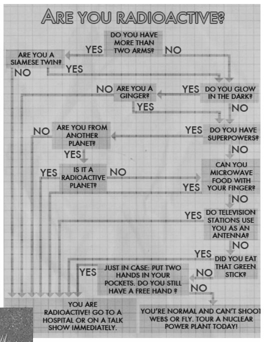
PET = Positron Emission Tomography
MRI = Magnetic Resonance Imaging

MAR

End of Chapter 21

See:

- *Chapter Twenty-one Study Guide*
- *Chapter Twenty-one Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*



MAR

Important Equations, Constants, and Handouts from this Chapter:

- **all** *of the* **first order kinetics** *equations apply. See the* **Nuclear Chemistry Guide** *(handout)*
- **decay** *or* **emission = product**
- **capture = reactant**
- **know how to balance nuclear reactions**

$$ -E_b = \Delta E = \Delta mc^2 $$

c = *speed of light* =
2.998 x 10^8 m/s
use **kg/mol** *for* Δ**m**

1st Order Integrated Rate Law;

$$ ln\frac{[R]}{[R_0]} = -kt $$

$$ t_{\frac{1}{2}} = \frac{0.693}{k} $$

MAR

End of Chapter Problems: *Test Yourself*

1. What particle is emitted when Gold-198 decays to mercury-198?
2. What particle is emitted when radon-222 decays to polonium-218?
3. What particle is emitted when indium-110 decays to cadmium-110?
4. What is emitted when hafnium-178m decays?
5. Boron has two stable isotopes, ^{10}B and ^{11}B. Calculate the binding energies per nucleon of these two nuclei. The required masses (in grams per mole) are proton = 1.00783, neutron = 1.00867, boron-10 = 10.01294, and boron-11 = 11.00931.
6. Gold-198 is used in the diagnosis of liver problems. The half-life of ^{198}Au is 2.69 days. If you begin with 2.8 μg of this gold isotope, what mass remains after 10.8 days?

MAR

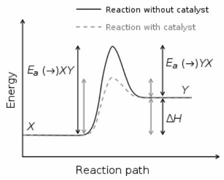
End of Chapter Problems: ***Answers***

1. beta particle
2. alpha particle
3. positron particle
4. hafnium-178
5. boron-10: 6.26×10^8 kJ/mol·nucleons; boron-11: 6.70×10^8 kJ/mol·nucleons
6. 0.17 μg

MAR

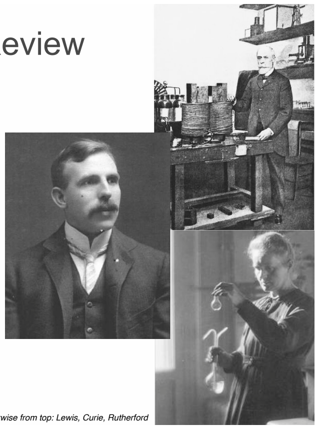
Chemistry 222 Final Exam Review

Chapters 12 and 21



Last update: 4/29/24

Chemistry 222
Professor Michael Russell



Clockwise from top: Lewis, Curie, Rutherford

For the reaction below, what is $\Delta[CH_2O]/\Delta t$ with respect to $\Delta[O_2]/\Delta t$?

$$C_2H_4(g) + O_3(g) \rightarrow 2\ CH_2O(g) + \tfrac{1}{2}\ O_2(g)$$

A. $\Delta[CH_2O]/\Delta t = 2 \bullet \Delta[O_2]/\Delta t$
B. $\Delta[CH_2O]/\Delta t = \tfrac{1}{4} \bullet \Delta[O_2]/\Delta t$
C. $\Delta[CH_2O]/\Delta t = -4 \bullet \Delta[O_2]/\Delta t$
D. $\Delta[CH_2O]/\Delta t = 4 \bullet \Delta[O_2]/\Delta t$
E. $\Delta[CH_2O]/\Delta t = \Delta[O_2]/\Delta t$

The reduction of NO with hydrogen produces nitrogen and water.

$$2\ NO + 2\ H_2 \rightarrow N_2 + 2\ H_2O$$

The reaction is second order in NO and third order overall. The rate law for the reaction is

A. Rate = k $[NO][H_2]$
B. Rate = k $[NO][H_2]^2$
C. Rate = k $[NO]^2[H_2]$
D. Rate = k $[NO]^2[H_2]^2$
E. Rate = k $[NO_2][H^+]^3$

The reduction of NO with hydrogen produces nitrogen and water.

$$2\ NO + 2\ H_2 \rightarrow N_2 + 2\ H_2O$$

The reaction is second order in NO and third order overall. The [NO] is increased by a factor of 5, the rate will increase by a factor of

A. 0
B. 1
C. 5
D. 10
E. 25

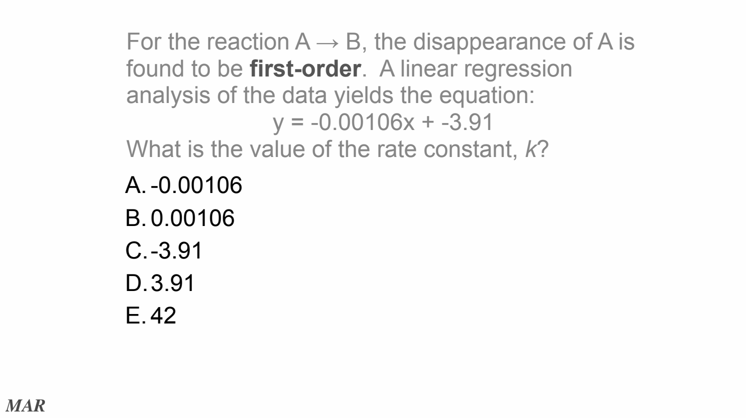
Given the initial rate data for the reaction A + B → C, determine the rate expression for the reaction.

[A] (M)	[B] (M)	Δ[C]/Δt (M/s)
0.10	0.20	40.
0.20	0.20	80.
0.10	0.10	40.

A. $\Delta[C]/\Delta t = 2000[A][B]$
B. $\Delta[C]/\Delta t = 40.[A]^2$
C. $\Delta[C]/\Delta t = 4.0[B]$
D. $\Delta[C]/\Delta t = 400[A]$
E. Δ[C]/Δt = #1[AC/DC]

For the reaction A → B, the disappearance of A is found to be **second-order**. Which of the following will produce a straight line graph?

A. log [A] vs. time
B. ln [A] vs. time
C. [A] vs. time
D. 1/[A] vs. time
E. $[A]^2$ vs. time



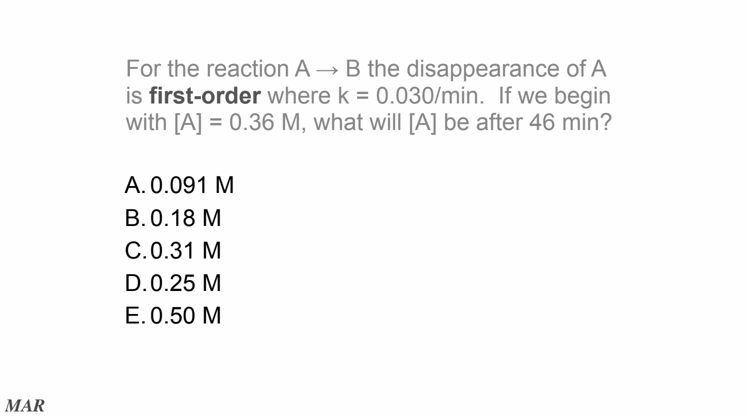
For the reaction A → B, the disappearance of A is found to be **first-order**. A linear regression analysis of the data yields the equation:

$$y = -0.00106x + -3.91$$

What is the value of the rate constant, *k*?

A. -0.00106
B. 0.00106
C. -3.91
D. 3.91
E. 42

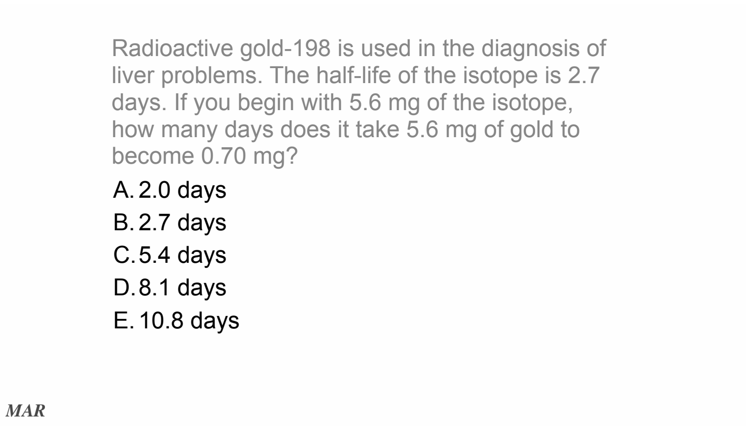
MAR



For the reaction A → B the disappearance of A is **first-order** where k = 0.030/min. If we begin with [A] = 0.36 M, what will [A] be after 46 min?

A. 0.091 M
B. 0.18 M
C. 0.31 M
D. 0.25 M
E. 0.50 M

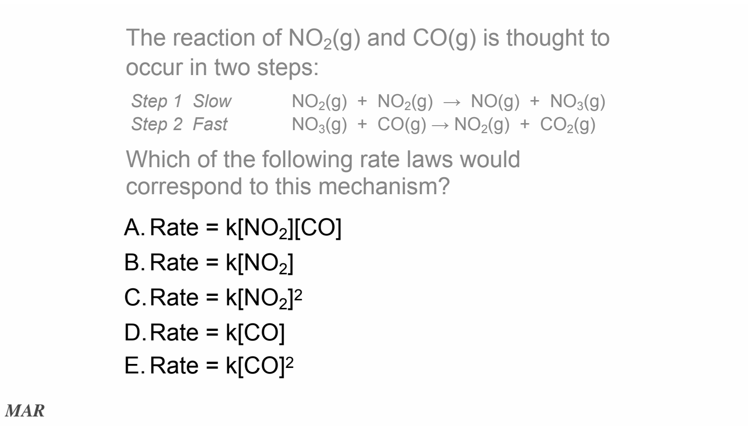
MAR



Radioactive gold-198 is used in the diagnosis of liver problems. The half-life of the isotope is 2.7 days. If you begin with 5.6 mg of the isotope, how many days does it take 5.6 mg of gold to become 0.70 mg?

A. 2.0 days
B. 2.7 days
C. 5.4 days
D. 8.1 days
E. 10.8 days

MAR



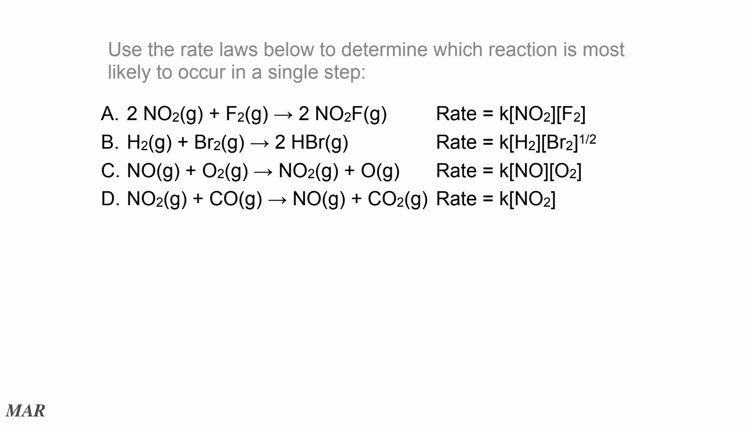
The reaction of $NO_2(g)$ and CO(g) is thought to occur in two steps:

Step 1 Slow $NO_2(g) + NO_2(g) \rightarrow NO(g) + NO_3(g)$
Step 2 Fast $NO_3(g) + CO(g) \rightarrow NO_2(g) + CO_2(g)$

Which of the following rate laws would correspond to this mechanism?

A. Rate = $k[NO_2][CO]$
B. Rate = $k[NO_2]$
C. Rate = $k[NO_2]^2$
D. Rate = $k[CO]$
E. Rate = $k[CO]^2$

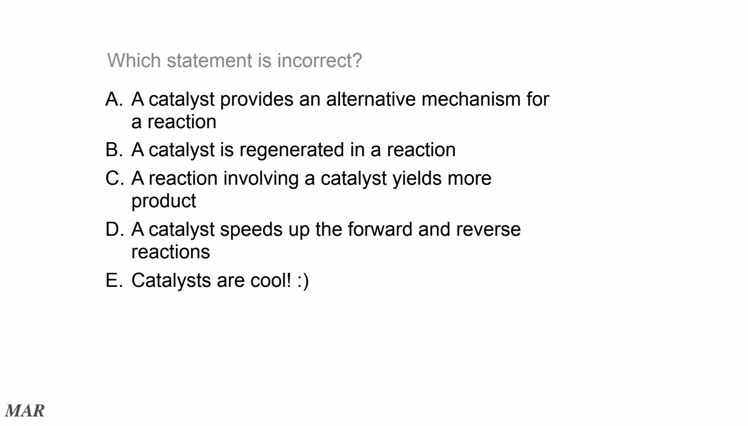
MAR



Use the rate laws below to determine which reaction is most likely to occur in a single step:

A. $2\ NO_2(g) + F_2(g) \rightarrow 2\ NO_2F(g)$ Rate = $k[NO_2][F_2]$
B. $H_2(g) + Br_2(g) \rightarrow 2\ HBr(g)$ Rate = $k[H_2][Br_2]^{1/2}$
C. $NO(g) + O_2(g) \rightarrow NO_2(g) + O(g)$ Rate = $k[NO][O_2]$
D. $NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g)$ Rate = $k[NO_2]$

MAR



Which statement is incorrect?

A. A catalyst provides an alternative mechanism for a reaction
B. A catalyst is regenerated in a reaction
C. A reaction involving a catalyst yields more product
D. A catalyst speeds up the forward and reverse reactions
E. Catalysts are cool! :)

MAR

How many neutrons and protons are there in the radioisotope ^{60}Co that is used in cancer therapy?

A. 60 neutrons and 27 protons
B. 27 neutrons and 60 protons
C. 33 neutrons and 27 protons
D. 27 neutrons and 33 protons
E. 0 neutrons and 0 protons

MAR

What is the unknown particle in the following nuclear reaction?

$^{239}_{92}U$ → particle + $^{239}_{93}Np$

A. alpha
B. beta
C. gamma
D. neutron
E. positron

MAR

What is the unknown particle in the following nuclear reaction?

$^{207}_{84}Po$ → particle + $^{207}_{83}Bi$

A. alpha
B. beta
C. gamma
D. neutron
E. positron

MAR

What new nucleus is produced in the following nuclear reaction?

$^{41}_{20}Ca + ^{0}_{-1}e \rightarrow$?

A. $^{41}_{21}Sc$
B. $^{42}_{19}K$
C. $^{41}_{19}K$

MAR

Which of the following nuclei has the highest binding energy per nucleon?

A. ^{7}Li
B. ^{59}Ni
C. ^{4}He
D. ^{232}Th
E. ^{0}Jq

MAR

Calculate the **binding energy per mol of nucleons (E_b) for carbon-12**. *Helpful values:*

- 1 proton = 1.007825 g/mol
- 1 neutron = 1.008665 g/mol
- carbon-12 = 12.000000 g/mol
- 2.998×10^8 m/s = c

A. 7.411×10^8 kJ
B. 8.893×10^9 kJ
C. 1.482×10^9 kJ
D. -0.098940 kJ
E. 0.0001×10^8 kJ

MAR

Radioactive iodine-131 is used to treat hyperthyroidism. It has a half-life of 8.04 days. If you begin with 8.8 micrograms, what mass remains after 32.3 days?

A. 4.4 micrograms
B. 2.2 micrograms
C. 1.1 micrograms
D. 0.54 micrograms
E. 0.23 micrograms

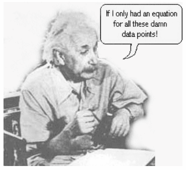
MAR

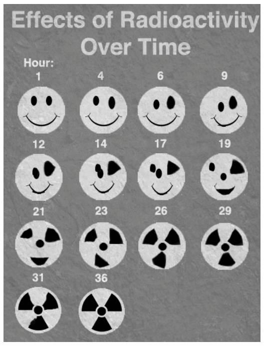
Gallium citrate, containing radioactive ^{67}Ga, is used as a tumor-seeking agent. It has a half-life of 78.2 hours. How long will it take for a sample of gallium citrate to decay to 15% of its original activity?

A. 86.5 hours
B. 157 hours
C. 214 hours
D. 235 hours
E. 150 seconds

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

Building Lewis Dot Structures

1) **Decide on the central atom**

   Central atom has lowest electron affinity
   Never Hydrogen! H only terminal atom

2) **Count Valence Electrons & Electron Pairs**

   # Valence Electrons = Group Number
   Electron Pairs = (Valence Electrons) / 2

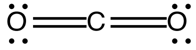
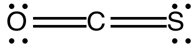
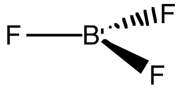
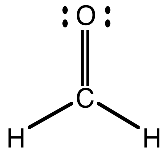
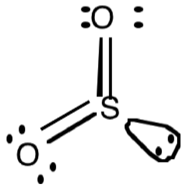
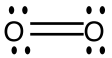
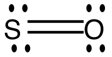
3) **Form sigma bond(s) between central atom and surrounding atom(s)**

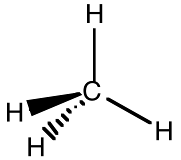
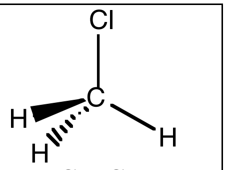
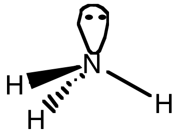
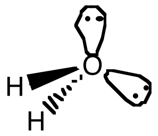
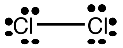
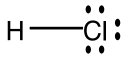
4) **Remaining electrons form lone pairs**

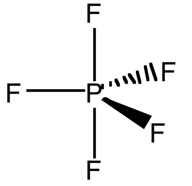
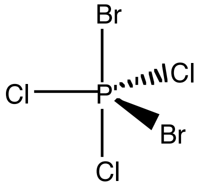
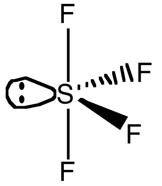
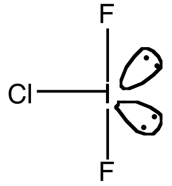
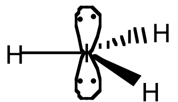
   Use the ***Octet Rule*** to assign electrons

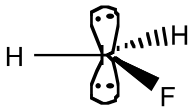
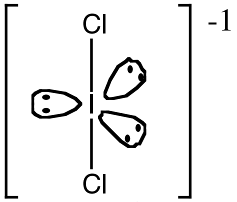
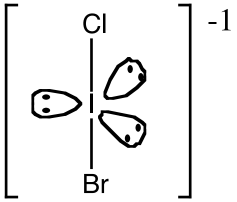
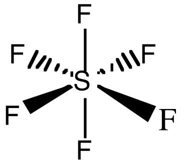
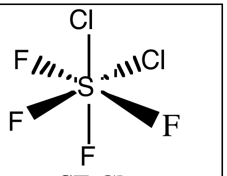
5) If an atom(s) does not have octet, **create pi bond(s) using lone pairs on adjacent atoms**

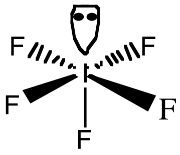
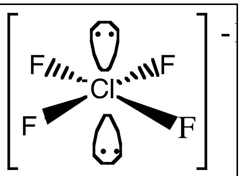
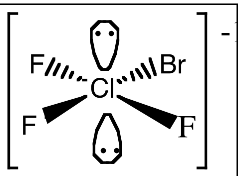
CH 222 Geometry & Polarity Guide

General Structure	VSEPR / Electron Pair Geometry	Molecular / Shape Geometry	Bond Angle(s) (Degrees)	Polarity	Example
<u>Two Clouds</u>: **2 atoms, no electron pairs,** ***symmetric***	**Linear**	**Linear**	**180**	**Non-polar**	 **CO_2**
<u>Two Clouds</u>: **2 atoms, no electron pairs,** ***asymmetric***	**Linear**	**Linear**	**180**	**Polar**	 **COS**
<u>Three Clouds</u>: **3 atoms** **no electron pairs** ***symmetric***	**Trigonal Planar**	**Trigonal Planar**	**120**	**Nonpolar**	 **BF_3**
<u>Three Clouds</u>: **3 atoms** **no electron pairs** ***asymmetric***	**Trigonal Planar**	**Trigonal Planar**	**120**	**Polar**	 **H_2CO**
<u>Three Clouds</u>: **2 atoms** **1 electron pair** ***asymmetric***	**Trigonal Planar**	**Bent** *or* **Angular**	**120**	**Polar**	 **SO_2**
<u>Three Clouds</u>: **1 atom, 2 electron pairs,** ***symmetric***	**Trigonal Planar**	**Linear**	**180**	**Non-polar**	 **O_2**
<u>Three Clouds</u>: **1 atom, 2 electron pairs,** ***asymmetric***	**Trigonal Planar**	**Linear**	**180**	**Polar**	 **SO**

General Structure	VSEPR / Electron Pair Geometry	Molecular / Shape Geometry	Bond Angle(s) (Degrees)	Polarity	Example
<u>Four Clouds</u>: **4 atoms** **no electron pairs** ***symmetric***	**Tetrahedral**	**Tetrahedral**	**109.5**	**Nonpolar**	 **CH_4**
<u>Four Clouds</u>: **4 atoms** **no electron pairs** ***asymmetric***	**Tetrahedral**	**Tetrahedral**	**109.5**	**Polar**	 **CH_3Cl**
<u>Four Clouds</u>: **3 atoms** **1 electron pair** ***asymmetric***	**Tetrahedral**	**Trigonal Pyramid**	**109.5**	**Polar**	 **NH_3**
<u>Four Clouds</u>: **2 atoms** **2 electron pairs** ***asymmetric***	**Tetrahedral**	**Bent *or* Angular**	**109.5**	**Polar**	 **H_2O**
<u>Four Clouds</u>: **1 atom** **3 electron pairs** ***symmetric***	**Tetrahedral**	**Linear**	**180**	**Nonpolar**	 **Cl_2**
<u>Four Clouds</u>: **1 atom** **3 electron pairs** ***asymmetric***	**Tetrahedral**	**Linear**	**180**	**Polar**	 **HCl**

General Structure	VSEPR / Electron Pair Geometry	Molecular / Shape Geometry	Bond Angle(s) (Degrees)	Polarity	Example
Five Clouds: **5 atoms** **no electron pairs** ***symmetric***	**Trigonal Bipyramid**	**Trigonal Bipyramid**	**90, 120, 180**	**Nonpolar**	 **PF_5**
Five Clouds: **5 atoms** **no electron pairs** ***asymmetric***	**Trigonal Bipyramid**	**Trigonal Bipyramid**	**90, 120, 180**	**Polar** *…but can be nonpolar in certain AX_2Y_3 configurations*	 **PCl_3Br_2**
Five Clouds: **4 atoms** **1 electron pair** ***asymmetric***	**Trigonal Bipyramid**	**Seesaw**	**90, 120, 180**	**Polar**	 **SF_4**
Five Clouds: **3 atoms** **2 electron pairs** ***asymmetric***	**Trigonal Bipyramid**	**T-shaped** *dominant trigonal bipyramid structure with 3 atoms*	**90, 180**	**Polar**	 **IF_2Cl**
Five Clouds: **3 atoms** **2 electron pairs** ***symmetric***	**Trigonal Bipyramid**	**Trigonal Planar** *rare*	**120**	**Nonpolar**	 **IH_3**

General Structure	VSEPR / Electron Pair Geometry	Molecular / Shape Geometry	Bond Angle(s) (Degrees)	Polarity	Example
<u>Five Clouds</u>: **3 atoms** **2 electron pairs** ***asymmetric***	**Trigonal Bipyramid**	**Trigonal Planar** *rare*	**120**	**Polar**	 **IH_2F**
<u>Five Clouds</u>: **2 atoms** **3 electron pairs,** ***symmetric***	**Trigonal Bipyramid**	**Linear**	**180**	**Nonpolar**	 **ICl_2^{-1}**
<u>Five Clouds</u>: **2 atoms** **3 electron pairs** ***asymmetric***	**Trigonal Bipyramid**	**Linear**	**180**	**Polar**	 **$IClBr^{-1}$**
<u>Six Clouds</u>: **6 atoms** **no electron pairs** ***symmetric***	**Octahedral**	**Octahedral**	**90, 180**	**Nonpolar**	 **SF_6**
<u>Six Clouds</u>: **6 atoms** **no electron pairs** ***asymmetric***	**Octahedral**	**Octahedral**	**90, 180**	**Polar** *…but can be nonpolar in certain AX_4Y_2 configurations*	 **SF_4Cl_2**

General Structure	VSEPR / Electron Pair Geometry	Molecular / Shape Geometry	Bond Angle(s) (Degrees)	Polarity	Example
<u>Six Clouds</u>: **5 atoms** **1 electron pair** ***asymmetric***	**Octahedral**	**Square Pyramid**	**90, 180**	**Polar**	 **IF_5**
<u>Six Clouds</u>: **4 atoms** **2 electron pairs** ***symmetric***	**Octahedral**	**Square Planar**	**90, 180**	**Nonpolar**	 **ClF_4^{-1}**
<u>Six Clouds</u>: **4 atoms** **2 electron pairs** ***asymmetric***	**Octahedral**	**Square Planar**	**90, 180**	**Polar**	 **ClF_3Br^{-1}**

CH 222 Guide to Polarity

Polarity arises when two atoms in a bond receive unequal distributions of electron density; i.e. one atom is slightly more negative than the other.

Example: In HCl, the Cl is more negative than the H

Polarity is an important property of molecules.
It affects physical properties such as melting point, boiling point and solubility.
Chemical properties also depend on polarity.
Dipole moment, μ, is a quantitative measure of the polarity of a molecule.

A molecule is **nonpolar** if the central atom is symmetrically substituted by identical atoms.

Examples: CO_2, CH_4 , CCl_4

A molecule will be **polar** if the geometry is not symmetrical.

Examples: H_2O, NH_3, CH_2Cl_2

The **degree of polarity** is a function of the *number* and *type* of polar bonds as well as the *geometry*.

For a molecule to be polar, the effects of bond polarity must not cancel out.

CH 222 Guide to Formal Charges

The **formal charge** for an atom in a molecule or ion is the charge calculated for that atom based on the Lewis structure of the molecule or ion using the following equation:

$$\textbf{Formal Charge} = \textbf{GN} - \textbf{lpe} - {}^{1}/_{2}(\textbf{bpe})$$

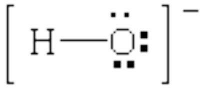
where **GN** = group number for the atom

lpe = number of lone pair electrons on the atom

bpe = number of bonding pair electrons on the atom

- The sum of the formal charges on the atoms in a molecule or ion always equals its net ionic charge.
- Formal charges can be helpful when deciding on the *most likely resonance structure*; the structure with the lowest overall formal charges will be the most likely resonance structure
- Alternatively, the most likely resonance structure will have the lowest absolute value of individual formal charges; i.e. the resonance structure with the lowest summation of the absolute values of *each* formal charge in the molecule or ion will be the most preferred structure

Example: Hydroxide, OH^-, has a net ionic charge of –1 and it has the Lewis structure:

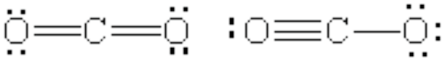


The hydrogen atom in hydroxide has a group number = 1, it has no lone pair electrons, and it has two bonding pair electrons in the single bond. Therefore, the formal charge on the hydrogen atom is: $1 - 0 - {}^{1}/_{2}\,(2) = \mathbf{0}$

The oxygen atom in hydroxide has a group number = 6, it has six lone pair electrons, and it has two bonding pair electrons in the single bond. Therefore, the formal charge on the oxygen atom is: $6 - 6 - {}^{1}/_{2}\,(2) = \mathbf{-1}$

The sum of formal charges is $0 + (-1) = \mathbf{-1}$, which is equal to the net ionic charge of -1.

Example: Carbon dioxide, CO_2, has a net ionic charge of zero. Two possible Lewis structures:



Structure A Structure B

The carbon atom in *both* **A** and **B** has a group number = 4, it has no lone pair electrons, and it has eight bonding pair electrons (in **A**, eight electrons come two double bonds; in **B**, two come from the single bond and six from the triple bond.). Therefore, the formal charge on the carbon atom is: $4 - 0 - {}^{1}/_{2}\,(8) = \mathbf{0}$

In structure **A**, *each* oxygen atom has a group number = 6, four lone pair electrons, and four bonding pair electrons in the double bond. Therefore, the formal charge on the oxygen atom is: $6 - 4 - {}^{1}/_{2}\,(4) = \mathbf{0}$

In structure **B**, the oxygen atom with the triple bond has a group number = 6, two lone pair electrons, and six bonding pair electrons from the triple bond. Therefore, the formal charge on this oxygen atom is: $6 - 2 - {}^{1}/_{2}\,(6) = \mathbf{+1}$

In structure **B**, the oxygen atom with the single bond has a group number = 6, six lone pair electrons, and two bonding pair electrons from the single bond. Therefore, the formal charge on this oxygen atom is: $6 - 6 - {}^{1}/_{2}\,(2) = \mathbf{-1}$

Note that the sum of the formal charges in both structures equals the ionic charge:

In **A**, $0 + 0 + 0 = 0$

In **B**, $0 + 1 - 1 = 0$

However, structure **A** is preferred over **B** because **A** has lower formal charges overall. Nature prefers resonance structures with lower formal charges.

The most preferred structure will have the lowest absolute value of individual formal charges:

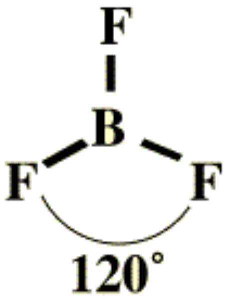
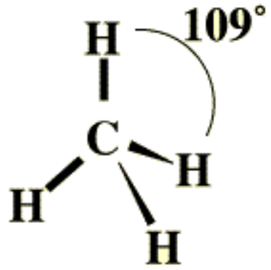
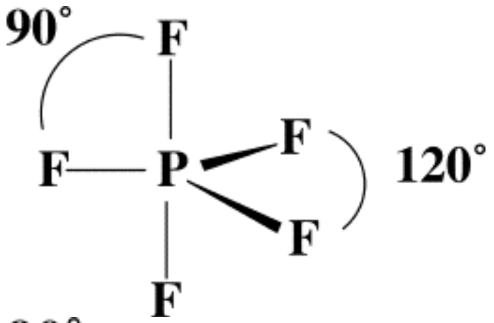
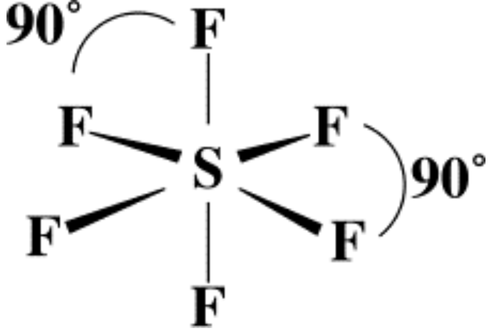
A: $|\,0\,| + |\,0\,| + |\,0\,| = 0$

B: $|\,0\,| + |\,1\,| + |\,-1\,| = 2$

CH 222 Chapter Seven Study Guide

- Be able to distinguish between valence and core electrons for any given element.
- Know the three primary attractive and repulsive forces that constitute atomic bonds.
- Know the key differences between covalent and ionic bonding. Be able to predict from a formula whether a compound is ionic or covalent.
- Be able to draw Lewis electron dot structures for a given compound.
- Understand the significance of the octet rule in assigning Lewis electron dot structures.
- Realize that certain elements (boron, etc.) do not always obey the octet rule.
- Understand the significance of resonance structures.
- Be able to define and predict trends in bond order, bond length and bond dissociation energy.
- Be able to use bond dissociation energy in calculations.
- Be able to recognize polar bonds in molecules using the concept of electronegativity.
- Be able to calculate the formal charge for a given atom in a molecule.
- Know the difference between formal charge and oxidation number.
- Be able to predict the most likely resonance structure using formal charges.
- Know how to apply VSEPR theory to predict the geometry on a molecule.
- Be able to predict if a molecule is polar or non-polar.
- Be able to solve and understand the assigned problems in problem set #1.

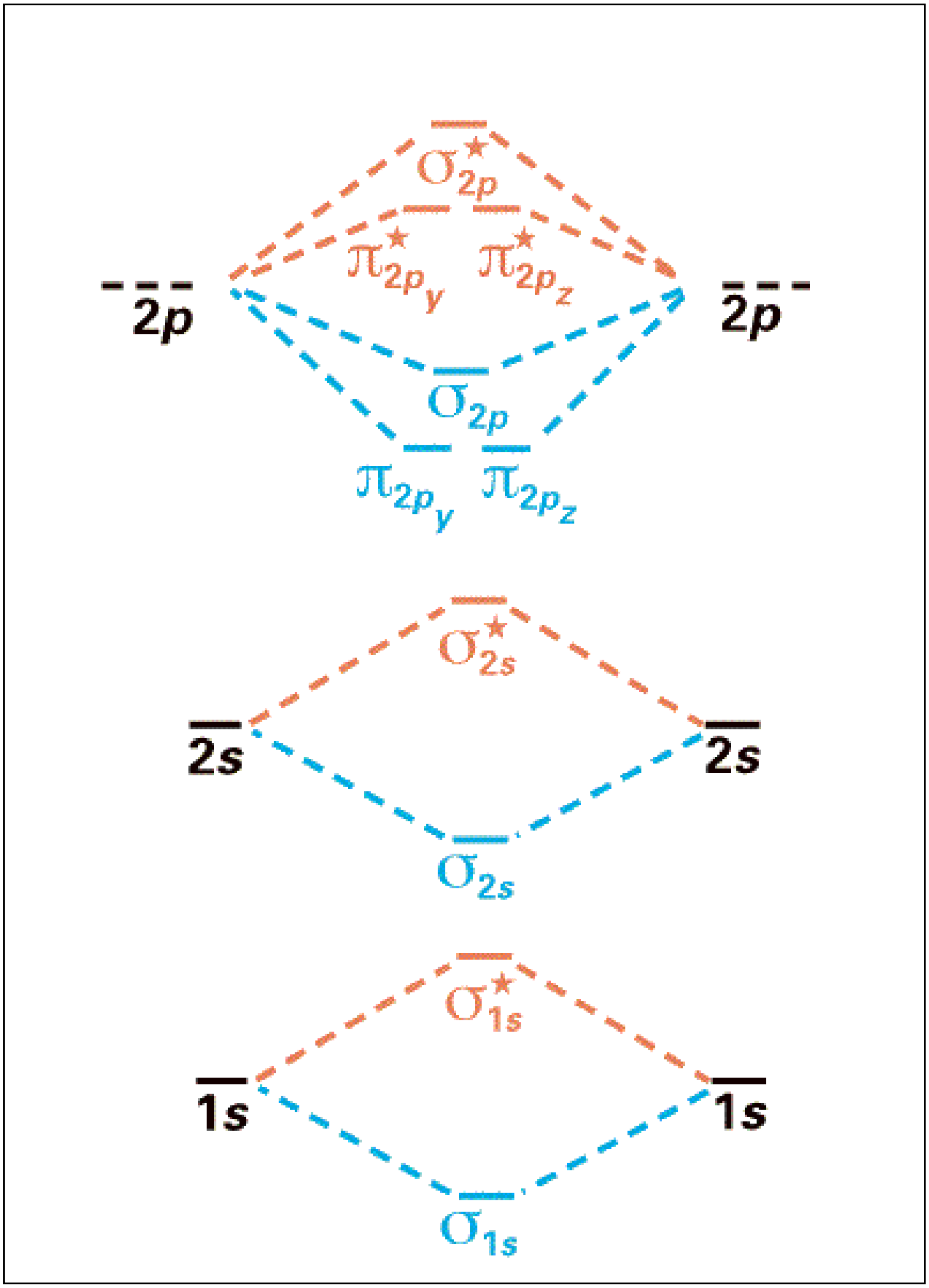
CH 222 VSEPR Guide

Number of Electron Clouds	*VSEPR Hybridization*	*Example*	*Electron Pair Geometry*
2	***sp***		*linear*
3	***sp^2***		*trigonal planar*
4	***sp^3***		*tetrahedral*
5	***dsp^3***		*trigonal bipyramid*
6	***d^{2}sp^3***		*octahedral*

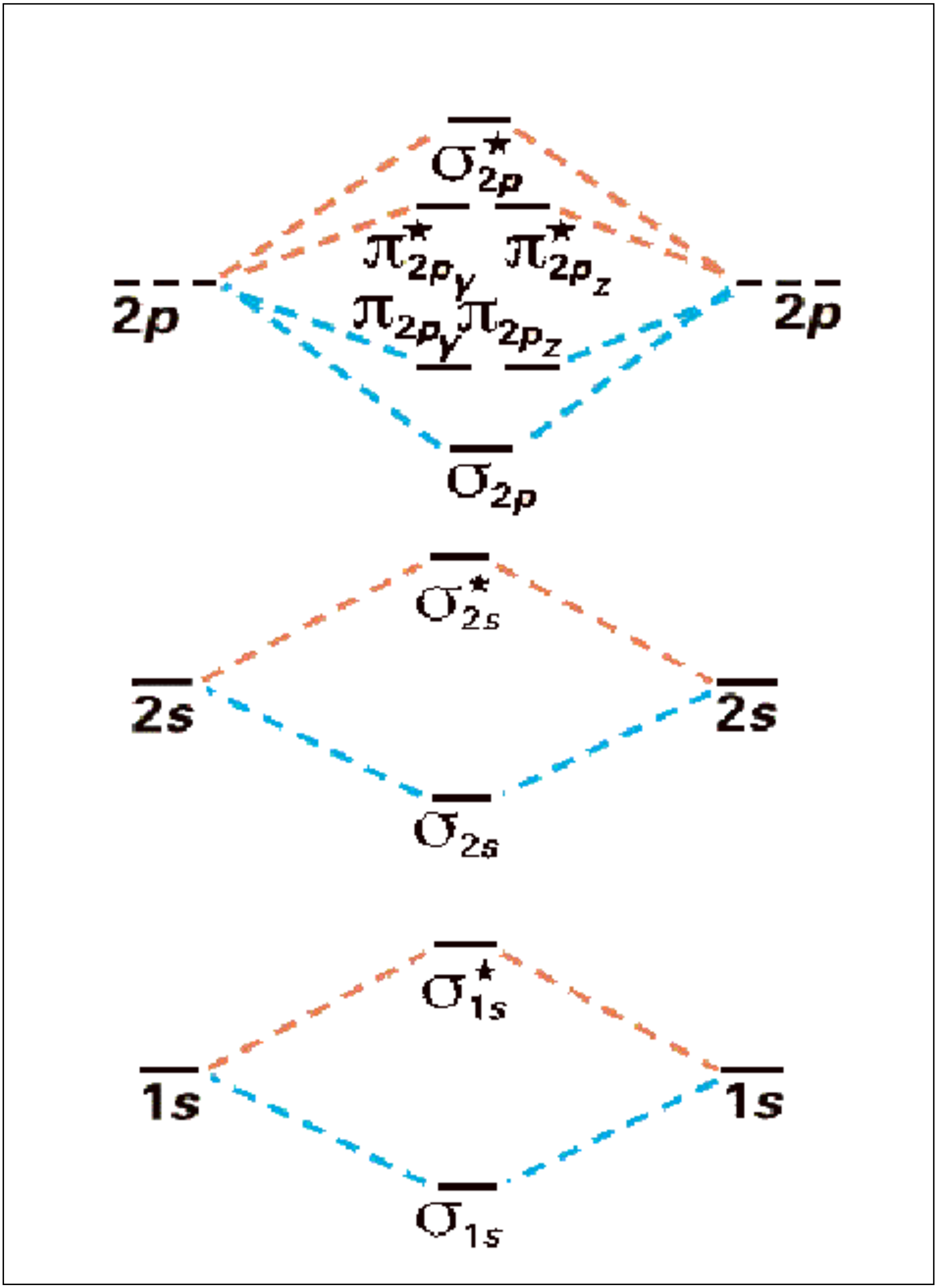
Four Principles of Molecular Orbital Theory

Principle #1:	The total number of molecular orbitals produced is always equal to the total number of atomic orbitals contributed by the atoms that have combined.
Principle #2:	The bonding molecular orbital is lower in energy than the parent orbital, and the antibonding orbital is higher in energy than the parent orbital.
Principle #3:	Electrons of the molecule are assigned to orbitals of successively higher energy.
Principle #4:	Atomic orbitals combine to form molecular orbitals most effectively when the atomic orbitals are of similar energy.

MO Diagram for B_2 Through N_2



MO Diagram for O_2 Through Ne_2



CH 222 Chapter Eight Study Guide

- Be able to describe the main features of valence bond theory. Realize that valence bond theory views bonding as arising from the idealized overlap between two atomic orbitals; the electron density remains localized along the bond axis, but the overall energy of the system does not change.

- Be able to describe the main features of molecular orbital theory. Realize that the number of molecular orbitals always equals the number of atomic orbitals used in the combining atoms. Know that some of the bonds will be bonding and others will be antibonding.

- Understand how a sigma bond forms. Know the three types of sigma bond formation (s+s, p+s, p+p).

- Understand the concept of pi bonds under both molecular orbital and valence bond theories.

- Be able to use the concept of orbital hybridization to describe the bonding form of atoms in molecules. Know how to assign sp, sp^2, sp^3, sp^3d, and sp^3d^2 hybridization to the appropriate atoms, and know what geometry each of these hybridization forms will give the molecule (ex: sp = linear, etc.)

- Be able to draw molecular orbital diagrams for all diatomic molecules (Li_2, O_2, etc.) up to neon. Know how to handle a charged diatomic molecule (ex: O_2^{2-}) using molecular orbital theory. Realize that you will be expected to determine the ***correct*** molecular orbital diagram for oxygen, fluorine and neon diatomics.

- Be able to apply Hund's Rule and the Pauli Exclusion Principle when assigning electrons to molecular orbitals.

- Be able to predict paramagnetic behavior and the bond order from molecular orbital diagrams. Know what these predictions mean for the molecule.

- Realize the implications of the application of molecular orbital theory to solids (band theory), especially for industrial processes.

- Be able to solve and understand the assigned problems in problem set #2.

Chemistry 222 Organic Chemistry Nomenclature Guide

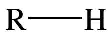
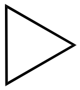
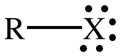
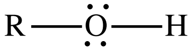
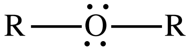
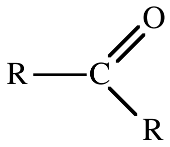
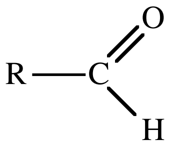
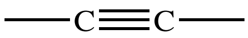
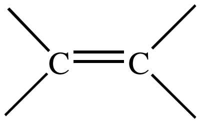
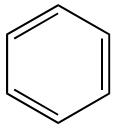
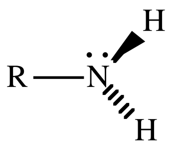
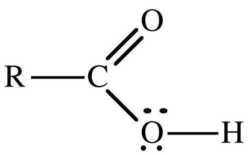
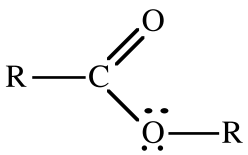
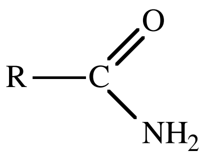
Many molecules in organic chemistry can be named using alkyl groups. *MEMORIZE THEM!*

Common Alkyl (R) Groups

Number of Carbons	Formula	Name
1	$-CH_3$	methyl
2	$-CH_2CH_3$	ethyl
3	$-CH_2CH_2CH_3$	propyl
4	$-CH_2CH_2CH_2CH_3$	butyl
5	$-CH_2CH_2CH_2CH_2CH_3$	pentyl
6	$-CH_2(CH_2)_4CH_3$	hexyl
7	$-CH_2(CH_2)_5CH_3$	heptyl
8	$-CH_2(CH_2)_6CH_3$	octyl
9	$-CH_2(CH_2)_7CH_3$	nonyl

Alkyl groups are generically referred to as R-groups, where **R** could be a methyl group, ethyl group, octyl group, etc.

Organic compounds are often lumped into families or classes of compounds. The classes we will study this term include the following:

 **Alkanes**	 **Cycloalkanes**	 **Alkyl Halides** *or* **haloalkanes**	 **Alcohols**	 **Ethers**
 **Ketones**	 **Aldehydes**	 **Alkynes**	 **Alkenes**	 **Aromatics**
 **Amines**	 **Carboxylic Acids**	 **Esters**	 **Amides**	***All of these families are detailed in the pages that follow.***

Alkanes

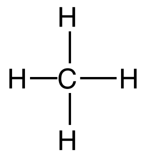
Elemental Formula: **C_nH_{2n+2}**

Nomenclature Guidelines: **-yl** on alkyl group, **+ane** to ending

Notes: An alkane is an alkyl group plus a hydrogen, often referenced as **R-H**. Alkanes contain only carbon and hydrogen atoms in long chains with no rings. Each carbon atom is *sp^3* hybridized. Alkanes make great fuels but are generally unreactive.

Example: **CH_4** - **methane** - is a *methyl* group plus a hydrogen (CH_3-H)

Example: **C_2H_6** - **ethane** - is a *ethyl* group plus a hydrogen (CH_3CH_2-H)



methane, CH_4

Cycloalkanes

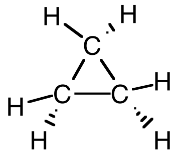
Elemental Formula: **C_nH_{2n}**

Nomenclature Guidelines: **cyclo+ -yl** on alkyl group, **+ane** to ending

Notes: Cycloalkanes are alkanes which form an internal ring within the molecule. Cycloalkanes have two less hydrogens than their straight chain alkane counterparts. Cycloalkanes are generally somewhat more reactive than alkanes.

Example: **C_3H_6** - **cyclopropane** - is a *propyl* group in a triangular ring

Example: **C_6H_{12}** - **cyclohexane** - is a *hexyl* group in a six sided ring



cyclopropane, C_3H_6

Alkyl Halides *or* Haloalkanes

Elemental Formula: **$C_nH_{2n+1}X$** *or* **R-X** *(where X = F, Cl, Br or I)*

Nomenclature Guidelines: **alkyl name + halide** (anionic halogen) *or* **haloalkane**

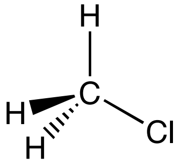
Notes: Alkyl halides are often reported using two nomenclature systems. The older *alkyl halide* system (such as *methyl chloride*) is in use, but IUPAC guidelines propose a gradual transition to *haloalkane* nomenclature (such as *chloromethane* instead of methyl chloride). The latter system is preferred due to its flexibility when dealing with multiple halogens (i.e. *1,2-dichloroethane*, etc.)

Example: **CH_3Cl – chloromethane** *or* **methyl chloride** – This is a *methyl* group plus a chloride

Example: **$CH_3CHI(CH_2)_4CH_3$ – 2-iodoheptane** *or* **2-heptyl iodide** - is a *heptyl* group plus an iodide on the *second* carbon of the heptyl group

Example: **CH_2Cl_2** - **dichloromethane**

Example: **$CH_2I\text{-}CHI\text{-}(CH_2)_4CIH_2$** - **1,2,7-triiodoheptane**



chloromethane, CH_3Cl

Alcohols

Elemental Formula: **$C_nH_{2n+2}O$** *or* **R-OH**

Nomenclature Guidelines: **-yl** on alkyl group, **+anol** to ending

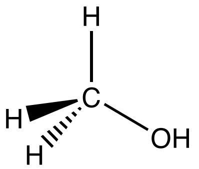
Notes: Alcohols are wonderful starting reagents for a variety of syntheses. If two –OH units are present, the compound is a *diol*, and if three –OH units are present, it is called a *triol*, etc.

Example: **CH_3OH** - **methanol** - is a *methyl* group plus an OH (CH_3-OH)

Example: **$CH_3(CH_2)_4CH_2OH$** - **hexan-1-ol** - OH on first carbon

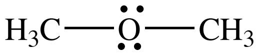
Example: **$CH_3CH(OH)(CH_2)_3CH_3$** - **hexan-2-ol** - OH on second carbon

Example: **$CH_2(OH)CH_2CH_2(OH)$** - **propane-1,3-diol**



methanol, CH_3OH

Ethers



dimethylether or methoxymethane

Elemental Formula: **R-O-R**

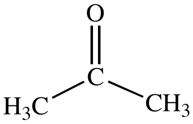
Common Nomenclature: **(alkyl name #1) + (alkyl name #2) + ether** Note that the alkyl names must be *alphabetical*; “ethyl methyl ether” and not “methyl ethyl ether”

IUPAC Nomenclature: **(shorter alkyl group - yl + oxy) + (longer alkyl group -yl +ane)** As an example, $CH_3OCH_2CH_3$ would be “methoxyethane” instead of “ethoxymethane” or “ethyl methyl ether"

Notes: Ethers have two common nomenclature systems, both of which appear often. The IUPAC system is more versatile, allowing for ether linkages which are difficult to name using the older "common" system.

Example: **CH_3OCH_3 - dimethyl ether** *or* **methyoxymethane**
Example: **$CH_3OCH_2CH_2CH_3$ – 1-methyl propyl ether** *or* **1-methoxypropane**
Example: **$CH_3CH(OCH_3)CH_2CH_3$ - 2-methoxybutane**

Ketones



propanone or acetone

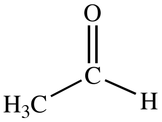
Elemental Formula: **R-CO-R** *where CO = carbonyl (C=O) group*

Nomenclature Guidelines: Count number of carbons; determine **alkane** name, then (# of CO position) **alkane -e +one**

Notes: Aldehydes and ketones are similar in structure, but ketones have two alkyl groups connected to the central carbonyl carbon.

Example: **CH_3COCH_3 - propanone** – this compound is close to *propane*; the CO is on the *second* carbon. Propanone is also known as **acetone**.
Example: **$CH_3COCH_2CH_2CH_3$ - 2-pentanone** - five carbons, *pentane*
Example: **$CH_3CH_2COCH_2CH_3$ - 3-pentanone**

Aldehydes



ethanal

Elemental Formula: **R-CO-H** *where CO = carbonyl (C=O) group*

Nomenclature Guidelines: Count number of carbons; determine **alkane** name, then **alkane -e +al**

Notes: Aldehydes and ketones are similar in structure, but aldehydes have at least one hydrogen connected to the central carbonyl carbon.

Example: **CH_3COH - ethanal** - two carbon atoms like *ethane*
Example: **$CH_3CH_2CH_2COH$ - butanal** - four carbon atoms like *butane*

Alkynes

2-butyne

Elemental Formula: **C_nH_{2n-2}** *or* **R-C≡C-R**

Nomenclature Guidelines: Count number of carbons; then **alkyl -yl +yne**

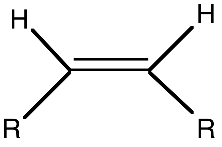
Example: **$CH_3C{\equiv}CH$ - propyne** - three carbons like *propane*
Example: **$CH_3C{\equiv}CCH_2CH_2CH_3$ - 2-hexyne** - six carbons, *hexane*
Example: **$CH_3CH_2CH_2C{\equiv}CH$ - 1-pentyne** – note that the number indicates *starting* position of triple bond, i.e. **2-pentyne** is **$CH_3CH_2C{\equiv}CCH_3$**

Alkenes

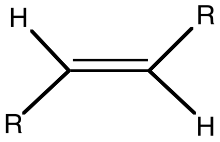
Elemental Formula: **C_nH_{2n}** *or* **R-(H)C=C(H)-R** *R can be H in these examples*

Nomenclature Guidelines: Count number of carbons; determine **alkane** name, then (# of C=C position) **alkane -ane +ene** *or (newer version):* **alkane -ane** (# of C=C position) **+ene**

Notes: Be sure to watch *cis* and *trans* configurations around a double bond where two of the groups off the double bonded carbons are the same.



cis



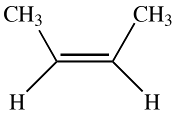
trans

Example: **$H_2C{=}CH_2$** - **ethene** - two carbons like *ethane*

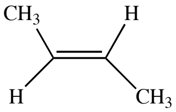
Example: **$H_2C{=}C(H)CH_3$** - **propene** - three carbons, *propane*

Example: **$(CH_3)(CH_3)C{=}CH_2$** - **2-methylpropene** – *longest chain has three carbons like propane*

Example: **$(CH_3)(H)C{=}C(CH_3)(H)$** - this is **2-butene (*or* but-2-ene),** but is it *cis* or *trans*? We cannot tell as written; it could be ***cis*-2-butene** (*cis-but-2-ene*) **or *trans*-2-butene** (*trans-but-2-ene*) (see below).



cis-2-butene



trans-2-butene

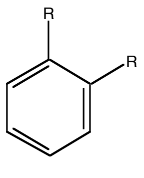
Aromatics

Elemental Formula: **$R\text{-}C_6H_5$** *where C_6H_5 = phenyl group and R = alkyl, halo, etc.*

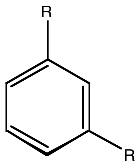
Nomenclature Guidelines: **alkylbenzene** (if **R** is alkyl) or **halobenzene** (if **R** is halo)

Notes: C_6H_5-H is **benzene**, an *aromatic* ringed compound with special stability. Substituents on the benzene ring must be *alphabetically* named.

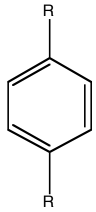
<u>Special Note</u>: Multiple alkyl or halo groups on a benzene ring can be named using appropriate numbers. Special cases involve *ortho* (1,2), *meta* (1,3) and *para* (1,4).



ortho (1,2)



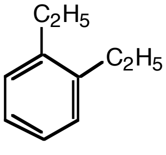
meta (1,3)



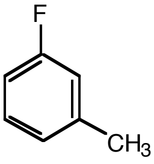
para (1,4)

Example: **$C_6H_5\text{-}CH_3$** - **methylbenzene** - *methyl* group on a benzene ring

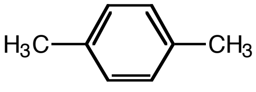
Example: **$C_6H_5\text{-}Cl$** - **chlorobenzene** - *chloro* group on a benzene ring



1,2-diethylbenzene or ortho-diethylbenzene

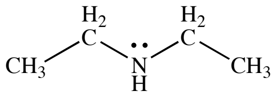


1,3-fluoromethylbenzene or meta-fluoromethylbenzene



1,4-dimethylbenzene or para-dimethylbenzene

Amines



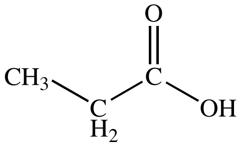
diethylamine

Elemental Formula: **R_x-NH_y** *where x + y = 3*

Nomenclature Guidelines: Count number of carbons; determine **alkane** name, then **alkane -ane +yl +amine**

Example: **CH_3NH_2** - **methylamine** - a *primary* amine (2 hydrogens on N)
Example: **$(CH_3)_2NH$** - **dimethylamine** - a *secondary* amine (only one hydrogen on N)
Example: **$(CH_3)_3N$** - **trimethylamine** - a *tertiary* amine (no hydrogens on N)
Example: **$(CH_3CH_2)(CH_3)NH$** - **ethylmethylamine**
Example: **$CH_3CH_2CH_2NH_2$ – 1-propylamine** *(choice of where N atom is connected to alkyl group)*

Carboxylic Acids



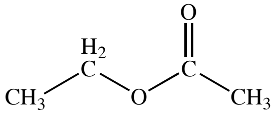
propanoic acid

Elemental Formula: **R-C(=O)OH** (R can be H)

Nomenclature Guidelines: Count number of carbons; determine **alkane** name, then **alkane -e +"oic acid"**

Example: **CH_3COOH** - **ethanoic acid** - also known as *acetic acid*
Example: **HCOOH** - **methanoic acid** - also known as *formic acid*
Example: **$CH_3(CH_2)_3COOH$** - **pentanoic acid** - also known as *valeric acid*

Esters



ethyl acetate

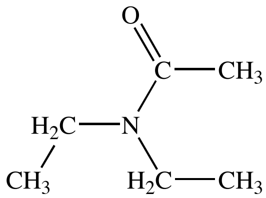
Elemental Formula: **R_1-C(=O)O-R_2**

Nomenclature Guidelines: Count number of carbons in *each R group*; determine **alkane** name *for each R*, then **{alkane(R_2) -ane +yl} {alkane(R_1) -e +oate}**

Note: if R_1 = CH_3, name it "acetate" over "ethanoate"

Example: **$CH_3(CH_2)_2COO(CH_2)_3CH_3$** - **butyl butanoate**
Example: **$CH_3COO(CH_2)_3CH_3$** - **butyl acetate**

Amides



N-diethylacetamide

Elemental Formula: **R_1-C(=O)-$NR_xH_{(2-x)}$** *(an **amine** on an **R-(C=O)** group)*

Nomenclature Guidelines: Count number of carbons; determine **alkane** name, then ***N*- alkane(R_x) -ane +yl +alkane(R_1) -ane +amide** *Note:* if R_1 = CH_3, name it "acet" over "meth"

Example: **$CH_3CONHCH_3$** - ***N*-methylacetamide**
Example: **$C_6H_5CON(CH_3)_2$** - ***N*-dimethylbenzamide**

Common Organic Compound Names

Many organic compounds have common names that do not follow the guidelines found in this handout. Here are some of the more common examples you might encounter at MHCC.

Structural Formula	"Official" Name	"Common" Name
$CH_3CH_2OCH_2CH_3$	diethyl ether	ether
CH_3COCH_3	propanone	acetone
$HCOH$	methanal	formaldehyde
CH_3COH	ethanal	acetaldehyde
$H_2C{=}CH_2$	ethene	ethylene
$H_2C{=}C(H)CH_3$	propene	propylene
$HC{\equiv}CH$	ethyne	acetylene
CH_3COOH	ethanoic acid	acetic acid
$C_6H_5\text{-}CH_3$	methylbenzene	toluene

-=-

Many more classes of organic compounds exist,
but you'll have to wait until CH 241 to see the rest of them! ☺

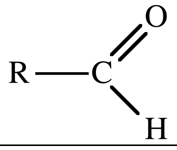
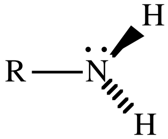
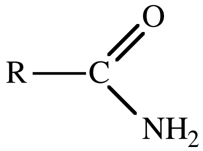
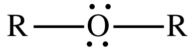
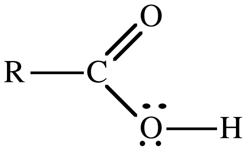
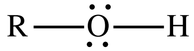
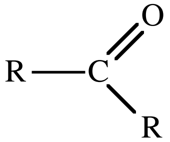
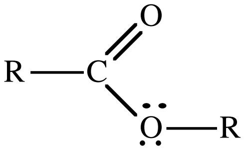
CH 222 "q&d" Guide to Organic Oxygen-Containing Classes

"q&d" = "Quick 'n' Dirty"

Remember: **"Triple As Eat-Cake"**

or

AAA E-CAKE

A	**Aldehyde**	$R-C(=O)-H$	
A	**Amine**	$R-NH_2$	
A	**Amide**	$R-C(=O)-NH_2$	
E	**Ether**	$R-O-R$	
C	**Carboxylic Acid**	$R-C(=O)-OH$	
A	**Alcohol**	$R-O-H$	
K	**Ketone**	$R-C(=O)-R$	
E	**Ester**	$R-C(=O)-O-R$	

CH 222 Chapter Twenty Study Guide

- Be able to identify and name representative examples from the following organic classes:

    * alkanes
    * cycloalkanes
    * alkyl halides
    * alkenes
    * alkynes
    * aromatic compounds (especially benzene and toluene)
    * alcohols
    * ethers
    * aldehydes
    * ketones
    * amines
    * carboxylic acids

- Know the systematic name for a compound matching one of these functional groups.

- Be able to draw structural formulas for unbranched and branched-chain alkanes as well as cyclohexanes. Be able to name these compounds.

- Understand the bonding behavior of alkenes, alkynes, aromatic compounds, aldehydes and ketones.

- Be able to identify possible isomers given a formula.

- Know the general reaction behavior for addition, elimination and substitution organic reactions.

- Be able to draw and name various *cis* and *trans* stereoisomers of alkenes.

- Recognize the importance of the carbonyl group in organic chemistry. Aldehydes, ketones, carboxylic acids, esters, etc. all contain a carbonyl group.

- Be able to name and write the structural formula of common polymers. Know how to write equations for the formation of addition polymers and condensation polymers.

- Be able to solve and understand the assigned problems in problem set #3.

Principal Assumptions of the Kinetic Molecular Theory

1. Gases consist of tiny (submicroscopic) molecules.

2. The distance between molecules is large compared with the size of the molecules themselves. The volume occupied by a gas consists mostly of empty space.

3. Gas molecules have no attraction for one another.

4. Gas molecules move in straight lines in all directions, colliding frequently with one another and with the walls of the container.

5. No energy is lost by the collision of a gas molecule with another gas molecule or with the walls of the container. All collisions are perfectly elastic.

6. The average kinetic energy for molecules is the same for all gases at the same temperature, and its value is directly proportional to the Kelvin temperature.

Explanation of Physical Properties and Gas Laws Based on the Kinetic Molecular Theory

1	Gases can be compressed easily	**1**	Because gas particles are far apart, they easily can be squeezed closer together by an outside force.
2	Gases expand to fill the volume of their container	**2**	Gas particles are constantly moving with no attractive forces between particles, so they will expand until they meet an outside force, namely, the wall of the container.
3	Gases have a low density	**3**	Because g gas is mostly empty space, there are few particles (low mass) per unit volume.
4	Gases can diffuse through each other	**4**	Gas particles are constantly moving and are separated by large distances. This leads to freedom for particles of one gas to move through the empty space of another gas.
5	Gases can exert a pressure on container walls	**5**	Moving gas particles collide with container walls, thus exerting a force on every square inch.
6	Boyle's law: $V \propto P^{-1}$	**6**	When the volume of gas is *decreased*, the particles collide with the walls more often, leading to a *greater pressure*. When volume of gas is *increased*, particles collide less often, leading to a *decreased pressure*.
7	Charles' law: $V \propto T$	**7**	Increased temperature causes particles to move faster, leading to more and "harder" collisions with walls. Pressure inside the walls is increased until the volume expands to the point where the pressure inside the walls is again equal to the pressure outside.
8	Dalton's Law of partial pressures: $P_{Total} = P_1 + P_2 + P_3 + P_4 + \ldots$	**8**	Since the particles move independently of one another, each gas in a mixture will exert a pressure independent of the pressure of the other gases. The total pressure will be the sum of the individual pressures.

CH 222 Chapter Nine Study Guide

- Be able to describe how pressure measurements are made.

- Be comfortable with the units of pressure, especially atm, mm Hg, torr, barr and Pa.

- Know what STP refers to and know its significance.

- Understand the Ideal Gas Law (PV = nRT); know when it applies and when it does not.

- Realize the importance of pressure, volume, temperature, number of moles, density, etc., especially when they are related to each other. Example: if pressure increases, temperature increases and volume decreases, etc.

- Know how to use the ideal gas law to solve problems similar to those in the problem set for this chapter.

- Be able to calculate the molecular weight (or molar mass) or the density of a gas knowing pressure, temperature, etc.

- MEMORIZE the value for the gas law constant, $R = 0.082057\ \text{L atm mol}^{-1}\ \text{K}^{-1}$. (Note that we will be using a different value for R in the near future - same R but expressed in different units.)

- Be able to use Dalton's Law of Partial Pressures to solve for individual contributions to the pressure of the system. Understand the importance of mole fractions when solving for Dalton's Law contributions.

- Be able to explain the Kinetic Molecular Theory (KMT) and its major assumptions. Know when these assumptions can be limiting.

- Understand the phenomena of diffusion and effusion. Be able to solve Graham's Law to find the relative rates of the gases.

- Realize that gases do not often obey the ideal gas laws due to "realistic" volume and intermolecular forces; instead, they must be described using the van der Waals equation. Know the effect of going from "ideal" gases to "real" gases using the van der Waals equation.

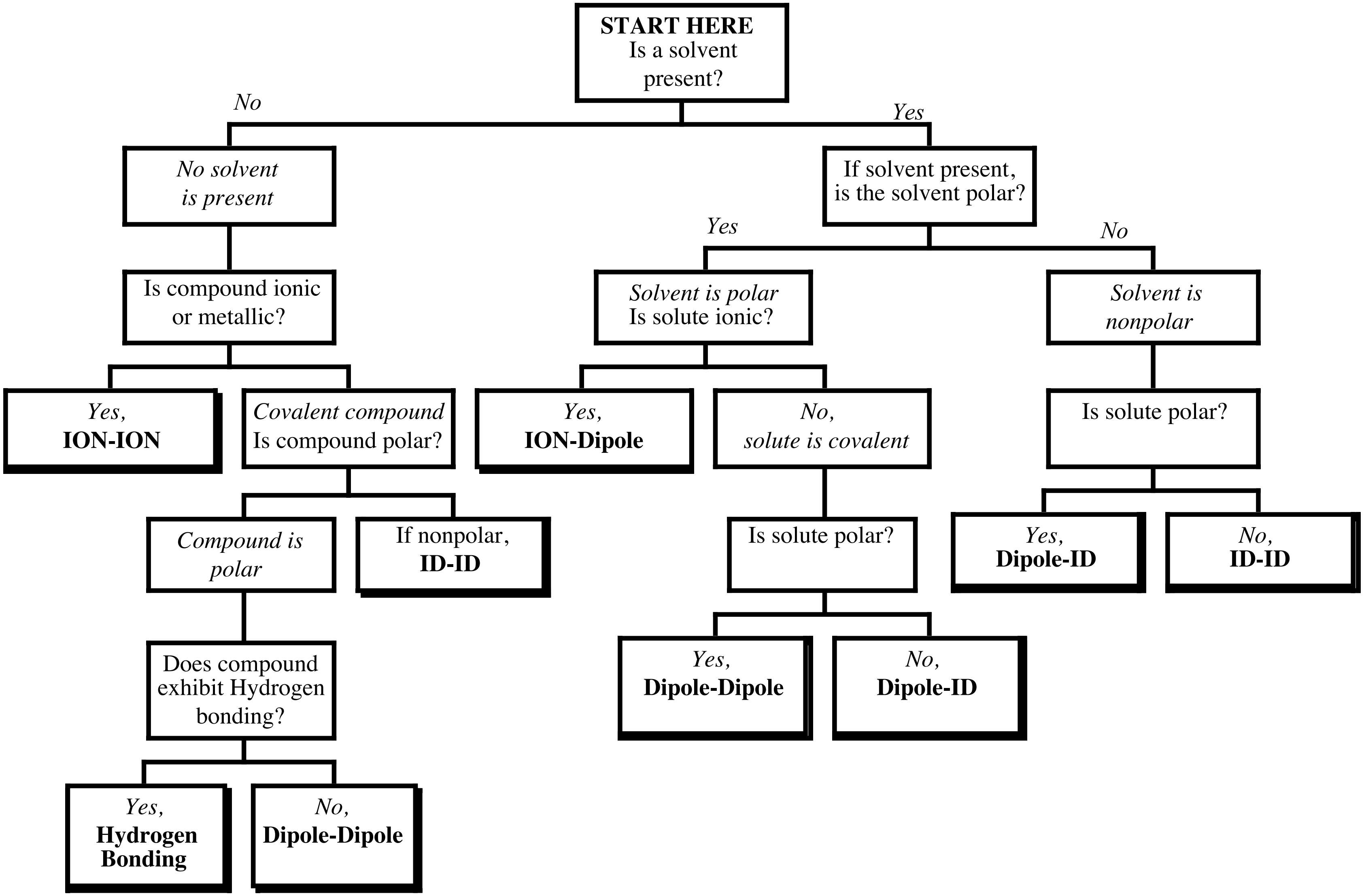
- Be able to solve and understand the assigned problems in problem set #4.

CH 222 Guide to Intermolecular Forces

There are six types of Intermolecular Forces (in *roughly* the order of strongest to weakest) are:

The **Ion-Ion** Force (which is the same for us as the *metallic* force), the **Ion-Dipole** Force, **Hydrogen Bonding** Force, the **Dipole - Dipole** Force, **Dipole - Induced Dipole** Force (or Dipole - ID), and the **Induced Dipole - Induced Dipole** Force (or ID - ID, also known as the *London Dispersion* force).

The flowchart below can help you decide which intermolecular force acts upon a particular substance.



CH 222 Guide to Solubility

I. Like Polarities Dissolve (or "Like Dissolves Like")

- **Polar solutes dissolve in polar solvents**
  NaCl (ionic) dissolves in water (polar)
  $KMnO_4$ *(ionic) dissolves in water (polar)*
  CH_3CH_2OH *(polar) dissolves in water (but see III, below)*

- **Non-polar solutes dissolve in non-polar solvents**
  Oil (non-polar) dissolves in gasoline (non-polar)
  Benzene (non-polar) dissolves in toluene (non-polar)

II. Polar and Non-Polar Species Do Not Dissolve

- **Polar solutes are insoluble in non-polar solvents**
  NaCl (ionic) is insoluble in gasoline (non-polar)
  Na_2CrO_4 *(ionic) is insoluble in benzene (non-polar)*

- **Non-polar solutes are insoluble in polar solvents**
  Oil (non-polar) is insoluble in water (polar)
  Toluene (non-polar) is insoluble in water (polar)

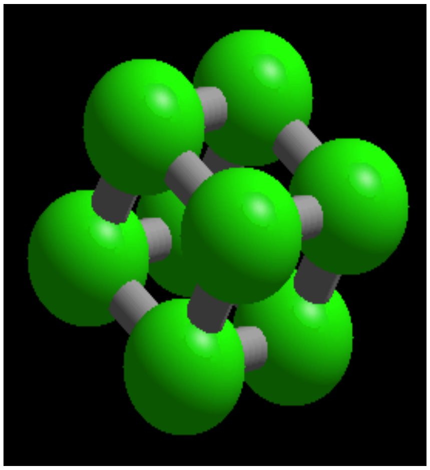
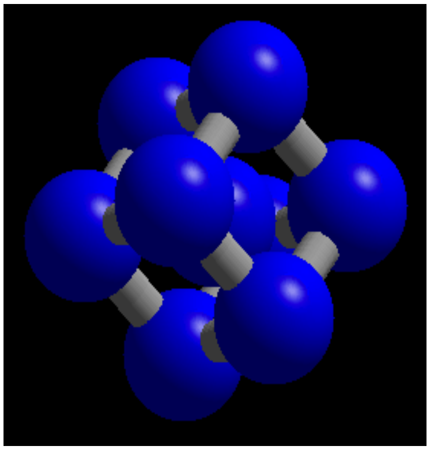
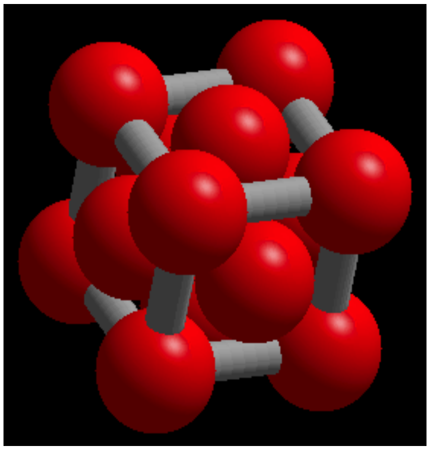
III. Organic Compounds, Water and Solubility

For organic compounds capable of hydrogen bonding (i.e. with a nitrogen or oxygen atom)

1 - 3 carbon atoms	water soluble	*ex: ethanol*
4 - 5 carbon atoms	borderline water solubility	*ex: n-butanol*
6 or more carbon atoms	water insoluble	*ex: n-hexanol*

Note: We will be discussing solubility in more detail during CH 223

"Quick and Dirty" Cubic Unit Cells for Metallic Solids Handout

Type	Atoms per Unit Cell	Edge - Radius Equation	Picture
Simple Cubic (SC)	1	Edge = 2*radius	
Body Centered Cubic (BCC)	2	$\text{Edge} = \frac{4 * \text{radius}}{\sqrt{3}}$	
Face Centered Cubic (FCC)	4	$\text{Edge} = \frac{4 * \text{radius}}{\sqrt{2}}$	

Volume = $(\text{edge})^3$ for simple cubic unit cells
1 pm = 10^{-12} m

CH 222 Chapter Ten Study Guide

- Be able to explain the Kinetic Molecular Theory (KMT) when describing the differences between solids, liquids and gases.

- Know the different intermolecular forces (dipole-dipole, induced dipole-induced dipole, etc.) in liquids and solids. Know their relative magnitudes - which one is strongest, which is weakest, etc.

- Be able to describe the major intermolecular forces acting upon a molecule or atom or a group of molecules or atoms.

- Know the requirements for hydrogen bonding - a hydrogen atom bonded to oxygen, nitrogen or fluorine.

- Be able to explain the process of evaporation and condensation of a liquid or its vapor.

- Be able to use the enthalpy of vaporization in calculations.

- Define and use the concept of the equilibrium vapor pressure of a liquid and its relation to the boiling point of a liquid. Know what is meant by normal boiling point.

- Know how to utilize the following concepts: cohesive forces, adhesive forces, surface tension, and viscosity.

- Be able to characterize different types of solids: metallic, ionic, molecular, network and amorphous.

- Be able to describe the three types of cubic units cells: simple cubic (or primitive), body centered cubic and face-centered cubic. Metals can utilize all three, but only the sc and fcc arrangements can occur for ionic compounds.

- Be comfortable with deriving the formula of an ionic compound from its unit cell. Know how many net atoms can exist in a given cubic cell.

- Be able to define the enthalpy of fusion and be able to use this in a calculation.

- Be able to identify the different points and regions of a phase diagram.

- Be able to solve and understand the assigned problems in problem set #4 & #5.

A Comparison of Various Concentration Units

Name	Units	Advantages	Disadvantages
Molarity (M)	mol solute / L solution	Useful in stoichiometry; measure by volume	Temperature dependent; must know density to find solvent mass
Molality (m)	mol solute / kg solvent	Temperature independent; useful in special applications *Examples:* Boiling point elevation: $\Delta T = K_{bp}m$ Freezing point depression: $\Delta T = K_{fp}m$	Measure by mass; must know density to convert to molarity
Mole Fraction (χ)	*none*	Temperature independent; useful for special applications *Example:* Raoult's Law: $P_{solvent} = \chi_{solvent}P^{\circ}_{solvent}$	Measure by mass; must know density to convert to molarity
Weight Percent	*none* (%)	Temperature independent; useful for small amounts	Measure by mass; must know density to convert to molarity

CH 222 Chapter Eleven Study Guide

- Define the terms solution, solvent, solute, colligative properties, miscible and immiscible.
- Be able to use molarity, molality, weight percent, mole fraction and parts per million (ppm) interchangeably.
- Understand the difference between saturated, unsaturated and supersaturated solutions.
- Be able to use Henry's Law and Raoult's Law.
- Be able to use colligative properties to find the molar mass (molecular weight) of a solute. Also be able to use colligative properties to find the elevation in boiling point or depression of freezing point. Be able to describe these phenomena using the models discussed in class.
- Know how to use the van't Hoff *i* factor in regards to colligative properties.
- Give a molecular-level explanation for osmosis.
- Be able to use osmotic pressure to determine concentrations and/or molecular weights of solutes.
- Be aware of the many enthalpy terms: lattice energy, energy of hydration, enthalpy of solution, energy of formation, etc.
- Know the effect of pressure and temperature on solubility. Pay special attention to Le Chatelier's Principle - this is an important concept.
- Recognize the differences between a homogeneous solution, a suspension and a colloid (or colloidal dispersion).
- Be familiar with the concepts of colloid and surfactant. Be able to characterize hydrophilic and hydrophobic substances.
- Be able to solve and understand the assigned problems in problem set #5.

CH 222 Quick & Dirty Kinetics Cheat Sheet

For the general reaction: **aA → Products**

	Zero Order	**First Order**	**Second Order**
Rate Law	Rate = k	Rate = $k[A]$	Rate = $k[A]^2$
Integrated Rate Law	$[A] = -kt + [A]_0$	$\ln [A] = -kt + \ln [A]_0$	$[A]^{-1} = kt + [A]_0^{-1}$
Plot Needed For Straight Line	[A] versus t	ln [A] versus t	$[A]^{-1}$ versus t
Relationship of Rate Constant to the Slope of Straight Line	Slope = $-k$	Slope = $-k$	Slope = k
Half-life	$t_{1/2} = [A]_0/2k$	$t_{1/2} = 0.693/k$	$t_{1/2} = 1/(k[A]_0)$

CH 222 "q&d" Guide to Reaction Mechanisms

"q&d" = "Quick 'n' Dirty"

- A **reaction mechanism** is the sequence of steps a reaction goes through at a molecular level in proceeding from reactants to products that control the speed and outcome of a reaction.

- Each step is called an **elementary step**.

- Each elementary step is called **unimolecular**, **bimolecular** or **termolecular** based upon the number of molecules colliding.

- A **reaction intermediate** is a species that is produced in one step of a reaction mechanism and completely consumed in a later step.

- The **rate law** for an elementary step does follow the balanced equation for that step.

- The slowest elementary step in a multi-step reaction mechanism is called the **rate-determining step** (RDS).

- The **sum** of the elementary steps in a multi-step reaction must give the balanced overall chemical equation**.**

- The mechanism must agree with the experimentally determined rate law.

Example: The following reaction has been studied at 500 K:

$$NO_{2(g)} + CO_{(g)} \rightarrow NO_{(g)} + CO_{2(g)}$$

An experimental rate law at 500 K has been found:

$$\text{rate} = k[NO_2]^2$$

and the following reaction mechanism proposed:

$$\begin{array}{ll} 2\,NO_{2(g)} \rightarrow NO_{3(g)} + NO_{(g)} & (1) \\ NO_{3(g)} + CO_{(g)} \rightarrow NO_{2(g)} + CO_{2(g)} & (2) \end{array}$$

How many elementary steps are present? Is this a valid mechanism? Are there any reaction intermediates? Which step is the rate-determining step? What is the molecularity of the rate-determining step?

Solution: There are **two** elementary steps in this mechanism.
To test validity, add the elementary steps together:

$$\begin{array}{ll} \cancel{2}\,NO_{2(g)} \rightarrow \cancel{NO}_{3(g)} + NO_{(g)} & (1) \\ \cancel{NO}_{3(g)} + CO_{(g)} \rightarrow \cancel{NO}_{2(g)} + CO_{2(g)} & (2) \\ \hline NO_{2(g)} + CO_{(g)} \rightarrow NO_{(g)} + CO_{2(g)} & (3) \end{array}$$

Reaction (3) is the same as the overall equation, so this mechanism is valid. $NO_{3(g)}$ and one molecule of $NO_{2(g)}$ are reaction intermediates. Since the rate depends on the squared concentration of $NO_{2(g)}$, elementary step (1) is the rate-determining step.

The molecularity of the rate-determining step is bimolecular. The molecularity of step (2) is also bimolecular.

CH 222 Chapter Twelve Study Guide

- Be able to explain the concept of reaction rate.

- Be able to derive the average and instantaneous rate from experimental information.

- Know the possible conditions that might affect reaction rate (such as reactant concentrations, temperature, presence of a catalyst, etc.)

- Know how to use a rate equation, and know how to find the rate constant and order of reaction from such an equation. Be able to derive a rate equation from experimental data.

- **MEMORIZE** the relationships between reactant concentration and time for zero order, first order and second order reactions. Know the integrated rate laws for these relationships.

- Know how to apply graphical techniques for determining the rate or reaction, reaction order and the rate constant.

- Be able to use the idea of half-life in kinetic equations.

- Know the principles behind the collision theory of reaction rates.

- Memorize the Arrhenius equation; be able to calculate the activation energy from experimental data.

- Be able to comprehend the concept of the reaction mechanism. Know how the mechanism relates to the overall stoichiometric equation for a reaction.

- Be able to describe the elementary steps of a mechanism and give their molecularity.

- Know how to find the rate-determining step in a mechanism and describe any reaction intermediates.

- Describe the role of the catalyst in reactions; know how it affects the activation energy and mechanism of a reaction. Understand the difference between a homogeneous and heterogeneous catalyst.

- Be able to solve and understand the assigned problems in problem set #6.

CH 222 "q&d" Guide to Nuclear Chemistry

"q&d" = "Quick 'n' Dirty"

$$ {}^{A}_{Z}\mathrm{X}^{n+/-} $$

X = element symbol
n = element charge (if any)
Z = atomic number (number of protons)
A = mass number (number of protons + neutrons)

Types of Radiative Processes

Alpha Decay:	Lose ${}^{4}_{2}\mathrm{He}$	*Example:*	${}^{234}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{230}_{90}\mathrm{Th}$	*Note 1*
Beta Decay:	Lose ${}^{0}_{-1}\mathrm{e}$	*Example:*	${}^{235}_{92}\mathrm{U} \rightarrow {}^{0}_{-1}\mathrm{e} + {}^{235}_{93}\mathrm{Np}$	*Note 4*
Gamma Decay:	Emit Energy	*Example:*	${}^{99m}_{43}\mathrm{Tc} \rightarrow \gamma + {}^{99}_{43}\mathrm{Tc}$	*Note 2*
Positron Emission:	Lose ${}^{0}_{+1}\mathrm{e}$	*Example:*	${}^{207}_{84}\mathrm{Po} \rightarrow {}^{0}_{+1}\mathrm{e} + {}^{207}_{83}\mathrm{Bi}$	*Notes 3, 4*
Electron Capture:	Gain ${}^{0}_{-1}\mathrm{e}$	*Example:*	${}^{7}_{4}\mathrm{Be} + {}^{0}_{-1}\mathrm{e} \rightarrow {}^{7}_{3}\mathrm{Li}$	*Note 4*
Neutron Capture:	Gain ${}^{1}_{0}\mathrm{n}$	*Example:*	${}^{6}_{3}\mathrm{Li} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{3}_{1}\mathrm{H}$	

Note 1: The alpha particle is actually charged, having a charge of +2. This makes the other product have a negative -2 charge (conservation of charge), but the charges of the ions are rarely considered in nuclear chemistry.

Note 2: Gamma emissions have energies in the range of roughly 1 MeV ($1.6 * 10^{-13}$ J).

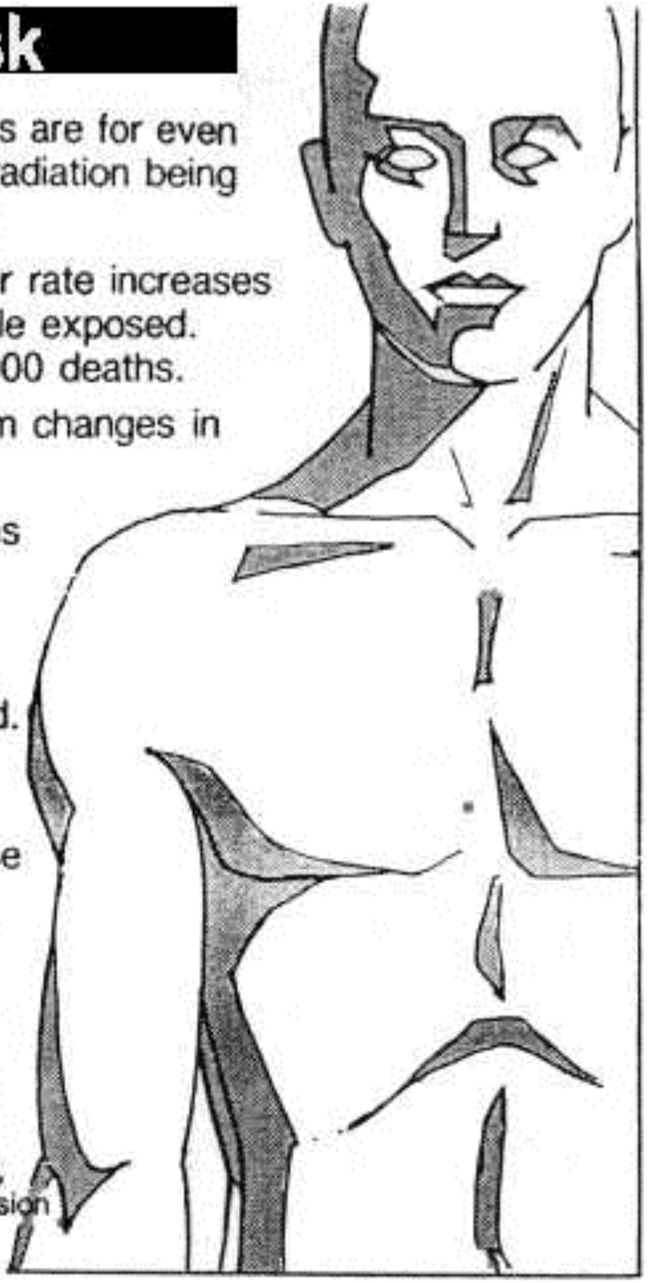
Note 3: A positron is an *antielectron* (a particle of antimatter) - when a positron and an electron collide, they annihilate each other (${}^{0}_{+1}\mathrm{e} + {}^{0}_{-1}\mathrm{e} \rightarrow 2\gamma$).

Note 4: The beta decay process produces an antineutrino in addition to the other products, while the positron emission and electron capture processes result in the creation of a neutrino. This is due to the conservation of spin concept, but you need not concern yourself about neutrinos in *this* CH 222 class!

Radiation: Exposure vs. health risk

REM [Roentgen Equivalent Man] is standard unit of radiation. Doses are for even exposure over entire body. Authorities are unable to estimate the radiation being released in Soviet Union tragedy.

- **.01 REM:** Dental X-ray.
- **.03 REM:** Chest X-ray.
- **0.14-.2 REM:** Average annual dose for Chicago area residents from environment and medical sources.
- **0.5 REM:** Federal exposure limit for pregnant workers in nuclear industry.
- **0.3-.6 REM:** Mammogram.
- **1 REM:** Spinal X-ray.
- **1-5 REM:** Public protection needed if nuclear accident is anticipated at this size.
- **5 REM:** Annual exposure limi for most nuclear workers.
- **10 REM:** 30-year fatal cancer rate increases by about 0.7% per million people exposed. Without exposure, rate is 170,000 deaths.
- **10-25 REM:** Small, short-term changes in blood of the exposed.
- **25-50 REM:** Almost all victims show low white- and red-blood cell counts within 24 hours.
- **50-100 REM:** Nausea within 12 hours for 5% of the exposed.
- **100-200 REM:** 50% will experience vomiting within 3 hours of exposure; 10% will lose hair within 5 to 10 days.
- **300-400 REM:** Death within 60 days for 50%.
- **500-600 REM:** Death within 60 days for 90%.



Chicago Tribune Graphic; Sources: University of Chicago, American Nuclear Society, Argonne Laboratories, Rensselaer Polytechnic Institute, Nuclear Regulatory Commission

CH 222 Chapter Twenty-one Study Guide

- Be able to characterize the three major types of radiation observed in radioactive decay (namely alpha, beta and gamma radiation.)
- Be able to write a balanced equation for nuclear reactions or transmutations.
- Know how to predict if a particular radioactive particle will decay by alpha, beta or positron emission or by electron capture.
- Know how to calculate the binding energy for a particular isotope and understand what this energy means in terms of nuclear stability.
- Be able to use the various first order rate equations as utilized in this chapter with respect to radioactivity.
- Be able to calculate the radioactive half-life of a radioactive isotope ($t_{1/2}$) from the activity of a sample. Also be able to find the time required for an isotope to decay using the half-life.
- Be able to describe nuclear chain reactions, nuclear fission and nuclear fusion.
- Understand the concept of background radiation. Know some of the sources of background radiation. Know the units used to measure intensity and understand how they pertain to health issues.
- Know some of the uses of radioactivity (carbon dating, medicine, etc.)
- Be able to solve and understand the assigned problems in problem set #6 from the kinetics chapter in our text - many of the radiation techniques stem from this chapter, so a thorough understanding is critical to success.
- Be able to solve and understand the assigned problems in problem set #6.

CH 222 Chapter Seven Concept Guide

1. Lewis Structures

Problem

Draw the Lewis Dot Structure for cyanide ion, CN^-.

Solution

Step 1. Add valence electrons.

1 C at 4 electrons = 4 electrons

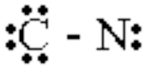
1 N at 5 electrons = 5 electrons

-1 charge = + 1 electron

Total = 10 electrons

Step 2. Place a bond between C and N to represent the sharing of 2 electrons. 8 electrons remain.
C-N

Step 3. Place lone pairs of the remaining 8 electrons around the outside atoms.



Step 4. In Step 3, nitrogen has only 4 valence electrons. Multiple bonds are needed to satisfy the octet rule. Two pairs of electrons are moved from C to form bonds between C and N.
:C≡N:

2. Lewis Structures

Problem

Draw the Lewis Dot Structure for acetaldehyde, CH_3CHO. Note: the right-hand H atom is bonded to the right-hand C atom.

Solution

Step 1. Add valence electrons.

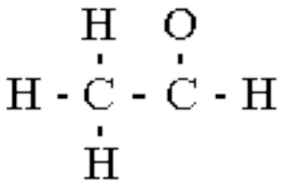
2 C at 4 electrons each = 8 electrons

4 H at 1 electron each = 4 electrons

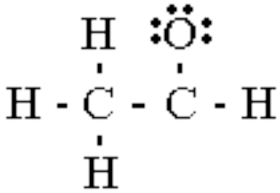
1 O at 6 electrons = 6 electrons

Total = 18 electrons

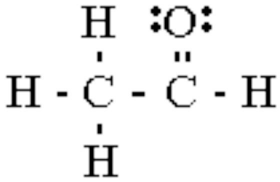
Step 2. Write the structure of CH_3CHO and place a bond between atoms to represent the sharing of 2 electrons. This uses 12 electrons. 6 electrons remain.



Step 3. The structure in Step 2 shows one of two carbon atoms and the oxygen atom without octets. Place lone pairs using the remaining 6 electrons around the oxygen atom.



Step 4. Multiple bonds are necessary to give carbon an octet, so a lone pair is moved from the oxygen atom to form a second bond between C and O.



3. Resonance Structures

Problem
Draw all resonance structures for O_3.

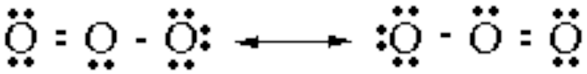
Solution
Step 1. Add valence electrons.
3 O at 6 electrons each = 18 electrons

Step 2. Place a bond between atoms. This uses 4 electrons. 14 electrons remain.
O-O-O

Step 3. Place lone pairs of the remaining 14 electrons around the oxygen atoms, starting with the terminal atoms. After each terminal O atom has 8 electrons, 2 electrons remain and are placed on the central atom.

Step 4. Use multiple bonds to obtain octets around each O.

Step 5. Electron delocalization leads to two resonance structures for O_3.



4. Lewis Structures

Problem
Draw the Lewis Dot Structure for $AlCl_3$.

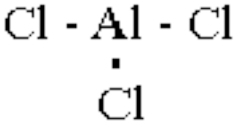
Solution
Step 1. Add valence electrons.

1 Al at 3 electrons = 3 electrons

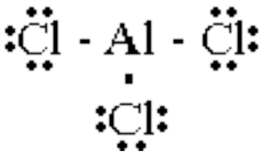
3 Cl at 7 electrons each = 21 electrons

Total = 24 electrons

Step 2. Place a bond between the atoms to represent the sharing of 2 electrons. 18 electrons remain.



Step 3. Place lone pairs of the remaining 18 electrons around the outside, terminal atoms. Each chlorine atom should have an octet of electrons. Neither Al nor Cl can form multiple bonds, so even though Al has only 6 electrons, this is the final Lewis structure.



5. Lewis Structures

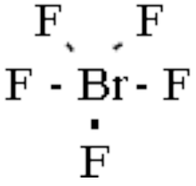
Problem
Draw the Lewis Dot Structure for BrF_5.
Solution:
Step 1. Add valence electrons.

1 Br at 7 electrons = 7 electrons

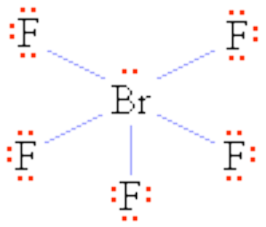
5 F at 7 electrons each = 35 electrons

Total = 42 electrons

Step 2. Place a bond between Br and F atoms to represent the sharing of 2 electrons. 32 electrons remain.

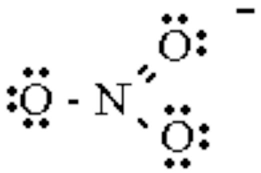


Step 3. Place lone pairs of the remaining 32 electrons around the 5 fluorine atoms. Each F atom has an octet of electrons. The remaining electrons are placed on the central atom.



6. Oxidation Numbers and Formal Charge

Problem
Determine the oxidation number and formal charge for N in the nitrate ion, NO_3^-.



Element	Electronegativity Values
H	2.1
B	2.0
C	2.5
N	3.0
O	3.5
F	4.0
Cl	3.0

Solution
Step 1. According to oxidation number rules, all valence electrons are considered to be held by the more electronegative atom. Therefore, each O atom has an oxidation number of –2. There are three oxygens for a total of –6. The charge on the ion is –1. Solving for x, where x is the oxidation number for nitrogen:

$x + (3)(-2) = -1$

$x = +5$

Nitrogen has an oxidation number of +5.

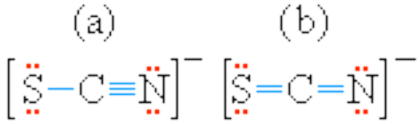
Step 2. The formal charge is based on the assumption that electrons are shared equally between covalently bonded atoms.

$$\text{Formal charge} = (\text{group number in periodic table}) - (\text{number of lone electrons}) - {}^{1}/_{2}\,(\text{number of bonding electrons})$$

For nitrogen, formal charge = 5 – 0 – 4 = +1. In this case, N has a formal charge of +1.

7. Resonance Structures

Question
Which is the better resonance structure for thiocyanate ion, SCN^-?



Solution
Step 1. Determine the formal charges on each element in both resonance structures.

Formal charge = (group number in periodic table) - (number of lone electrons)

– $^1/_2$ (number of bonding electrons)

Formal Charge

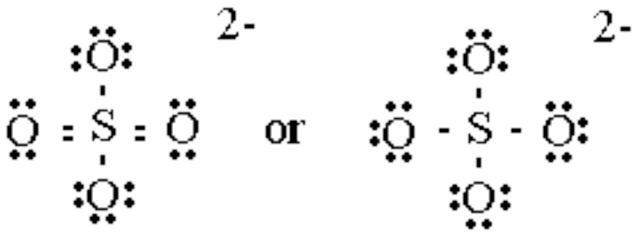
	Structure (a)	Structure (b)
Sulfur	+1	0
Carbon	0	0
Nitrogen	–2	-1

Step 2. The structure that has the fewest number of atoms with a formal charge is the better structure. Consequently, structure (b) is the correct one.

8. Using VSEPR: Predicting Geometries

Question
Using the VSEPR theory, what is the geometry of SO_4^{2-}?

Solution
Step 1. Draw the Lewis Dot Structure. There is more than one.

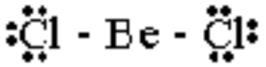


Step 2. Both structures have 4 bonded atoms and no lone pairs on the central S atom. Therefore, both the electron pair geometry and the molecular geometry are tetrahedral.
(Note: the VSEPR theory applies to both ions and molecules.) If the central S had one or more lone pairs, the electron pair and molecular geometries would differ.

9. Using VSEPR: Predicting Geometries

Question
Using the VSEPR theory, what is the geometry of $BeCl_2$?

Solution
Step 1. Draw the Lewis Dot Structure.

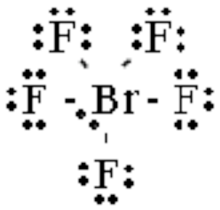


Step 2. Neither Be nor Cl is able to form multiple bonds, so the central atom Be remains electron deficient. There are no lone pairs on the central atom, Be. The electron pair geometry is linear. This molecule is linear.

10. Using VSEPR: Geometry and Bond Angles

Problem
Based on the VSEPR theory, predict the electron-pair geometry for BrF_5 and F-Br -F bond angles.

Solution
Step 1. Draw the Lewis structure of the molecule.



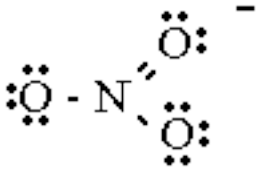
Step 2. Determine the number of bonded pairs and lone pairs of electrons around the central atom. In this molecule, there are 5 bonded atoms and 1 lone pair around Br. There are 6 structurally significant electron pairs.

Step 3. The electron pair geometry is octahedral. The ideal molecular geometry is square pyramidal.

Step 4. The ideal bond angles in an octahedral structure are 90 degrees. The angles in BrF_5 should be compressed slightly from the ideal 90 degrees angle due to the lone pair of electrons on Br.

11. Bond Properties and Resonance

Problem
NO_3^- has three resonance structures and a bond order of 1.3. Predict the bond length for N-O.



N-O single bond length = 136 pm
N-O double bond length = 115 pm
N-O triple bond length = 108 pm

Solution

Bond length depends partly on bond order. An N-O bond order of 1.3 suggests that the N-O bond length is a value between that for a single and double N-O bond. The bond length is 122 pm, which is in fact the length of a single N-O bond added to one third the difference between single and double N-O bond lengths.

Note: in order to use the $\Delta H_{formation}$ data, all reactants must be in the gas phase.

12. Calculating ΔH

Problem

Calculate ΔH for

$H_2(g) + F_2(g) \rightarrow 2\ HF(g)$.

using the bond energy data below.

H-H = 436 kJ/mol
F-F = 159 kJ/mol
H-F = 569 kJ/mol

Solution

In this reaction, we need to break the bonds between H-H and F-F and will form 2 H-F bonds. Because all the reactants and products are in the gas phase, we can relate ΔH for the reaction to the strengths of bonds broken and formed.

ΔH = (energy of bonds broken) - (energy of bonds formed)

$\Delta H = [(1\text{mol})(436\ \text{kJ/mol}) + (1\ \text{mol})(159\ \text{kJ/mol})] - (2\ \text{mol})(569\ \text{kJ/mol})$

$\Delta H = -543\ \text{kJ}$

13. Calculating ΔH

Problem

Calculate ΔH for

$H_2(g) + Br_2(g) \rightarrow 2\ HBr(g)$.

using the bond energy data given below.

H-H = 436 kJ/mol
Br-Br = 192 kJ/mol
H-Br = 368 kJ/mol

Solution

In this reaction, we need to break the bonds between H-H and Br-Br and will form 2 H-Br bonds.

ΔH = (energy of bonds broken) - (energy of bonds formed)

$\Delta H = [(1\ \text{mol})(436\ \text{kJ/mol}) + (1\ \text{mol})(192\ \text{kJ/mol})] - (2\ \text{mol})(368\ \text{kJ/mol})]$

$\Delta H = -108\ \text{kJ}$

14. Calculating ΔH

Problem

Calculate ΔH for

$C_4H_{10}(g) + 13/2\ O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)$

using the bond energy data given below.

C-C = 331 kJ/mol
C-H = 414 kJ/mol
O=O = 498 kJ/mol
C=O = 803 kJ/mol
O-H = 464 kJ/mol

Solution

In this reaction, find that 3 C-C, 10 C-H, and $^{13}/_{2}$ O=O bonds are being broken, and 8 C=O and 10 O-H bonds are being formed. Thus

ΔH = (energy of bonds broken) - (energy of bonds formed)

ΔH = [(3 mol)(331 kJ/mol) + (10 mol)(414 kJ/mol) + (13/2 mol)(498 kJ/mol)] - [(8 mol)(803 kJ/mol) + (10 mol)(464 kJ/mol)]

ΔH = -2694 kJ

15. Bond Polarity

Problem

Arrange the following covalent bonds in order of increasing polarity:

O-H I-Br C-F P-H S-Cl.

Electronegativity Values	
H	2.1
B	2.0
P	2.1
I	2.5
C	2.5
S	2.5
Br	2.8
N	3.0
O	3.5
F	4.0
Cl	3.0

Approach
Consider the differences in electronegativity values for each bonded pair of atoms.

Solution
The polarity of a bond increases with increasing difference in electronegativity of the bonded atoms. The covalent bonds, therefore, have the following order of increasing polarity with the electronegativity differences shown in parentheses:

→ **Increasing electronegativity** →

P-H (0.0) < I-Br (0.3) < S-Cl (0.5) < O-H (1.4) < C-F (1.5)

16. Molecular Polarity

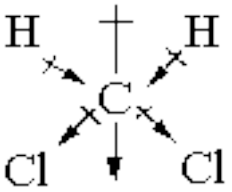
Question
Is NF_3 polar or nonpolar?

Solution
NF_3 has the same pyramidal structure as NH_3. Fluorine is more electronegative than N, thus each bond is polar. The NF_3 molecule is asymmetrical and polar.

17. Molecular Polarity

Question
Is CH_2Cl_2 polar or nonpolar?

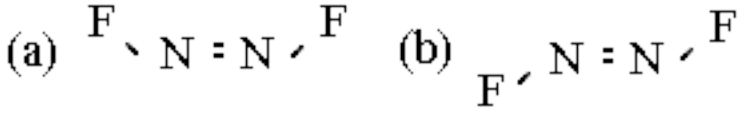
Solution:
In CH_2Cl_2, chlorine atoms are the most electronegative, followed by carbon atoms, then hydrogen atoms.



There is a net movement of electron density away from H atoms and toward Cl atoms. The asymmetric arrangement of the atoms, and the negative end of the bond, is toward the two Cl atoms. The positive end toward the 2 H atoms makes CH_2Cl_2 a polar molecule.

18. Molecular Polarity

Problem
There are two different molecules with the formula N_2F_2. Is either molecule polar?



Solution
Fluorine atoms are more electronegative than nitrogen atoms. The negative ends of the bonds are toward the two fluorine atoms and the positive ends are toward the two nitrogen atoms. In molecule (a), the two dipoles do not cancel each other, thus this molecule is polar. In molecule (b), however, the two dipoles are opposite in direction and do cancel each other, making this molecule nonpolar.

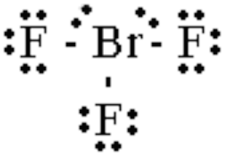
CH 222 Chapter Eight Concept Guide

According to **valence bond theory**, the s and p orbitals of any atom give a maximum of four hybrid orbitals. Atoms of second-period elements, which have only s and p orbitals, may form up to four hybrid orbitals and no more than four covalent bonds around the central atom. Atoms of the elements in the third and higher periods can form a larger number of covalent bonds by involving d orbitals in hybridization. For sp^3d hybridization, the number of hybrid orbitals is 5 and the electronic geometry is trigonal bipyramidal. For sp^3d^2 hybridization, there are 6 hybrid orbitals and the electronic geometry is octahedral.

1. Valence Bond Theory

Problem
Describe the bonding in BrF_3 in terms of valence bond theory.

Solution
The Lewis structure shows that 3 single bonds and two lone pairs surround the bromine atom:



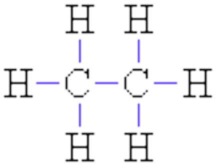
Five equivalent hybridized orbitals on the bromine atom are necessary. The outer electron configuration for the bromine atom is: $4s^24p^5$. By using the 4s orbital, 3 4p orbitals, and one of the empty 4d orbitals in sp^3d hybridization, five hybrid orbitals can be formed. Two of the hybrid orbitals contain lone pairs of electrons. The other hybrid orbitals contain single electrons, which will form single bonds with the fluorine atoms.

The five hybridized orbitals will be arranged in the shape of a trigonal bipyramid.

2. Valence Bond Theory

Problem
Describe the bonding in C_2H_6 in terms of valence bond theory.

Solution



Each carbon atom must have four equivalent hybridized orbitals that are formed by sp^3 hybridization. Three of these orbitals on each carbon atom contain a single electron that will form a single bond with a hydrogen atom. The fourth orbital contains a single electron that will form a single bond with the other carbon atom.

3. Valence Bond Theory

Problem
Describe the bonding in H_2O_2 in terms of valence bond theory.

Solution

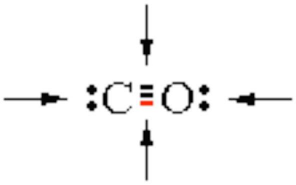
Each oxygen has four equivalent hybrid orbitals formed by sp^3 hybridization. Two of these orbitals contain lone pairs of electrons, one contains a single electron that forms a single bond with a hydrogen atom, and one contains a single electron that will form a single bond with the other oxygen atom.

4. Multiple Bonds

Problem
Describe the bonding in carbon monoxide, CO, using hybrid orbital theory.

Solution
The Lewis structure depicts C and O being bonded by a triple bond and each having a single lone pair.

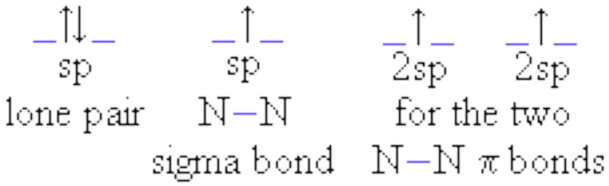
Each atom has a half-filled sp hybrid orbital it uses for sigma bond formation. Each atom also has a sp hybrid orbital that contains a lone electron pair. In addition, two pairs of electrons occur in unhybridized p orbitals and are used to form two π bonds.



5. Multiple Bonds

Problem
Describe the bonding in a nitrogen molecule, N_2, using hybrid orbital theory.

Solution
The bonding in a nitrogen molecule is identical to that in CO, except that both atoms are nitrogen: there is one sigma bond, two π bonds, and one lone pair on each atom. Each N has linear electron-pair geometry and is, therefore, sp-hybridized. The assignments of the five valence electrons on each N and their roles in bonding are:



6. Multiple Bonds

Problem
Describe the hybridization of both carbon atoms and of the nitrogen atom in acetonitrile:
$CH_3\text{-}C\equiv N\colon$

Solution
The CH_3 carbon has tetrahedral electron-pair geometry, and is, therefore, sp^3 hybridized. The CN carbon has linear electron-pair geometry and is, therefore, sp-hybridized. Last, the N atom has linear electron-pair geometry and is sp-hybridized. Two unhybridized orbitals on the central carbon and two on the nitrogen are used to form two π bonds.

7. Molecular Orbitals and Bond Order

Question
Write the electron configuration of the H_2^- ion in molecular orbital terms. What is the bond order of this ion?

Solution
This molecular ion has three electrons: one from each H atom and one due to the negative charge. Its configuration, therefore, is $(\sigma_{1s})^2\ (\sigma^*{}_{1s})^1$.

H_2^- has a net bond order of $^1/_2$ and the ion is predicted to exist under special circumstances:
$^1/_2$(2 bonding electrons - 1 antibonding electron) = $^1/_2$

8. Molecular Orbitals and Bond Order

Question
Write the configuration of the H_2^+ ion in molecular orbital terms. Compare the bond order of this ion to He_2^+ and H_2^-. Do you expect H_2^+ to exist?

Solution
The molecular orbital configuration for H_2^+ is $(\sigma_{1s})^1$. This ion has a bond order of $^1/_2$, as do He_2^+ and H_2^-:
$^1/_2$(2 bonding electrons - 1 antibonding electron) = $^1/_2$

H_2^+, therefore, is expected to exist.

9. Molecular Orbitals in Diatomic Molecules

Question
Knowing that Be_2 does not exist, describe the electron configuration in molecular orbital terms for Be_2^+ and give its net bond order. Do you expect Be_2^+ to exist?

Solution
The Be_2^+ molecular ion has seven electrons. Four of the seven electrons are core electrons, and are assigned to σ_{1s} and σ_{1s}^* molecular orbitals. The remaining three electrons are assigned to the σ_{2s} and σ_{2s}^* orbitals. The molecular orbital configuration is:
[core electrons] $(\sigma_{2s})^2\ (\sigma_{2s}^*)^1$

The net bond order is: $^1/_2$(2 bonding electrons - 1 antibonding electron) = $^1/_2$

Be_2^+ has a net bond order of $^1/_2$, thus it is expected to exist.

10. Metallic bonding

Problem
Magnesium has its highest energy band filled by two 3s electrons contributed by each atom, thus it is expected to be an insulator. Explain why magnesium is not an insulator, but a conductor of electricity.

Solution:
The empty 3p orbitals form an energy band that overlaps the 3s band, creating a larger, partially filled band. Electrons may, therefore, move up and out of 3s orbitals to occupy vacant 3p orbitals.

11. Metals and Semiconductors

Question
As the temperature increases, more electrons in a semiconductor gain energy needed to jump out of the valence band into the conduction band. How does this affect the resistivity of the semiconductor?

Solution
With an increase in temperature, as a greater number of electrons can jump out of the valence band into the conduction band, the resistivity of the semiconductor decreases. The amount of decrease is different for each semiconductor and depends on the band gap. At low temperatures, the conductivity of semiconductors is similar to that of insulators; at high temperatures, it is like that of metals.

12. P-type and N-type Semiconductors

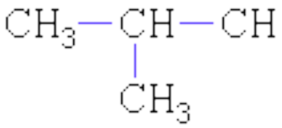
Question
Recall that a p-type semiconductor has been defined in this lesson as a semiconductor that conducts a positive charge. What would you expect from an n-type semiconductor? What is meant by a p-n junction?

Solution
An n-type semiconductor conducts a negative charge. It is formed by doping Si with a group 5A element. This is so the solid has orbitals with extra electrons that can jump to the conduction band. A p-n junction is the boundary between p-type and n-type semiconductors. Doping adjacent areas in the same crystal with Group 3A and Group 5A elements creates a p-n junction.

CH 222 Chapter Twenty Concept Guide

1. Organic Nomenclature

Question
Is the following compound an alkane, alkene, or alkyne; saturated or unsaturated; branched or straight chain?



$CH_3-\underset{\underset{\displaystyle CH_3}{|}}{CH}-CH$

Solution:
It is a branched-chain, saturated alkane. All organic compounds with all C-C single bonds are saturated. Compounds with only C-C single bonds are alkanes.

2. Organic Nomenclature

Question
Is the following compound an alkane, alkene, or alkyne; saturated or unsaturated; branched or straight chain?
$CH_3-CH_2-C\equiv C-CH_3$

Solution:
This is 2-butyne. It is a straight chain, unsaturated alkyne. All organic compounds with double or triple C-C bonds are unsaturated. Compounds with C-C triple bonds are alkynes.

3. Reactions of Organic Compounds

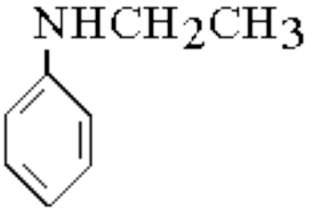
Problem
Predict the product of the hydrogenation reaction of 1-butene and $H_2(g)$: $CH_2{=}CHCH_2CH_3 + H_2$

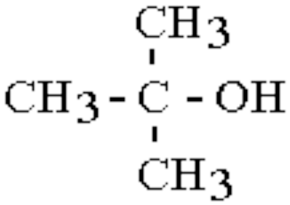
Approach
This is a hydrogenation reaction, thus H atoms will add across the C-C double bond forming an alkane.

Solution:
The product is butane: $CH_3CH_2CH_2CH_3$

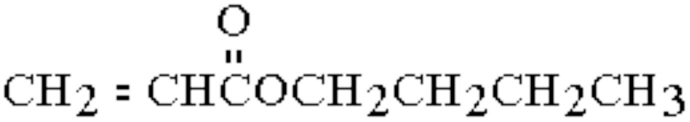
4. Classification of Organic Compounds

Problem
Classify the following compounds according to the types of compounds listed below.

(a)


(b)


$CH_3-\underset{CH_3}{\overset{CH_3}{C}}-OH$

(c)


$CH_2=CH\overset{O}{\overset{\|}{C}}OCH_2CH_2CH_2CH_3$

<u>Types of Organic Compounds</u>
Alcohol Aldehyde Amine Carboxylic Acid
Ester Ketone Phenol

Approach
Identify the functional groups and the hydrocarbon portions in each molecule.

Solution
(a). An amine (secondary). The hydrocarbon is C_6H_5 and C_2CH_3.
(b). An alcohol. The hydrocarbon is C_4H_9.
(c). An ester. The hydrocarbon is $CH_2{=}CH$ and $CH_2CH_2CH_2CH_3$.

5. Synthesizing Carboxylic Acids

Problem
The reaction of methanol and carbon monoxide yields a carboxylic acid that is produced in bread when leavened by a particular yeast, *Saccharomyces exigus*. Predict this product.

Approach
Write out the reaction of methanol and carbon monoxide. Balance the equation.

Solution

$$CH_3OH(l) + CO(g) \rightarrow CH_3CO_2H(l)$$

The product is acetic acid.

CH 222 Chapter Nine Concept Guide

1. Boyle's Law

Question
A sample of gaseous nitrogen in an automobile airbag has a pressure of 745.0 mm Hg in a 35.00-L bag. If this sample is transferred to a 15.00-L bag at the same temperature, what is the pressure (in mm Hg and atm) of the gas in the new bag?

Approach
Boyle's Law says that pressure is indirectly proportional to volume: $P_1V_1 = P_2V_2$.
We know P_1, V_1, and V_2 and need to solve for P_2.

Solution
$P_2 = P_1V_1/V_2 = (745.0 \text{ mm Hg} \times 35.00 \text{ L})/15.00 \text{ L} = 1738 \text{ mm Hg}$
$1738 \text{ mm Hg} \times (1 \text{ atm}/760 \text{ mm Hg}) = 2.29 \text{ atm}$

As expected, the pressure of a gas increases as volume decreases.

2. Charles's Law

Question
The gas volume of CO_2 in a syringe is 25.0 L at 20 °C. What is the final volume of the gas if you hold the syringe in your hand until the temperature reaches 39 °C?

Approach
Charles's Law says that the volume of a gas is directly proportional to the absolute temperature:
$V_1/T_1 = V_2/T_2$. We know T_1, V_1, and T_2. Solve for V_2. Note: Don't forget to convert temperature to Kelvins.

Solution
$V_2 = T_2(V_1/T_1) = (312 \text{ K})(25.0 \text{ mL}/293 \text{ K}) = 26.6 \text{ mL}$

As expected, the volume of a gas increases with an increase in temperature.

3. Ideal Gas Law

Question
If you wanted to use the Ideal Gas Law to solve for volume, how would the equation look?

Approach
Isolate volume from the other variables in the Ideal Gas Law.

Solution
$V = nRT/P$

4. Gas Density and Molar Mass

Question
The density of an unknown gas is 1.50 g/L at STP. What is its molar mass?

Approach
Use the Ideal Gas Law, substituting density for n and molar mass (M; g/mol) for V. Solve for M.

Solution
Density is mass per unit volume and can be used to convert volume into mass: d = m / V. Therefore,
PM = dRT
M = dRT/P
M = [(1.50 g/L)(0.082057 L atm/K mol)(273.15 K)]/1.000 atm = 33.6 g/mol

5. Root-mean-square Speed

A relationship exists among molecular mass, average speed, and temperature. Two gases with different molecular masses must have the same kinetic energy at the same temperature because the average kinetic energy is fixed by temperature. The heavier gas molecules, therefore, must have a lower average speed. Sometimes called the "Maxwell Equation," the **root-mean-square (rms) speed** expresses this idea in quantitative form:

$$\sqrt{\overline{u^2}} = \sqrt{3RT/M}$$

where $\sqrt{\overline{u^2}}$ is called the rms speed, temperature (T) is in Kelvins, M is molar mass, and R is expressed in units related to energy, 8.314 J / K mol.

6. Graham's Law

Gases have the ability to diffuse and effuse. Diffusion is the mixing of two or more substances by random molecular motion; effusion is the movement of gas through a tiny opening in a container into another container.

Thomas Graham (1805 - 1869) studied effusion and found that the rates of effusion of two gases were inversely proportional to the square roots of the molar masses at the same temperature and pressure.

$$\frac{\text{Rate of effusion of gas 1}}{\text{Rate of effusion of gas 2}} = \sqrt{\frac{\text{molar mass of gas 2}}{\text{molar mass of gas 1}}}$$

Therefore, the rate at which a gas will escape through an opening depends on how fast the gas molecules move. This equation is derived from Maxwell's equation.

$$\frac{\text{Rate of effusion of gas 1}}{\text{Rate of effusion of gas 2}} = \frac{\sqrt{\overline{u^2}}\text{ of gas 1}}{\sqrt{\overline{u^2}}\text{ of gas 2}} = \sqrt{\frac{3RT/(M\text{ of gas 1})}{3RT/(M\text{ of gas 2})}}$$

Canceling out like terms leaves us with the simple equation developed by Graham.

7. Molecular Speeds

Problem

Place the following gases in order of decreasing average molecular speed at 25 °C:

Ar CH_4 N_2 CH_2F_2

Approach

Remember that the average kinetic energy of gas molecules is determined by temperature, thus all these gases have the same average kinetic energy and heavier molecules must move with a slower average speed than lighter molecules at the same temperature.

Solution:

Compound	Molecular Weights
CH_4	16.04 g/mol
N_2	28.01 g/mol
Ar	39.90 g/mol
CH_2F_2	52.02 g/mol

8. Molecular Speeds

Problem

Molecular speed is important in the atmospheric diffusion of gases. Rank O_2, O_3, NO_2, and NO in the order of increasing average molecular speed.

Approach

Remember that the average kinetic energy of gas molecules is determined by temperature, thus all these gases have the same average kinetic energy and heavier molecules must move with a slower average speed than lighter molecules at the same temperature.

Solution:

Compound	Molecular Weights
O_3	48.00 g/mol
NO_2	46.01 g/mol
O_2	32.00 g/mol
NO	30.01 g/mol

$O_3 < NO_2 < O_2 < NO$

9. Dalton's Law of Partial Pressures

Question
A 5.00 L sample of N_2 at 738 Torr is mixed at constant temperature with 15.5 L of O_2 at 325 Torr. The gaseous mixture is placed in a 10.0 L container. What is the pressure of the mixture?

Approach
We need to apply Boyle's Law to each gas, then use Dalton's Law of partial pressures to find the pressure of the mixture.

Solution
For N_2: $P_1 = 738$ Torr, $P_2 = ?$, $V_1 = 5.00$ L, and $V_2 = 10.0$ L. When we solve for the unknown, $P_2 = 369$ Torr.

For O_2: $P_1 = 325$ Torr, $P_2 = ?$, $V_1 = 15.5$ L, and $V_2 = 10.0$ L. When we solve for the unknown, $P_2 = 504$ Torr.

According to Dalton's Law of partial pressure, the total pressure of the mixture is:

$$P_{total} = P_{O2} + P_{N2} = 369 \text{ Torr} + 504 \text{ Torr} = 873 \text{ Torr}$$

10. Using Partial Pressures to Measure a Gas Collected over Water

Problem
A volume of 550 mL of gas was collected over water at 21 °C. The atmospheric pressure was 0.980 atm. The vapor pressure of water at this temperature is 0.025 atm. Calculate the partial pressure in atmospheres of the gas collected over water.

Approach
Use the following equation to find the partial pressure of the gas:

$$P_{atm} = P_{H2O} + P_{gas}$$

Solution

$$P_{gas} = P_{atm} - P_{H2O}$$

$$P_{gas} = 0.980 \text{ atm} - 0.025 \text{ atm} = 0.955 \text{ atm}$$

11. Root-mean-square Speed

Problem
Calculate the rms speed of oxygen molecules at 27 °C.

Approach
We need Maxwell's equation to calculate the rms speed. Before plugging in numbers, however, we need to convert M to units of kilograms per mole because R is in units of J / K mol. The necessary conversion factor is: $1 \text{ J} = 1 \text{ kg m}^2 / \text{s}^2$.

Solution

$$\sqrt{\overline{u^2}} = \sqrt{\frac{3RT}{M}}$$

$$= \sqrt{\frac{3(8.314 \text{ J / K mol})(300 \text{ K})}{3.20 \times 10^{-2} \text{ kg / mol}}} = \sqrt{2.34 \times 10^5 \text{ J/kg}}$$

To obtain the answer in meters per second, we use the relation I J = 1 kg m^2 / s^2, which means we have

$$\sqrt{\overline{u^2}} = \sqrt{\frac{3RT}{M}}$$

$$= \sqrt{2.34 \times 10^5 \text{ J/kg}} = \sqrt{2.34 \times 10^5 \text{ kg m}^2\text{/kg s}^2} = 483 \text{ m/s}$$

12. Graham's Law

Question
How fast do helium molecules effuse through a barrier relative to oxygen molecules?

Approach
We need Graham's Law to determine the rate differential between helium and oxygen molecules.

Solution

$$\frac{\text{Rate of effusion of He}}{\text{Rate of effusion of O}_2} = \sqrt{\frac{\text{molar mass of O}_2}{\text{molar mass of He}}}$$

$$= \sqrt{\frac{32.00 \text{ g/mol}}{4.00 \text{ g/mol}}} = 2.83$$

Helium molecules effuse through a barrier 2.83 times faster than oxygen molecules.

CH 222 Chapter Ten Concept Guide

1. Types of Intermolecular Forces

Question

What type if intermolecular force is involved in each case below? Place the following cases in order of decreasing strength of interaction.

(a) Liquid methane, CH_4
(b) Mixture of water and methanol, H_2O and CH_3OH
(c) Solution of LiCl in water

Solution

(a) No ions are involved with CH_4. It is a simple molecule with covalent bonds. It is also nonpolar, thus the only way methane molecules can interact is through induced dipole forces.

(b) Similarly to CH_4, there are no ions involved with these covalently bonded molecules. H_2O and CH_3OH are, however, polar and both have an O-H bond. They interact, therefore, through special dipole-dipole forces: hydrogen bonding.

(c) LiCl is an ionic compound composed of Li^+ and Cl^- ions and water is a polar molecule. In this case, the salt dissociates in water and the ions interact with water molecules through ion-dipole forces.

In order of decreasing strength, the interactions are:

LiCl in H_2O H_2O in CH_3OH liquid CH_4.

2. Types of Intermolecular Forces

Question

What type of intermolecular forces is involved in each case below? Place the following cases in order of decreasing strength of interaction.

(a) liquid O_2
(b) $MgSO_4$ dissolved in water
(c) O_2 dissolved in H_2O

Solution

(a) O_2 interactions occur by induced dipole-induced dipole forces. These are the weakest of all forces.

(b) $MgSO_4$ dissociates into Mg^{2+} and SO_4^{2-} ions when dissolved in water. The ions interact with water through ion-dipole forces.

(c) Water is a polar molecular compound. O_2 is composed of nonpolar molecules. Dipole-induced dipole forces exist between H_2O and O_2.

In order of decreasing strength, the interactions are:

$MgSO_4$ dissolved in H_2O O_2 dissolved in H_2O liquid O_2.

3. Explaining Differences in Boiling Points

Problem

Explain why the boiling point of H_2S is lower than that of H_2O.

Solution

There is significant hydrogen bonding between the molecules in H_2O, but not in H_2S. This difference arises because oxygen's electronegativity is much greater than sulfur's electronegativity. Hydrogen bonding increases a substance's boiling point, thus H_2O has a higher boiling point than H_2S. Recall from this lesson that the difference between the boiling points of ethanol and dimethyl ether is also due to differences in the strength of forces, and more specifically, the lack of hydrogen bonding in dimethyl ether.

4. Characteristics of Liquids

Question

Why do gases completely fill their containers and liquids spread to take the shape of their containers?

Answer

Gases and liquids flow simply due to the movement of molecules. In comparison, a solid has too rigid a structure to allow for flow. Its molecules cannot move past one another.

5. Characteristics of Liquids

Question

What factors determine the viscosity of a liquid? How does viscosity change with a decrease in temperature?

Answer

The size, shape, and chemical nature of molecules, as well as a sample's temperature, determine the viscosity of a liquid. As temperature decreases, viscosity increases due to less rapid movement of the molecules and a lessened ability to overcome intermolecular forces in order to move past one another.

6. Normal Boiling Point

Question

The normal boiling point of CH_3Cl, chloromethane, is -24 °C and the normal boiling point for CH_3I is 42.4 °C. Which compound has the stronger intermolecular forces in the liquid phase? What types of intermolecular forces are involved?

Answer

The forces involved are dipole-dipole attractions and dispersion (induced dipole) forces. CH_3Cl has a larger dipole (1.892 D) than CH_3I, therefore one might expect CH_3Cl to have stronger intermolecular forces as a liquid. However, the larger iodine atom in CH_3I is more polarizable, yielding greater induced dipole attractive forces than the chlorine atom in CH_3Cl, giving CH_3I a higher normal boiling point.

7. Vapor Pressure: Water

Question

Why does water boil at a lower temperature at the top of a mountain than it does at sea level?

Answer

External atmospheric pressure decreases with increasing altitude. When water is heated on a mountaintop, its vapor pressure reaches this lower atmospheric pressure at a lower temperature. The water boils sooner, yet it will take longer to cook food in this water because the boiling point is lower.

8. Lattice Energy

Question

Lattice energy is defined as the net force of attraction between ions in an extended solid. It is the energy released if one mole of gas phase ions would come together to form the solid lattice. Write a chemical equation illustrating the reaction between M^{2+} and X^{2-} that fits this definition. How does the value of the lattice energy change as ionic charge decreases and as ionic radius decreases?

Answer

The energy liberated as gaseous ions combine to give a crystalline ionic structure may be illustrated as:

$$M^{2+}\ (g) + X^{2-}\ (g) \rightarrow MX\ (s)$$

The lattice energy depends on Coulomb’s Law. Therefore, lattice energy decreases with decreasing ionic charge and increases with decreasing ionic radius.

9. Lattice Energy

Question

Which compound should have the higher lattice energy: $CaCl_2$ or $BaCl_2$?

Answer

Lattice energy depends inversely on the size of the ions involved: the smaller the ions, the stronger the electrostatic attraction between them, and the larger the lattice energy. Here, Ca^{2+} is smaller than Ba^{2+}, thus $CaCl_2$ is anticipated to have the higher lattice energy.

10. Summary of Cubic Unit Cells

Question

How many net atoms or ions does a (a) simple cubic unit cell, (b) body-centered cubic unit cell, and (c) face-centered cubic unit cell have?

Answer

(a) A simple cubic unit cell has 8 atoms in corner positions, and one eighth of each is contained in each unit cell. So, a simple cubic unit cell of atoms or ions always contains 1 net atom or ion.

(b) A body-centered cubic unit cell is similar to a simple cubic unit cell, but also contains an atom or ion at its center. Therefore, a body-centered cubic unit cell of atoms always contains 2 net atoms within the cell.

(c) A face-centered cubic unit cell of X atoms or ions always contains 4 net atoms or ions within the cell.

11. Unit Cell - Volume and Density

Question

The unit cell length of diamond was measured as 0.3567 nm.

(a) What is the volume of this cubic unit cell in cubic centimeters?

(b) If the mass of unit cell of a diamond is 1.60×10^{-22} g, what is the theoretical density of diamond?

Approach

The volume of the cubic unit is found by cubing the length of the side. Density can be calculated by dividing the mass of diamond by the volume.

Solution

(a) Conversion factors are needed to express the unit cell length in centimeters. Cubing this number gives the volume of the cubic unit.

$$\left(0.3567\ \text{nm}\right)\left(\frac{1\ \text{m}}{10^{9}\ \text{nm}}\right)\left(\frac{100\ \text{cm}}{1\ \text{m}}\right) = 3.567 \times 10^{-8}\ \text{cm}$$

Volume = $(3.567 \times 10^{-8})^{3} = 4.538 \times 10^{-23}\ \text{cm}^{3}$

(b) To calculate the theoretical density of diamond having a mass of 1.60×10^{-22} g, divide the mass of the unit cell by its volume.

$$\frac{1.60 \times 10^{-22}\ \text{g}}{4.538 \times 10^{-23}\ \text{cm}^{3}} = 3.50\ \text{g/cm}^{3}$$

12. Structures and Formulas of Ionic Solids

Question

Can calcium chloride have a unit cell like that of sodium chloride?

Solution

A unit cell of NaCl can only have a 1:1 cation to anion ration. $CaCl_2$ has a 1:2 cation to anion ratio, therefore it cannot have a unit cell like NaCl.

13. Determination of an Atomic Radius from Measurements of a Crystal Lattice: Aluminum

Problem

Aluminum has a density of 2.70 g/cm^3 and the atoms are packed into a face-centered cubic unit cell. Calculate the radius of an aluminum atom.

Approach

First, calculate the mass of a unit cell with the knowledge that it is face-centered cubic, and thus has 4 atoms per unit cell. Second, combine the density of aluminum with the mass of the unit cell to calculate the volume of the unit cell. The volume may then be used to find the length of an edge of the unit cell. Last, calculate the radius of an aluminum atom from the edge dimension.

Solution

Step 1. The mass of the unit cell is

$$\left(26.98\ \mathrm{g/mol\ Al}\right)\left(\frac{1\ \mathrm{mol}}{6.022 \times 10^{+23}\ \mathrm{atoms}}\right)\left(4\ \mathrm{atoms/unit\ cell}\right) = 1.792 \times 10^{-22}\ \mathrm{g/unit\ cell}$$

Step 2. Calculate the volume of the unit cell by dividing the mass of the unit cell by the density of aluminum.

$$\left(1.792 \times 10^{-22}\ \mathrm{g/unit\ cell}\right)\left(\frac{1\ \mathrm{cm}^3}{2.70\ \mathrm{g}}\right) = 6.64 \times 10^{-23}\ \mathrm{cm}^3\mathrm{/unit\ cell}$$

Step 3. To find the length of a unit cell edge, calculate the cube root of the cube volume.

$$\sqrt[3]{6.64 \times 10^{-23}}\ \mathrm{cm}^3\mathrm{/unit\ cell} = 4.05 \times 10^{-8}\ \mathrm{cm}$$

Step 4. The diagonal distance across a face-centered cubic unit cell is

$$\text{Diagonal distance}^2 = (\text{edge})^2 + (\text{edge})^2 = 2\ (\text{edge})^2$$

Taking the square root of both sides yields:

$$\text{diagonal distance} = \left(\sqrt{2}\right)\left(\text{cell edge}\right)$$

Plugging in the cell edge:

$$\text{diagonal distance} = \left(\sqrt{2}\right)\left(4.05 \times 10^{-8}\ \text{cm}\right) = 5.73 \times 10^{-8}\ \text{cm}$$

To find the atomic radius of aluminum, divide the diagonal distance by 4.

$$\frac{5.73 \times 10^{-8}\ \text{cm}}{4} = 1.43 \times 10^{-8}\ \text{cm}$$

The atomic radius for aluminum is 1.43×10^{-8} cm or 143 pm.

14. Critical Points

Question

What is the critical point? Will a substance always be a liquid below the critical point?

Answer

The critical point is the temperature above which no amount of pressure will cause condensation (movement of molecules from the gas phase to the liquid phase). No, a substance will not always be a liquid below the critical point. All solids, for example, are below the critical point and many gases exist below their critical point.

15. Triple Points

Question

How many phases can co-exist at equilibrium at a triple point? Refer to the phase diagram for water in this lesson.

Answer

There are three phases at a triple point: solid, liquid, and gas.

16. Phase Changes: CO_2

Question

Refer to a phase diagram for CO_2. If a sample of CO_2 is at its triple point, what phase is present after the pressure has been increased at constant temperature?

Answer

After increasing pressure, CO_2 will be in its solid form. The solid form is favored over the liquid at high pressure because $CO_2(s)$ is more dense than $CO_2(\ell)$.

CH 222 Chapter Eleven Concept Guide

1. Molality

Question

A 4.5 M nitric acid solution contains 65.0 g of HNO_3 in 288 g of solution. What is the molality of this solution?

Approach

Molality is calculated by dividing the moles of solute by the kilograms of solvent.

Solution

Step 1. Calculate the number of moles of HNO_3.

$$\left(65.0\text{ g }HNO_3\right)\left(\frac{1\text{ mol }HNO_3}{63.01\text{ g }HNO_3}\right) = 1.03\text{ mol }HNO_3$$

Step 2. Calculate the molality of the solution.

There are 223 g of solvent: 288 g solution - 65.0 g solute = 223 g solvent.

Molality = moles of solute/kilograms of solvent

$$\text{Molality} = \frac{1.03\text{ mol }HNO_3}{0.223\text{ kg}} = 4.62\text{ mol/ kg}$$

2. Weight percent

Question

What is the weight percent of methanol in a solution of 1.0 L of methanol in 2.5 L of diethyl ether? The density of methanol is 0.79 g/mL and the density of diethyl ether is 0.71 g/mL.

Approach

To find the weight percent of methanol, we must know the masses of each component in the solution. Then, we can calculate the weight percent by dividing the mass of methanol by the mass of solution, and then multiplying by 100.

Solution

Step 1. Calculate the mass of methanol and the mass of diethyl ether.

$$\left(1.0\text{ L methanol}\right)\left(\frac{1000\text{ mL}}{1\text{ L}}\right)\left(0.79\text{ g/ mL}\right) = 790\text{ g methanol}$$

$$\left(2.5\text{ L diethyl ether}\right)\left(\frac{1000\text{ mL}}{1\text{ L}}\right)\left(0.71\text{ g/ mL}\right) = 1800\text{ g diethyl ether}$$

Step 2. Calculate the mass of the solution.

790 g methanol + 1800 g diethyl ether = 2600 g solution

Step 3. The weight percent of methanol in the solution is calculated by dividing the mass of methanol by the mass of the solution, then multiplying by 100.

$$\left(\frac{790\text{ g methanol}}{2600\text{ g solution}}\right)\left(100\right) = 30.\%\text{ methanol}$$

3. Solution Concentration

Problem

Commercial vinegar is an aqueous solution that must contain at least 4 weight percent acetic acid. The density of this solution is 1.0058 g/mL and the volume of the sample is 100.00 g. Calculate the concentration of acetic acid, if this vinegar is exactly 4.000% acetic acid, in terms of (a) mole fraction, (b) molality, and (c) molarity. Acetic acid is CH_3CO_2H and has a molar mass of 60.05 g.

Approach

To find the mole fraction, we need to calculate the numbers of moles of solute and solvent in 100.00 g of solution. Then, divide the number of moles of acetic acid by the total number of moles of solution. To find the molality, we will need to use a ratio of moles of acetic acid to kilograms of solvent. Last, to find the molarity of the solution, simply divide the moles of acetic acid by the number of liters of solution.

Solution

(a) Mole Fraction

Step 1. Calculate the mass and moles of both acetic acid and water.

$$4\text{ mass percent }CH_3CO_2H = \frac{\text{mass }CH_3CO_2H}{\text{mass solution}}$$

Mass CH_3CO_2H = 0.04 x 100.00 = 4.00 g CH_3CO_2H

$$\text{Moles of }CH_3CO_2H = \left(4.00\text{ g }CH_3CO_2H\right)\left(\frac{1\text{ mol }CH_3CO_2H}{60.05\text{ g }CH_3CO_2H}\right) = 0.0666\text{ mol }CH_3CO_2H$$

Mass of H_2O = mass of solution - mass of CH_3CO_2H
= 100.00 g solution - 4.00 g CH_3CO_2H = 96.00 g H_2O

$$= \left(96.00\text{ g }H_2O\right)\left(\frac{1\text{ mol }H_2O}{18.015\text{ g }H_2O}\right) = 5.329\text{ mol }H_2O$$

Step 2. Calculate the mole fraction of acetic acid.

$$\text{Mole fraction } CH_3CO_2H = \frac{\text{moles } CH_3CO_2H}{\text{moles solution}}$$

$$= \frac{0.0666 \text{ mol } CH_3CO_2H}{0.0666 \text{ mol } CH_3CO_2H + 5.329 \text{ mol } H_2O} = 0.0123$$

(b) Molality

$$\text{Molality} = \frac{\text{moles } CH_3CO_2H}{\text{kg solution}}$$

$$= \frac{0.0666 \text{ mol } CH_3CO_2H}{0.09600 \text{ kg } H_2O}$$

$$= 0.694 \text{ mol/kg } H_2O$$

(c) Molarity

Step 1. Calculate the volume of the solution.

$$(100 \text{ g solution})\left(\frac{1 \text{ mL}}{1.0058 \text{ g}}\right) = 99.42 \text{ mL solution}$$

Step 2. Calculate the molarity by dividing the moles of solute by the liters of solution.

$$= \frac{0.0666 \text{ mol } CH_3CO_2H}{0.09942 \text{ L solution}} = 0.670 \text{ mol/ L}$$

4. Parts Per Million

Problem

You have a solution consisting of 2.665 g of solute and 151.78 g of solution. Convert this to units of ppm.

Approach

To calculate ppm, convert grams of solute to milligrams, and grams of solution to kilograms.

Solution

$$\frac{2665 \text{ mg solute}}{0.15178 \text{ kg solution}} = 17{,}560 \text{ ppm}$$

5. Solubility and Henry's Law

Question

A soft drink has an aqueous CO_2 concentration of 0.0511 M at 25 °C. What is the pressure of CO_2 gas in the drink? Henry's law constant for CO_2 is 4.48 x 10^{-5} M/mm Hg at 25 °C.

Approach

We need to use Henry's law

$$S_g = k_H P_g$$

where S_g is gas solubility, k_H is Henry's law constant, and P_g is the partial pressure of CO_2. Substituting the gas solubility and Henry's law constant for CO_2 into the equation will yield the pressure of CO_2.

Solution

$$P_g = \frac{S_g}{k_H}$$

$$P_g = \frac{0.0511 \text{ M}}{4.48 \times 10^{-5} \text{ M/mm Hg}} = 1.14 \times 10^{3} \text{ mm Hg}$$

6. Solubility and Henry's Law

Question

The partial pressure of O_2 in a person's lungs varies from 22 mm Hg to 40 mm Hg. How much O_2 can dissolve in water at 25 °C if the partial pressure of O_2 is 35 mm Hg? Henry's law constant for O_2 is 1.66 x 10^{-6} M/mm Hg at 25 °C.

Solution

We need to use Henry's law

$$S_g = k_H P_g$$

where S_g is gas solubility, k_H is Henry's law constant, and P_g is the partial pressure of O_2. Substituting Henry's law constant for O_2 and the partial pressure into the equation will yield the solubility of O_2 in water.

$$S_g = \left(\frac{1.66 \times 10^{-6} \text{ M}}{\text{mm Hg}}\right)\left(35 \text{ mm Hg}\right) = 5.8 \times 10^{-5} \text{ M}$$

7. Solubility and Henry's Law

Question

What is the concentration of O_2 (in grams of O_2 per liter of water) in a freshwater stream in equilibrium with air at 25 °C? The atmospheric pressure is 1.0 atm and Henry's law constant for O_2 is 1.66 x 10^{-6} M/mm Hg at 25 °C. Assume air contains 21% oxygen.

Approach

Henry's Law can be used to calculate the solubility of oxygen. First, calculate the partial pressure of oxygen in air (21% of air is oxygen and the mole fraction of O_2 is 0.21). Then, calculate the solubility of oxygen using Henry's law.

Solution

Step 1. Calculate the partial pressure of oxygen.

$$P(O_2) = (1.0\ \text{atm})\left(\frac{760\ \text{mm Hg}}{1\ \text{atm}}\right)(0.21) = 160\ \text{mm Hg}$$

Step 2. Calculate the solubility of oxygen in units of grams of oxygen per liter of water.

$$\left(\frac{1.66 \times 10^{-6}\ \text{mol / L}}{1\ \text{mm Hg}}\right)(160\ \text{mm Hg}) = 2.7 \times 10^{-4}\ \text{mol / L}$$

$(2.7 \text{ x } 10^{-4}\ \text{mol/L})(32.\ \text{g/mol}) = 0.0085\ \text{g/L}$

8. Vapor Pressure and Raoult's Law

Question

What is the vapor pressure at 25 °C of a benzene-toluene solution of composition $X_{benz} = 0.40$ and $X_{tol} = 0.60$? The vapor pressures of the pure substances are 73 Torr for benzene, C_6H_6, and 27 Torr for toluene, C_7H_8, at 25 °C. Assume that the benzene and toluene form an ideal solution.

Approach

According to Raoult's law, the partial pressure of each component in the vapor phase is directly proportional to its mole fraction in the solution:

$P_A = X_A P°_A$

The total vapor pressure of the solution is the sum of the partial pressures.

Solution

Step 1. Calculate the partial pressures of benzene and toluene.

$P_{benz} = X_{benz}P°_{benz} = (0.40)(73\ \text{Torr}) = 29\ \text{Torr}$

$P_{tol} = X_{tol}P°_{tol} = (0.60)(27\ \text{Torr}) = 16\ \text{Torr}$

Step 2. Calculate the total vapor pressure of the solution.

$P_{soln} = P_{benz} + P_{tol} = 29\ \text{Torr} + 16\ \text{Torr} = 45\ \text{Torr}$

9. Colligative Properties: Boiling Point Elevation and Freezing Point Depression

Question

A solution contains a mixture of 5 sugars: 0.50 mol fructose, 0.60 mol glucose, 0.50 mol lactose, 0.60 mol maltose, and 0.60 mol sucrose dissolved in 1.00 kg of water. What are the boiling point and the freezing point of this solution? The boiling point elevation constant for water is 0.512 °C / m, and the freezing point depression constant for water is -1.86 °C / m.

Approach

We need to first calculate the total number of moles of solute in the solution. Then, we will need to use the following relationship to calculate the boiling point elevation:

$\Delta T_{bp} = K_{bp} * m_{solute}$.

Similarly, for freezing point depression, we will need to use the relationship:

$\Delta T_{fp} = K_{fp} * m_{solute}$.

The solution's boiling point can be calculated by adding the change in temperature found to the boiling point of water. The solution's freezing point can be calculated by subtracting the change in temperature found from the freezing point of water.

Solution

Step 1. Calculate the total number of moles of solute in the solution and the total concentration of solute.

0.50 mol fructose + 0.60 mol glucose + 0.50 mol lactose + 0.60 mol maltose + 0.60 mol sucrose = 2.80 mol solute

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kg of solvent}} = \frac{2.80\text{ mol}}{1.00\text{ kg H}_2\text{O}} = 2.80\text{ mol/kg}$$

Step 2. Calculate the boiling point elevation of the solution and the boiling point of the solution.

$\Delta T_{bp} = K_{bp} * m_{solute} = (0.512\ °C * kg/mol)(2.80\ mol/kg) = 1.43\ °C$

$T_{bp,\ solution} = T_{bp,\ solvent} + \Delta T_{bp} = 100.00\ °C + 1.43\ °C = 101.43\ °C$

Step 3. Calculate the freezing point depression of the solution and the freezing point of the solution.

$\Delta T_{fp} = K_{fp} * m_{solute} = (1.86\ °C *kg/mol)(2.80\ mol/kg) = -5.21\ °C$

$T_{fp,\ solution} = T_{fp,\ solvent} - \Delta T_{fp} = 0.00\ °C - 5.21\ °C = -5.21\ °C$

10. Osmotic Pressure

Problem

An aqueous solution contains 77.1 g of insulin $(C_6H_{10}O_5)_x$, a high molecular mass sugar, per liter of solution. The osmotic pressure at 20 °C of this solution is 0.58 atm. Calculate the molar mass of insulin.

Approach

The molarity of the solution can be calculated from the osmotic pressure, using the following relationship: $\pi = cRT$, where π is the osmotic pressure, c is concentration (in moles per liter), R is the gas constant, and T is the absolute temperature. Then, the molar mass can be found from the mass and molarity.

Solution

Step 1. Calculate the solution concentration.

$$\pi = cRT$$

$$c = \frac{0.58 \text{ atm}}{(0.0821 \text{ L atm/K mol})(293 \text{ K})} = 0.024 \text{ mol/L}$$

Step 2. Calculate the molar mass of insulin.

$$\left(\frac{77.1 \text{ g}}{1 \text{ L}}\right)\left(\frac{1 \text{ L}}{0.024 \text{ mol}}\right) = 3200 \text{ g/mol}$$

CH 222 Chapter Twelve Concept Guide

1. Reaction Rate

Problem

Express the rate of reaction for the following reaction:

$H_2(g) + I_2(g) \rightarrow 2HI(g)$

Approach

The rate of reaction is defined as the change in either reactant or product concentration divided by the change in time.

Solution

In terms of reactants, the rate of reaction *could* be:

$$\text{Rate} = -\frac{\Delta[H_2]}{\Delta t} = -\frac{\Delta[I_2]}{\Delta t}$$

One molecule of H_2 and one molecule of I_2 react with each other, thus the concentration of these two reactants changes at the same rate.

In terms of product, the rate of reaction *could* be:

$$\text{Rate} = \frac{1}{2}\frac{\Delta[HI]}{\Delta t}$$

Two molecules of HI are produced for every one molecule of H_2 and one molecule of I_2 that react. Thus, the change in concentration of HI will be twice as large as the rate of reaction defined in terms of H_2 or I_2.

2. Reaction Rate

Problem

Define the rate of reaction for the following equation:

$2\,O_3(g) \rightarrow 3\,O_2(g)$

Approach

The rate of reaction is defined as the difference in either reactant or product concentration divided by the change in time.

Solution

In terms of the reactant, the rate of reaction *could* be:

$$\text{Rate} = -\frac{1}{2}\frac{\Delta[O_3]}{\Delta t}$$

In terms of product, the rate of reaction *could* be:

$$\text{Rate} = +\frac{1}{3}\frac{\Delta[O_2]}{\Delta t}$$

The relationship between the rate of O_3 consumption and O_2 production is:

$$\frac{\Delta[O_3]}{\Delta t} = -\frac{2}{3}\frac{\Delta[O_2]}{\Delta t}$$

3. Reaction Order

Question

The rate equation for the following reaction in the presence of platinum is as follows:

$2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$

Rate = $k[SO_2][SO_3]^{-1/2}$

What is the order of the reaction with respect to each species in the rate equation? What is the reaction order?

Approach

The order of the reaction for each substance is given by the exponent on the concentration term in the rate equation.

Solution

The reaction is first order with respect to SO_2 and one-half order with respect to SO_3. The overall reaction order is **$-^{1}/_{2}$**.

4. Reaction Order

Question

The iodide-catalyzed decomposition of hydrogen peroxide, H_2O_2, is known to be first order with respect to H_2O_2 and first order with respect to I^-.

$$2\ H_2O_2(aq) \xrightarrow{I^-} 2\ H_2O(\ell) + O_2(g)$$

Write the rate equation for this reaction. What is the overall order of this reaction?

Solution

The rate equation is: Rate = $k[H_2O_2][I^-]$. The overall order of this reaction is the sum of the exponents of the concentration terms in the rate equation. This reaction is second order overall.

5. Method of Initial Rates

Problem

The rate equation for the reaction of iodide ion with hypochlorite ion in an alkaline solution

$$I^- + OCl^- \xrightarrow{OH^-} Cl^- + IO^-$$

was thought to be in the form of the following rate equation:

Rate = $k[I^-]^x[OCl^-]^y[OH^-]^z$

Using the method of initial rates, the following data were collected at 25 °C:

Experiment	**$[ClO^-]_0$**	**$[I^-]_0$**	**$[OH^-]_0$**	**Initial Rate (mol/L s)**
(a)	2.00×10^{-3}	2.00×10^{-3}	1.00	2.42×10^{-4}
(b)	4.00×10^{-3}	2.00×10^{-3}	1.00	4.82×10^{-4}
(c)	2.00×10^{-3}	4.00×10^{-3}	1.00	5.02×10^{-4}
(d)	2.00×10^{-3}	2.00×10^{-3}	0.500	4.64×10^{-4}

Determine the order of the reaction with respect to OCl^-, I^-, and OH^-, then calculate the rate constant k.

Solution

Step 1. Find two experiments where different concentrations of I^- are used but $[OH^-]$ and $[ClO^-]$ are the same.

The data for experiments (a) and (c) show that the reaction rate doubles as the concentration of I^- doubles. Thus, the reaction is first order with respect to $[I^-]$ and x = 1. Mathematically, this may be shown as:

$$\text{Rate} = \frac{5.02 \times 10^{-4}}{2.42 \times 10^{-4}} = \frac{k(2.00 \times 10^{-3})^x(4.00 \times 10^{-3})^y(1.00)^z}{k(2.00 \times 10^{-3})^x(2.00 \times 10^{-3})^y(1.00)^z}$$

$$\text{Rate} = 2.07 \approx 2.00$$

Thus, by inspection, y is 1. This may also be shown mathematically: ln 2 = y ln 2, thus y = 1.

Step 2. Find two experiments where different ClO^- concentrations are used but $[I^-]$ and $[OH^-]$ are the same.

The data for experiments (a) and (b) show that the reaction rate doubles as the concentration of ClO^- doubles. Thus, the reaction is first order with respect to $[OCl^-]$ and y = 1.

Step 3. Find two experiments where different OH^- concentrations are used but $[OCl^-]$ and $[I^-]$ are the same.

The data for experiments (a) and (d) show that the reaction rate doubles as the concentration of OH^- is halved. Thus, the reaction is negative first order with respect to $[OH^-]$ and z = -1. The reaction rate decreases with increasing $[OH^-]$.

The complete rate of reaction is:

Rate = $k[I^-][OCl^-][OH^-]^{-1}$

Step 4. The rate constant can be calculated by rearranging the rate equation and choosing data from one experiment.

Using the data from experiment (a):

$$k = \frac{\text{rate}[OH^-]}{[I^-][OCl^-]}$$

$$= \frac{(2.42 \times 10^{-4}\ \text{mol / Ls})(1.00\ \text{mol/ L})}{(2.00 \times 10^{-3}\ \text{mol / L})(2.00 \times 10^{-3}\ \text{mol / L})}$$

$$= 60.5\ s^{-1}$$

Using data from experiment (b), $k = 60.3\ s^{-1}$; from (c), $k = 62.8\ s^{-1}$; and from (d), $k = 58.0\ s^{-1}$. The average value of $k = 60.4\ s^{-1}$.

6. Half-life

Question

Compound A decomposes to form B and C in a reaction that is first order with respect to A and first order overall. At 25 °C, $k = 0.0470\ s^{-1}$. What is the half-life for A at 25 °C?

Approach

We will need to use the equation for $t_{1/2}$:

$$t_{1/2} = 0.693 / k$$

Solution

$$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{0.0470\ s^{-1}} = 14.7\ s$$

Half of the original reactant remains after 14.7 s of reaction.

7. Reaction Mechanisms

Question

The proposed mechanism for a reaction is:

$$\text{Step 1: } A + B \underset{k2}{\overset{k1}{\rightleftarrows}} C + D$$

$$\text{Step 2: } C + A \underset{k2}{\overset{k1}{\rightleftarrows}} E$$

$$\text{Step 3: } E + B \underset{k2}{\overset{k1}{\rightleftarrows}} F$$

What is the overall reaction?

Approach

We need to add the reactants on the left side of each arrow and add the products on the right side of each arrow, canceling out like terms to determine the overall reaction.

Solution

The overall reaction is: $2A + 2B = D + F$.

8. Determining Rate Laws

Question

Refer to question 7. What is the experimental rate law if the first step is the slow step?

Solution

The experimental rate law, contingent upon the first step being the slow step, is: $R = k[A][B]$.

9. Catalysts

Question

One pathway for the destruction of ozone in the upper atmosphere is:

$$\begin{array}{ll} O_3(g) + NO(g) \longrightarrow NO_2(g) + O_2(g) & \text{slow} \\ NO_2(g) + O(g) \longrightarrow NO(g) + O_2(g) & \text{fast} \\ \hline O_3(g) + O(g) \longrightarrow 2\,O_2(g) & \end{array}$$

Which species is the catalyst and which species is the intermediate?

Solution

The catalyst is a species that increases the rate of a chemical reaction without itself undergoing any permanent change. It is involved in one step of the mechanism and regenerated at a later step. It is not a reactant or product in the overall reaction. In this example, the catalyst is NO. The intermediate is a species that is produced during the course of a reaction but always reacts further and is not among the final products. Here, the intermediate is NO_2.

10. Determining Activation Energy

Question

A plot of $\ln k \text{ vs. } \frac{1}{T}$ gives a slope of -1.50 x 10^4 K. What is the activation energy for this particular reaction?

Approach

We will need to use the following equation:

$$\ln k = \ln A - \left(\frac{E_a}{R}\right)\left(\frac{1}{T}\right)$$

When plotting ln k vs. (1/T), the slope of the line is equal to $-\frac{E_a}{R}$

Solution

$$\ln k = \ln A - \left(\frac{E_a}{R}\right)\left(\frac{1}{T}\right)$$

E_a = -(-1.50 x 10^4 K)(8.314 J/K) = 125000 J = 125 kJ

11. Activation Energy

Question

Do reactions with larger activation energies have smaller or larger values of k, and as a result, are these reactions slow or fast?

Solution

Large values of activation energy lead to small values of k. These reactions are slower, as we would expect from our understanding of effective collisions and energy relationships. With a large activation energy, only a small fraction of the collisions between molecules lead to products.

12. Activation Energy and Temperature

Question

The rate constant for the following reaction

$2\ N_2O_5 \rightarrow 4\ NO_2 + O_2$

was measured at two different temperatures: at $T_1 = 25.0\ °C$, $k_1 = 5.60 \times 10^{-5}\ s^{-1}$, and at $T_2 = 67.0\ °C$, $k_2 = 9.32 \times 10^{-3}\ s^{-1}$. What is the activation energy for this reaction?

Approach

The solution to this problem requires solving for E_a using the following equation and substituting in k_1, k_2, T_1, and T_2:

$$\ln \frac{k_2}{k_1} = -\left(\frac{E_a}{R}\right)\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

Solution

$$\left(\frac{-E_a}{8.314\ J/\ K \cdot mol}\right)\left(\frac{1}{340.0} - \frac{1}{298.0}\right) = \ln \frac{9.32 \times 10^{-3}}{5.60 \times 10^{-5}}$$

$E_a = 1.0 \times 10^5$ J/mol

CH 222 Chapter Twenty-one Concept Guide

1. Terminology

Alpha Radiation (α):	Alpha particles are positively charged particles ejected at high speed from certain radioactive substances; a helium nucleus
Beta Radiation (β):	Beta particles are electrons that are ejected at high speed from certain radioactive substances
Gamma Radiation (γ):	High-energy electromagnetic radiation
Nuclear Reaction:	A reaction involving one or more atomic nuclei, resulting in a change in the identities of the isotopes
Nucleons:	A nuclear particle, either a neutron or a proton
Radioactive Decay Series:	A series of nuclear reactions by which a radioactive isotopes decays to form a stable isotope
Positrons:	A nuclear particle having the same mass as an electron but a positive charge
Nuclear Binding Energy:	The energy required to separate a nucleus into its individual nucleons
Fission:	The highly exothermic process by which very heavy nuclei split to form lighter nuclei
Fusion:	the state change from solid to liquid
Half-life:	The time required for the concentration of one of the reactants to reach half of its initial value
Activity (A):	A measure of the rate of nuclear decay, the number of disintegrations observed in a sample per unit cell
Nuclear Reactor:	A container in which a controlled nuclear reaction occurs
Nuclear Fusion:	The highly exothermic process by which comparatively light nuclei combine to form heavier nuclei
Plasma:	A gas like phase of matter that consists of charged particles

Röntgen	A unit of radiation dosage proportional to the amount of ionization produced in air
Rad:	A unit of radiation dosage which measures the radiation dose to living tissue
Rem:	A unit of radiation dosage which takes into account the differing intensities of different radiation (alpha, beta and gamma) upon human tissue
Curie (Ci):	A unit of radioactivity which measures activity. One curie represents any radioactive isotope which undergoes 3.7 x 10^{10} disintegrations per second (dps).

2. α Radioactive Decay Series

Problem

1. A radioactive decay series begins with $^{235}_{92}U$ and ends with $^{207}_{82}Pb$. What is the total number of α and β particles emitted in this series?
2. The first three steps of this series involve (in order) α, β, and α emissions. Write the nuclear equations for each of these steps.

Solution

1. Mass declines by 28 mass units (235-207) in this series. Because a decrease in mass can only occur with α emission, we conclude that seven α particles must be emitted. For each α emission, the atomic number must decrease by 2, so emission of seven α particles causes the atomic number to decrease by 14. The actual decrease in atomic number is 10, however (92-82). Four β particles cause the atomic number to increase by 4. This radioactive decay sequence involves loss of seven α and four β particles.

2. Step 1. $^{235}_{92}U \rightarrow {}^{231}_{90}Th + {}^{4}_{2}\alpha$

Step 2. $^{231}_{90}{}^{Th} \rightarrow {}^{231}_{91}Pa + {}^{0}_{-1}\beta$

Step 3 $^{231}_{91}Pa \rightarrow {}^{227}_{89}Ac + {}^{4}_{2}\alpha$

3. Balancing Nuclear Reaction

Problem

Complete the following equations. Give the symbol, mass number, and atomic number of the species indicated by the question mark.

1. ${}^{37}_{18}\mathrm{Ar} + {-}{}^{0}_{1}\mathrm{e} \rightarrow ?$
2. ${}^{11}_{6}\mathrm{C} \rightarrow {}^{11}_{5}\mathrm{B} + ?$
3. ${}^{35}_{16}\mathrm{S} \rightarrow {}^{35}_{17}\mathrm{Cl} + ?$
4. ${}^{30}_{15}\mathrm{P} \rightarrow +{}^{0}_{1}\beta + ?$

Solution

1. This is an electron capture reaction. The product has a mass number of 37+0 = 37 and an atomic number of 18-1 = 17. The symbol for the product is $\mathbf{{}^{37}_{17}Cl}$.
2. This reaction is recognized as positron (${}^{0}_{+1}\beta$) emission. By choosing this particle, the sum of the atomic numbers (6 = 5+1) and the mass numbers (11) on either side of the reaction are equal.
3. Beta (${}^{0}_{-1}\beta$) emission is required to balance the mass numbers (35) and atomic numbers (16 = 17-1) on both sides of the equation.
4. The product nucleus is $\mathbf{{}^{30}_{14}Si}$. This balances the mass numbers (30) and atomic numbers (15 = 1+14) on both sides of the equation.

4. Binding Energy

Approach

Einstein's equation from the theory of special relativity states that the energy of a body is equivalent to the mass times the speed of light squared

$$\Delta E = (\Delta m)c^2$$

When comparing nuclear stabilities, scientists generally calculate the binding energy (E_b) per nucleon:

$$\frac{E_b}{\text{Mol nucleons}}$$

where the number of nucleons equals the number of protons plus the number of neutrons in an atom. The binding energy is related to the change in energy by

$$\Delta E = -E_b$$

Problem

Calculate the binding energy (in kJ/mole) and the binding energy per nucleon (in kJ/mole nucleons) for carbon-12.

Solution

The following reaction results in formation of carbon-12:

$$6\ {}^{1}_{1}H + 6\ {}^{1}_{0}n \rightarrow {}^{12}_{6}C$$

The mass of ${}^{1}_{1}H$ is 1.00783 g/mol and the mass of ${}^{1}_{0}n$ is 1.00867 g/mole. Carbon-12, ${}^{12}_{6}C$, is the standard for the atomic masses in the periodic table, and its mass is defined as exactly 12.000 g/mol.

To determine binding energy we must first determine the difference in mass of the products and reactants in this reaction:

$$\begin{aligned}\Delta m &= 12.0000000 - [(6 \times 1.00783) + (6 \times 1.00867)] \\ &= -9.9000 \times 10^{-2} \text{ g/mol}\end{aligned}$$

The binding energy is calculated using $\Delta E = (\Delta m)c^2$.

Using the **mass** in **kilograms** and the **speed of light** in **meters per second** gives an answer for the binding energy in joules:

$$\begin{aligned}E_b &= -(\Delta m)c^2 = -(9.9 \times 10^{-5} \text{ kg/mol})(3.00 \times 10^8 \text{ m/s})^2 \\ &= 8.91 \times 10^{12} \text{ J/mol} = 8.91 \times 10^9 \text{ kJ/mol}\end{aligned}$$

The **binding energy per nucleon** is determined by dividing the binding energy by the number of nucleons, which in this instance is 12.

$$\frac{E_b}{\text{mol nucleons}} = \frac{8.91 \times 10^9 \text{ kJ mol}^{-1}}{12 \text{ nucleons mol}^{-1}}$$

$$= \mathbf{7.43 \times 10^8 \text{ kJ/mol nucleons}}$$

5. Rate of Radioactive Decay and Half-Life

Approach

Radioactive decay processes always follow first order kinetics. The activity (A) of a nuclear decay process is proportional to the number of radioactive atoms present (N), or

$$A/A_0 = k(N/N_0)$$

where k equals the decay or rate constant. The first order integrated rate law equation relates the period over which a sample is observed (t) to the fraction of radioactive atoms present after that amount of time has passed

$$\ln N/N_0 = -kt$$

Another convenient method to find the decay constant is through the **half-life**, $t_{1/2}$. Half-life is defined as the time required for the concentration of a reactant to reach half of its initial value (i.e. $N/N_0 = 0.5$). An equation to determine the half-life from the rate constant can be derived from the first order integrated rate law; the results can be expressed as follow:

$$t_{1/2} = 0.693 / k$$

Example

Some high-level radioactive waste with a half-life $t_{1/2}$ of 200.0 years is stored in underground tanks. What time is required to reduce an activity of 6.50×10^{12} disintegrations per minute (dpm) to a fairly harmless activity of 3.00×10^{-3} dpm?

Solution

The data provides the initial activity ($A_o = 6.50 \times 10^{12}$ dpm) and the activity after some elapsed time ($A = 3.00 \times 10^{-3}$ dpm). To find the elapsed time t, first find k from the half-life:

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{200.\text{ years}} = 0.00347 \text{ year}^{-1}$$

With k known, the time t can be calculated:

$$\ln\left[\frac{3.00 \times 10^{-3}}{6.50 \times 10^{12}}\right] = -[0.00347 \text{ year}^{-1}]t$$

$$-35.312 = [0.00347 \text{ year}^{-1}]t$$

$$t = \frac{-35.312}{-[0.00347 \text{ year}^{-1}]}$$

$$= 1.02 \times 10^{4} \text{ years}$$

Note that the time unit t and rate constant k must share common units (i.e. years and years^{-1}) for this equation to work properly.

CH 222 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 Seven** *and* **Chapter Guide One**

*Important Tables and/or Constants: "***Bond Enthalpies and Electronegativities***" Handout (*after problem set #1) and "**Geometry and Polarity Guide**" (summary of molecular geometries and polarities *Handout* in *Companion)*

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1. Give the periodic group number and number of valence electrons for each of the following atoms.
   a. O b. B c. Na d. Mg e. F f. S
2. Which of the following elements are capable of forming compounds in which the indicated atom has more than four valence electron pairs?
   a. N b. As c. C d. O e. Br f. Be g. S h. Se
3. Which compound in each of the following pairs should require the higher temperature to melt?
   a. KBr or CsBr
   b. SrS or CaS
   c. LiF or BeO
4. Draw a Lewis structure for each of the following molecules.
   a. NF_3 b. $CHClF_2$ c. HOBr d. CH_3CN
5. Draw a Lewis structure for each of the following molecules or ions.
   a. BrF_3 b. I_3^{1-} c. XeO_2F_2 d. XeF_3^{1+}
6. Draw a Lewis structure for each of the following molecules or ions. Describe the electron-pair geometry and the molecular geometry around the central atom.
   a. NH_2Cl b. Cl_2O (O is the central atom) c. SCN^{1-} (C is the central atom) d. HOF
7. Draw a Lewis structure for each of the following molecules or ions. Describe the electron-pair geometry and the molecular geometry around the central atom.
   a. ClF_2^{1-} b. ClF_3 c. ClF_4^{1-} d. ClF_5
8. Give approximate values for the indicated bond angles.
   a. O-S-O in SO_2
   b. F-B-F angle in BF_3
   c. Cl-C-Cl angle in Cl_2CO
9. Determine the formal charge on each atom in the following molecules and ions.
   a. NO_2^{+1} b. NO_2^{1-} c. NF_3 d. HNO_3
10. For each of the bonds below, Tell which atom is the more negatively charged using values of electronegativity in your textbook to support your answer..
    a. C-O b. P-Cl c. B-O d. B-F
11. In the following list of compounds, which bond is the most polar? Which compound(s) are nonpolar? Which atom in ClF is more negatively charged?
    a. H_2O b. NH_3 c. CO_2 d. ClF e. CCl_4

12. Three resonance structures are possible for dinitrogen monoxide, N_2O.
    a. Draw the three resonance structures.
    b. Calculate the formal charge on each atom in each resonance structure.
    c. Based on formal charges and electronegativity, predict which resonance structure is the most reasonable.
13. Give the bond order for each bond in the following molecules or ions.
    a. CH_2O b. CO_2 c. NO_2^{1+} d. CH_4
14. The compound oxygen difluoride is quite reactive, giving oxygen and HF when treated with water:

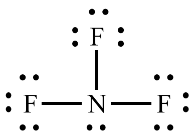
    $$OF_2(g) + H_2O(g) \rightarrow O_2(g) + 2\ HF(g) \quad \Delta H^\circ_{rxn} = -318\ kJ$$

    Using bond energies, calculate the bond dissociation energy of the O-F bond in OF_2.
15. Nitric acid, HNO_3, has three resonance structures. One of them, however, contributes much less to the resonance hybrid than the other two. Sketch the three resonance structures and assign a formal charge to each atom. Which one of your structures is the least important?
16. Draw the Lewis structure of ammonium nitrate.

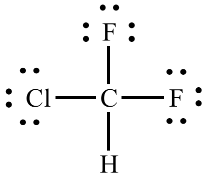
Answers to the Practice Problem Set:

1. *Answers:*
   a. O Group 6A 6 valence electrons
   b. B Group 3A 3 valence electrons
   c. Na Group 1A 1 valence electron
   d. Mg Group 2A 2 valence electrons
   e. F Group 7A 7 valence electrons
   f. S Group 6A 6 valence electrons
2. b. As e. Br g. S h. Se (all third period or lower)
3. a. KBr b. CaS c. BeO
4. *Answers:*

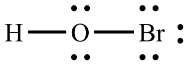
a.



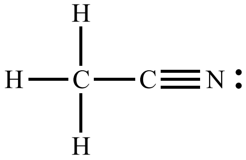
b.



c.

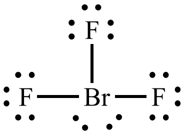


d.

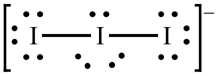


5. *Answers:*

a.

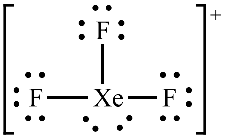


b.



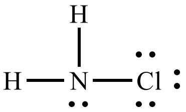
c.

d.



6. *Answers:*

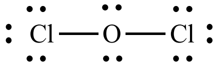
a.



electron pair geometry, tetrahedral

molecular geometry, trigonal pyramidal

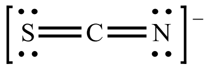
b.



electron pair geometry, tetrahedral

molecular geometry, bent

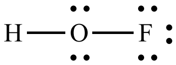
c.



electron pair geometry, linear

molecular geometry, linear

d.

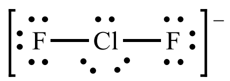


electron pair geometry, tetrahedral

molecular geometry, bent

7. *Answers:*

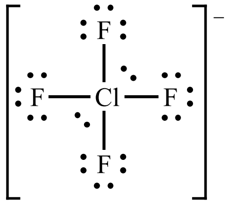
a.



electron-pair geometry, trigonal bipyramid

molecular geometry, linear

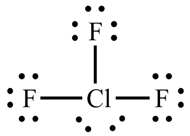
c.



electron-pair geometry, octahedral

molecular geometry, square planar

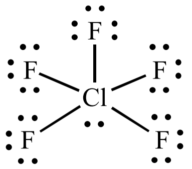
b.



electron-pair geometry = trigonal bipyramid

molecular geometry, T-shaped

d.



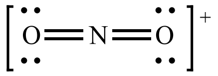
electron-pair geometry, octahedral

molecular geometry, square pyramidal

8. a. 120° b. 120° c. 120°

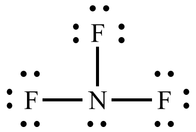
9. *Answers:*

a.



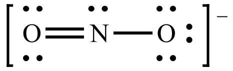
$O = 6 - 4 - \frac{1}{2}(4) = 0$
$N = 5 - 0 - \frac{1}{2}(8) = 1$

c.



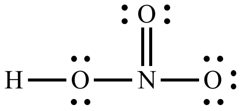
$F = 7 - 6 - \frac{1}{2}(2) = 0$
$N = 5 - 2 - \frac{1}{2}(6) = 0$

b.



$O = 6 - 4 - \frac{1}{2}(4) = 0$
$N = 5 - 2 - \frac{1}{2}(6) = 0$
$O = 6 - 6 - \frac{1}{2}(2) = -1$

d.



$H = 1 - 0 - \frac{1}{2}(2) = 0$
$O = 6 - 4 - \frac{1}{2}(4) = 0$
$N = 5 - 0 - \frac{1}{2}(8) = 1$
$O = 6 - 4 - \frac{1}{2}(4) = 0$
$O = 6 - 6 - \frac{1}{2}(2) = -1$

10. *Answers:*

a.	C—O	O	2.5 - 3.5
b.	P—Cl	Cl	2.2 - 3.2
c.	B—O	O	2.9 - 3.5
d.	B—F	F	2.0 - 4.0

11. Answers:

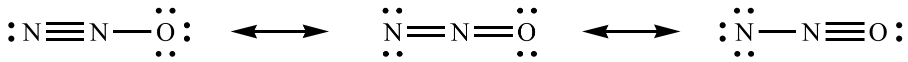
Molecule	$\Delta\chi$ for bond
H_2O	O—H = 3.5 – 2.1 = 1.4
NH_3	N—H = 3.0 – 2.1 = 0.9
CO_2	O—C = 3.5 – 2.5 = 1.0
ClF	F—Cl = 4.0 – 3.0 = 1.0
CCl_4	Cl—C = 3.0 – 2.5 = 0.5

(i) The bonds are most polar in H_2O (biggest $\Delta\chi$)

(ii) CO_2 and CCl_4 are nonpolar molecules.

(iii) The F atom in ClF is more negatively charged.

12. Answers:

a. 

b.

$N = 5 - 2 - {}^{1}/_{2}(6) = 0$	$N = 5 - 4 - {}^{1}/_{2}(4) = -1$	$N = 5 - 6 - {}^{1}/_{2}(2) = -2$
$N = 5 - 0 - {}^{1}/_{2}(8) = 1$	$N = 5 - 0 - {}^{1}/_{2}(8) = 1$	$N = 5 - 0 - {}^{1}/_{2}(8) = 1$
$O = 6 - 6 - {}^{1}/_{2}(2) = -1$	$O = 6 - 4 - {}^{1}/_{2}(4) = 0$	$O = 6 - 2 - {}^{1}/_{2}(6) = 1$

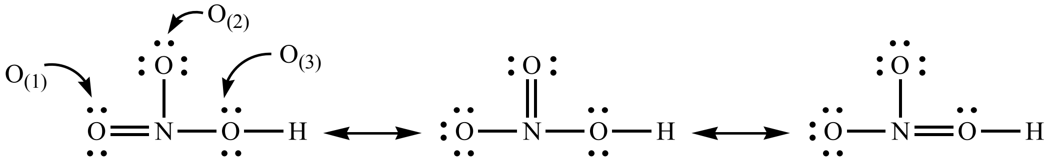
c. The first resonance structure is most reasonable (the most electronegative element, oxygen, has a negative formal charge).

13. Answers:

a.	H_2CO	two carbon-hydrogen single bonds	bond order = 1
		one carbon-oxygen double bond	bond order = 2
b.	CO_2	two carbon-oxygen double bonds	bond order = 2
c.	NO_2^+	two nitrogen-oxygen double bonds	bond order = 2
d.	CH_4	four carbon-hydrogen single bonds	bond order = 1

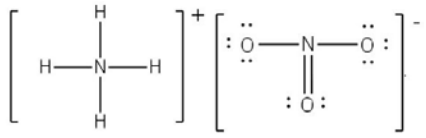
14. 195 kJ/mol

15. Answers:



$O_{(1)} = 0$	$O_{(1)} = -1$	$O_{(1)} = -1$
$O_{(2)} = -1$	$O_{(2)} = 0$	$O_{(2)} = -1$
$O_{(3)} = 0$	$O_{(3)} = 0$	$O_{(3)} = 1$
H = 0	H = 0	H = 0
N = 1	N = 1	N = 1

The third resonance structure is the least important since it has a positive formal charge on one of the oxygen atoms.

16. 

CH 222 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 Eight** *and* **Chapter Guide Two**

Important Tables and/or Constants: "**MO Diagram for B_2, C_2, and N_2**" *(Handout)*, "**MO Diagram for O_2, F_2, and Ne_2**" *(Handout)*, "**Geometry and Polarity Guide**" *(Handout)*

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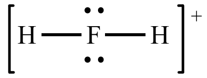
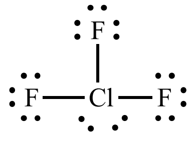
1. Draw the Lewis structure for chloroform, $CHCl_3$. What are its electron-pair and molecular geometries? What orbitals on C, H, and Cl overlap to form bonds involving these elements?
2. Specify the electron-pair and molecular geometry for each of the following. Describe the hybrid orbital set used by the central atom in each molecule or ion.
   a. BBr_3 b. CO_2 c. CH_2Cl_2 d. CO_3^{2-}
3. Draw the Lewis structure and then specify the electron- pair and molecular geometries for each of the following molecules or ions. Identify the hybridization of the central atom.
   a. SiF_6^{2-} b. SeF_4 c. ClF_2^{1-} d. XeF_4
4. The hydrogen molecular ion, H_2^+, can be detected spectroscopically. Write the electron configuration of the ion in molecular orbital terms. What is the bond order of the ion? Is the hydrogen–hydrogen bond stronger or weaker in H_2^+ than in H_2?
5. Calcium carbide, CaC_2, contains the acetylide ion, C_2^{2-}. Sketch the molecular orbital energy level diagram for the ion. How many net σ and π bonds does the ion have? What is the carbon–carbon bond order? How has the bond order changed on adding electrons to C_2 to obtain C_2^{2-}? Is the C_2^{2-} ion paramagnetic?
6. The simple valence bond picture of O_2 does not agree with the molecular orbital view. Compare these two theories with regard to the peroxide ion, O_2^{2-}.
   (a) Draw an electron dot structure for O_2^{2-}. What is the bond order of the ion?
   (b) Write the molecular orbital electron configuration for O_2^{2-}. What is the bond order based on this approach?
   (c) Do the two theories of bonding lead to the same magnetic character and bond order for O_2^{2-}?
7. Which of the homonuclear, diatomic molecules of the second-period elements (from Li_2 to Ne_2) are paramagnetic? Which have a bond order of 1? Which have a bond order of 2? Which diatomic molecule has the highest bond order?
8. Consider the three fluorides BF_4^{-1}, SiF_4, and SF_4.
   a. Identify a molecule that is isoelectronic with BF_4^{1-}.
   b. Are SiF_4 and SF_4 isoelectronic?
   c. What is the hybridization of the central atom in each of these species?
9. When is it desirable to use MO theory rather than valence bond theory?
10. Antimony pentafluoride reacts with HF according to the equation:
    $$2\ HF + SbF_5 \rightarrow [H_2F]^+[SbF_6]^-$$
    a. What is the hybridization of the Sb atom in the reactant and product?
    b. Draw a Lewis structure for H_2F. What is the geometry of H_2F^+? What is the hybridization of F in H_2F^+?

11. The CN molecule has been found in interstellar space. Using the appropriate molecular orbital energy level diagram, answer the following questions.
    a. What is the highest energy occupied molecular orbital (HOMO) to which an electron (or electrons) is (are) assigned?
    b. What is the bond order of the molecule?
    c. How many net σ bonds are there? How many net π bonds?
    d. Is the molecule paramagnetic or diamagnetic?
12. Draw the Lewis structure for ClF_3. What are its electron- pair and molecular geometries? What is the hybridization of the chlorine atom? What orbitals on Cl and F overlap to form bonds between these elements?

Answers to the Practice Problem Set:

1. The electron-pair and molecular geometries are tetrahedral. The C atom is *sp*^{3} hybridized. Three of these hybrid orbitals each overlap with a chlorine 3*p* orbital to form three C—Cl sigma bonds. One hybrid orbital overlaps with a hydrogen 1*s* orbital to from a C—H sigma bond.
2. *Answers:*

(a) BBr_3	trigonal planar	trigonal planar	*sp*^{2}
(b) CO_2	linear	linear	*sp*
(c) CH_2Cl_2	tetrahedral	tetrahedral	*sp*^{3}
(d) CO_3^{2-}	trigonal planar	trigonal planar	*sp*^{2}

3. a) octahedral, octahedral, *sp*^{3}*d*^{2} b) trigonal bipyramid, see-saw, *sp*^{3}*d* c) trigonal bipyramid, linear, *sp*^{3}*d* d) octahedral, square planar, *sp*^{3}*d*^{2}
4. H_2^+: $(\sigma_{1s})^1$ Bond order = $\frac{1}{2}(1 - 0) = \frac{1}{2}$, weaker H—H bond H_2: $(\sigma_{1s})^2$ Bond order = $\frac{1}{2}(2 - 0) = 1$, stronger H—H bond
5. C_2^{2-}: $(\sigma_{1s})^2(\sigma^*_{1s})^2(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^4(\sigma_{2p})^2$ C_2^{2-} ion has a bond order of $\frac{1}{2}(8 - 2) = 3$ (one σ bond and two π bonds). The C_2 molecule has two fewer electrons and a bond order of $\frac{1}{2}(6 - 2) = 2$. The C_2^{2-} ion is diamagnetic.
6. (a) bond order = 1 (b) [core electrons]$(\sigma_{2s})^2(\sigma^*_{2s})^2(\sigma_{2p})^2(\pi_{2p})^4(\pi^*_{2p})^4$ bond order = $\frac{1}{2}(8 - 6) = 1$ (c) The theories agree here.
7. B_2 and O_2 are paramagnetic, Li_2, B_2, and F_2 have a bond order of 1, C_2 and O_2 have a bond order of 2, and N_2 has the highest bond order, 3.
8. (a) CF_4 is isoelectronic with BF_4^- (32 valence electrons)
   (b) SiF_4 (32 valence electrons) and SF_4 (34 valence electrons) are not isoelectronic
   (c) BF_4^-: *sp*^{3} SiF_4: *sp*^{3} SF_4: *sp*^{3}*d*
9. Molecular orbital theory correctly predicts the electronic structures for odd-electron molecules and other molecules such as O_2 that do not follow the electron-pairing assumptions of the Lewis dot structure approach.
10. (a) *sp*^{3}*d* in SbF_5, *sp*^{3}*d*^{2} in SbF_6^-
    (b) 
    $[H—\ddot{\underset{..}{F}}—H]^+$
    The geometry of H_2F^+ is bent, and the F atom is *sp*^{3} hybridized.
11. CN [core electrons]$(\sigma_{2s})^2(\sigma^*_{2s})^2(\pi_{2p})^4(\sigma_{2p})^1$
    (a) The HOMO is σ_{2p}
    (b) Bond order = $\frac{1}{2}(7 - 2) = 2\frac{1}{2}$
    (c) One-half net σ bond and two net π bonds
    (d) paramagnetic
12. 
    The electron-pair geometry is trigonal bipyramidal, and the molecular geometry is T-shaped. The Cl atom is *sp*^{3}*d* hybridized. Three of these hybrid orbitals each overlap a fluorine 2*p* orbital to form three Cl—F sigma bonds.

CH 222 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 Twenty** *and* **Chapter Guide Three**

Important Tables and/or Constants: "**Organic Chemistry Nomenclature Guide**" *(Handout)*, "**Organic Chemistry Lab**" *(Handout in Lab Packet)*

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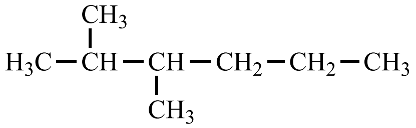
1. What is the name of the straight (unbranched) chain alkane with the formula C_7H_{16}? What is the molecular formula for an alkane with 9 carbon atoms?
2. Which of the following is an alkane? Which could be a cycloalkane?
   a. C_2H_4
   b. C_5H_{10}
   c. $C_{14}H_{30}$
   d. C_7H_8
3. Draw the structure of each of the following compounds: a. 2,3-dimethylhexane
   b. 3-ethylheptane
4. Draw structures for the *cis* and *trans* isomers of 4-methyl- 2-hexene.
5. Give the systematic name or structure for the following alcohols, amines and ethers.
   a. $CH_3CH_2CH_2OH$
   b. $CH_3CH_2CH_2CH_2OH$
   c. ethylamine
   d. dipropylamine
   e. dibutyl ether
   f. 1-methoxypropane
6. Draw structural formulas for a. 2-pentanone, b. hexanal, and c. pentanoic acid.
7. Draw structural formulas for the following compounds: a. 1,3-dichlorobenzene
   b. 1-bromo-4-methylbenzene
8. Draw structural formulas for the following acids and esters:
   a. 2-methylhexanoic acid
   b. pentyl butanoate (which has the odor of apricots)
   c. octyl acetate (which has the odor of oranges)
9. Aldehydes and carboxylic acids are formed by oxidation of primary alcohols, and ketones are formed when secondary alcohols are oxidized. Typical oxidizing agents include $K_2Cr_2O_7$ or $KMnO_4$. Give the name and formula for the alcohol that, when oxidized, gives the following products:
   a. $CH_3CH_2CH_2CHO$
   b. 2-hexanone
10. Ketones can be reduced with $LiAlH_4$ or $NaBH_4$ to create alcohols. Describe how to prepare 2-pentanol beginning with the appropriate ketone.
11. Draw the structure and give the systematic name for the products of the following reactions:
    a. $CH_3CH{=}CH_2 + Br_2 \rightarrow$
    b. $CH_3CH_2CH{=}CHCH_3 + H_2 \rightarrow$
12. The compound 2-bromobutane is a product of addition of HBr to an alkene. Identify the alkene and give its name.

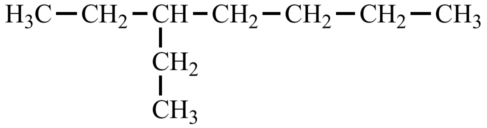
13. Draw structural formulas for all the alcohols with the formula $C_4H_{10}O$. Give the systematic name of each.
14. Draw structural formulas for all the primary amines with the formula $C_4H_9NH_2$. Name them.
15. Give structural formulas and systematic names for the three structural isomers of trimethylbenzene, $C_6H_3(CH_3)_3$.

Answers to the Practice Problem Set:

1. n-heptane; C_9H_{20}
2. c. $C_{14}H_{30}$ is an alkane b. C_5H_{10} could be a cycloalkane
3. *Answers:*

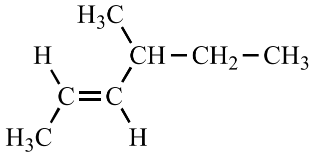
(a) 

$$\mathrm{H_3C-\overset{\substack{CH_3\\|}}{CH}-\underset{\substack{|\\CH_3}}{CH}-CH_2-CH_2-CH_3}$$

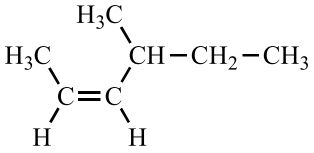
(b) 

$$\mathrm{H_3C-CH_2-\underset{\substack{|\\CH_2\\|\\CH_3}}{CH}-CH_2-CH_2-CH_2-CH_3}$$

4. *Answers:*



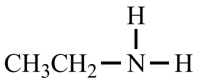
trans-4-methyl-2-hexene



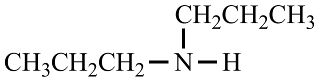
cis-4-methyl-2-hexene

5. *Answers:*

a. 1-propanol
b. 1-butanol

c. $C_2H_5NH_2$ 

$$\mathrm{CH_3CH_2-\overset{\substack{H\\|}}{N}-H}$$

d. $(C_3H_7)_2NH$ 

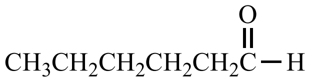
$$\mathrm{CH_3CH_2CH_2-\overset{\substack{CH_2CH_2CH_3\\|}}{N}-H}$$

e. $C_4H_9OC_4H_9$
f. $CH_3OCH_2CH_2CH_3$

6. *Answers:*

a. 

$$\mathrm{CH_3\overset{\substack{O\\\|}}{C}CH_2CH_2CH_3}$$

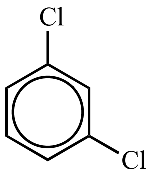
b. 

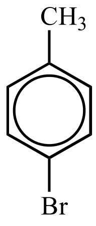
$$\mathrm{CH_3CH_2CH_2CH_2CH_2\overset{\substack{O\\\|}}{C}-H}$$

c. 

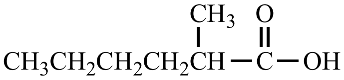
$$\mathrm{CH_3CH_2CH_2CH_2\overset{\substack{O\\\|}}{C}-OH}$$

7. *Answers:*

a. 

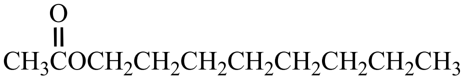
b. 

8. *Answers:*

a.

$CH_3CH_2CH_2CH_2\overset{CH_3}{\overset{|}{C}}H-\overset{O}{\overset{\|}{C}}-OH$

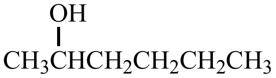
b.

$CH_3CH_2CH_2\overset{O}{\overset{\|}{C}}OCH_2CH_2CH_2CH_2CH_3$

c.

$CH_3\overset{O}{\overset{\|}{C}}OCH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$

9. *Answers:*

a. $CH_3CH_2CH_2CH_2-OH$ 1-butanol

b.

$CH_3\overset{OH}{\overset{|}{C}}HCH_2CH_2CH_2CH_3$ 2-hexanol

10. Reduction of 2-pentanone with $LiAlH_4$ or $NaBH_4$

11. *Answers:*

a.

$H_3C-\overset{Br}{\overset{|}{C}}H-CH_2-Br$ 1,2-dibromopropane

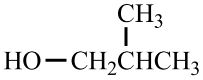
b. $CH_3CH_2CH_2CH_2CH_3$ pentane

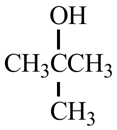
12. Using 1-butene, $H_2C{=}CHCH_2CH_3 + HBr \rightarrow CH_3CHBrCH_2CH_3$

13. *Answers:*

$HO-CH_2CH_2CH_2CH_3$ 1-butanol

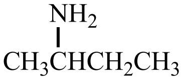

$CH_3\overset{OH}{\overset{|}{C}}HCH_2CH_3$ 2-butanol

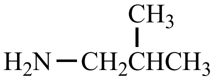

$HO-CH_2\overset{CH_3}{\overset{|}{C}}HCH_3$ 2-methyl-1-propanol

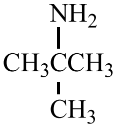

$CH_3\underset{CH_3}{\underset{|}{\overset{OH}{\overset{|}{C}}}}CH_3$ 2-methyl-2-propanol

14. *Answers:*

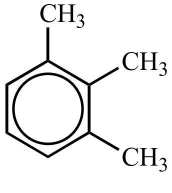
$H_2N-CH_2CH_2CH_2CH_3$ 1-butylamine


$CH_3\overset{NH_2}{\overset{|}{C}}HCH_2CH_3$ 2-butylamine

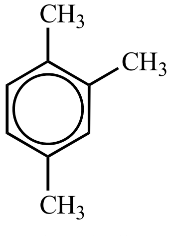

$H_2N-CH_2\overset{CH_3}{\overset{|}{C}}HCH_3$ 2-methyl-1-propylamine


$CH_3\underset{CH_3}{\underset{|}{\overset{NH_2}{\overset{|}{C}}}}CH_3$ 2-methyl-2-propylamine

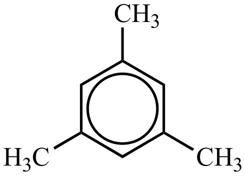
15. *Answers:*



1,2,3-trimethylbenzene



1,2,4-trimethylbenzene



1,3,5-trimethylbenzene

CH 222 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 Nine, Chapter Ten** *and* **Chapter Guide Four**

Important Tables and/or Constants: **R = 0.082057 L atm mol^{-1} K^{-1}, 760 mm Hg = 1 atm = 1013 mbar, 1 mbar = 1 hPa, 1 torr = 1 mm Hg**

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1. A sample of nitrogen gas has a pressure of 67.5 mm Hg in a 500. mL flask. What is the pressure of this gas sample when it is transferred to a 125 mL flask at the same temperature?
2. You have 3.5 L of NO at a temperature of 22.0 °C. What volume would the NO occupy at 37 °C? (Assume the pressure is constant.)
3. One of the cylinders of an automobile engine has a volume of 400. cm^3. The engine takes in air at a pressure of 1.00 atm and a temperature of 15 °C and compresses the air to a volume of 50.0 cm^3 at 77 °C. What is the final pressure of the gas in the cylinder?
4. A 1.25 g sample of CO_2 is contained in a 750. mL flask at 22.5 °C. What is the pressure of the gas?
5. A gaseous organofluorine compound has a density of 0.355 g/L at 17 °C and 189 mm Hg. What is the molar mass of the compound?
6. Acetaldehyde is a common liquid compound that vaporizes readily. Determine the molar mass of acetaldehyde from the following data:
   Sample mass = 0.107 g   Volume of gas = 125 mL
   Temperature = 0.0 °C   Pressure = 331 mm Hg
7. Iron reacts with hydrochloric acid to produce iron(II) chloride and hydrogen gas:
   $Fe(s) + 2\ HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
   The H_2 gas from the reaction of 2.2 g of iron with excess acid is collected in a 10.0-L flask at 25 °C. What is the pressure of the H_2 gas in this flask?
8. Sodium azide, the explosive compound in automobile air bags, decomposes according to the following equation:
   $2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$
   What mass of sodium azide is required to provide the nitrogen needed to inflate a 75.0 L bag to a pressure of 1.3 atm at 25 °C?
9. What is the total pressure in atmospheres of a gas mixture that contains 1.0 g of H_2 and 8.0 g of Ar in a 3.0 L container at 27 °C? What are the partial pressures of the two gases?
10. You have two flasks of equal volume. Flask A contains H_2 at 0 °C and 1 atm pressure. Flask B contains CO_2 gas at 25 °C and 2 atm pressure. Compare these two gases with respect to each of the following:
    a. average kinetic energy per molecule
    b. average molecular velocity
    c. number of molecules
    d. mass of gas
11. Place the following gases in order of increasing average molecular speed at 25 °C: Ar, CH_4, N_2, CH_2F_2.

12. There are five compounds in the family of sulfur–fluorine compounds with the general formula S_xF_y. One of these compounds is 25.23% S. If you place 0.0955 g of the compound in a 89 mL flask at 45 °C, the pressure of the gas is 83.8 mm Hg. What is the molecular formula of S_xF_y?
13. A miniature volcano can be made in the laboratory with ammonium dichromate. When ignited, it decomposes in a fiery display.

    $(NH_4)_2Cr_2O_7(s) \rightarrow N_2(g) + 4\ H_2O(g) + Cr_2O_3(s)$

    If 0.95 g of ammonium dichromate is used, and if the gases from this reaction are trapped in a 15.0 L flask at 23 °C, what is the total pressure of the gas in the flask? What are the partial pressures of N_2 and H_2O?
14. What type of intermolecular force must be overcome in converting each of the following from a liquid to a gas?

    a. liquid O_2 b. H_2O c. CH_3I d. CH_3CH_2OH
15. Rank the following atoms or molecules in order of increasing strength of intermolecular forces in the pure substance. Which exist as gases at 25 °C and 1 atm?

    a. Ne b. CH_4 c. CO d. CCl_4
16. In each pair of ionic compounds, which is more likely to have the greater heat of hydration? Briefly explain your reasoning in each case.
    - a. LiCl or CsCl
    - b. $NaNO_3$ or $Mg(NO_3)_2$
    - c. RbCl or $NiCl_2$
17. Ethanol, CH_3CH_2OH, has a vapor pressure of 59 mm Hg at 25 °C. What quantity of heat energy is required to evaporate 125 mL of the alcohol at 25 °C? The enthalpy of vaporization of the alcohol at 25 °C is 42.32 kJ/mol. The density of the liquid is 0.7849 g/mL.

Answers to the Practice Problem Set:

1. 270. mm Hg
2. 3.7 L
3. 9.72 atm
4. 0.919 atm
5. 34.0 g/mol
6. 44.1 g/mol
7. 0.096 atm
8. 170 g
9. 5.7 atm; 4.1 atm (H_2), 1.6 atm (Ar)
10. a. B > A   b. A > B   c. B > A   d. B > A
11. $CH_2F_2 < Ar < N_2 < CH_4$
12. S_2F_{10}
13. 0.031 atm; 0.0061 atm (N_2), 0.024 atm (H_2O)
14. a. induced dipole - induced dipole   b. hydrogen bonding   c. dipole-dipole   d. hydrogen bonding
15. $CH_4 < Ne < CO < CCl_4$ First three are gases
16. a. LiCl   b. $Mg(NO_3)_2$ c. $NiCl_2$
17. 90.1 kJ

CH 222 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 Ten, Chapter Eleven** *and* **Chapter Guide Five**

Important Tables and/or Constants: **R = 8.3145 J mol^{-1} K^{-1}**, "**Cubic Unit Cells Guide**" *(Handout)*, "**Solids**" *(Lab)*

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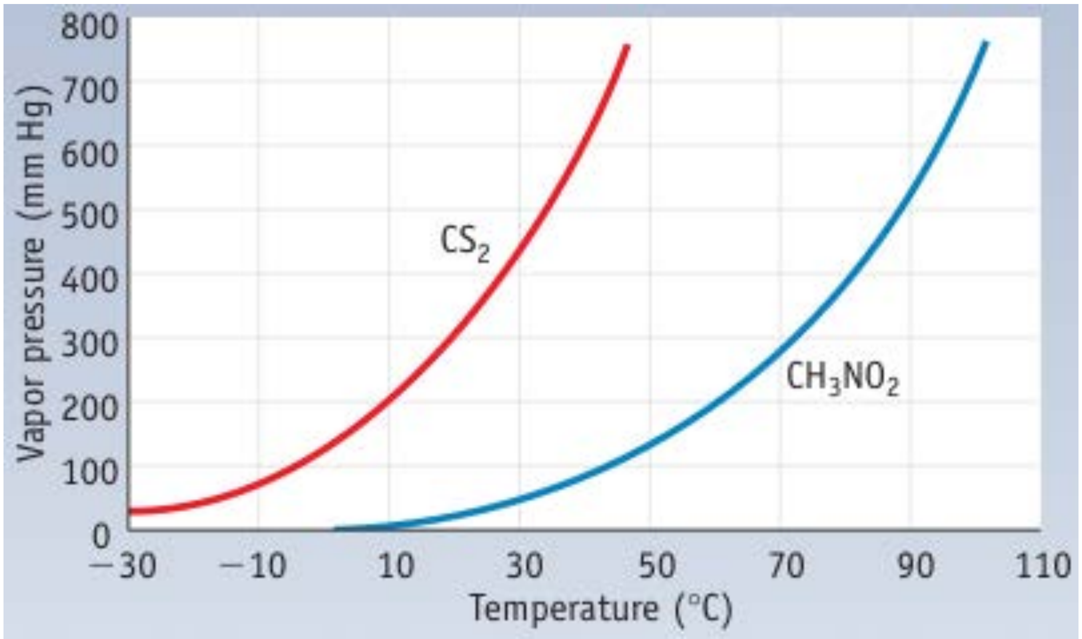
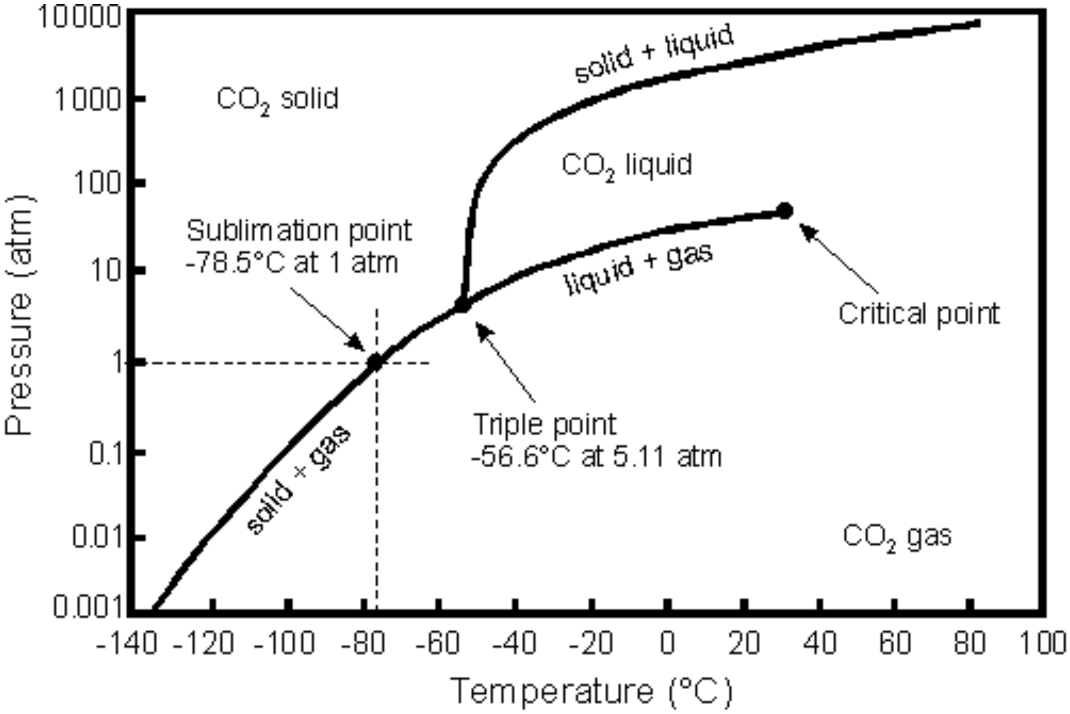
1. Vapor pressure curves for CS_2 (carbon disulfide) and CH_3NO_2 (nitromethane) are drawn here.
   a. What are the approximate vapor pressures of CS_2 and CH_3NO_2 at 40 °C?
   b. What types of intermolecular forces exist in the liquid phase of each compound?
   c. What is the normal boiling point of CS_2? Of CH_3NO_2?
   d. At what temperature does CS_2 have a vapor pressure of 600 mm Hg?
   e. At what temperature does CH_3NO_2 have a vapor pressure of 60 mm Hg?



2. Benzene, C_6H_6, is an organic liquid that freezes at 5.5 °C to form beautiful, feather-like crystals. How much heat is evolved when 15.5 g of benzene freezes at 5.5 °C? (The heat of fusion of benzene is 9.95 kJ/mol.) If the 15.5 g sample is remelted, again at 5.5 °C, what quantity of heat is required to convert it to a liquid?
3. Liquid ammonia, $NH_3(l)$, was once used in home refrigerators as the heat transfer fluid. The specific heat of the liquid is 4.7 J/g • K and that of the vapor is 2.2 J/g • K. The enthalpy of vaporization is 23.33 kJ/mol at the boiling point. If you heat 12 kg of liquid ammonia from -50.0 °C to its boiling point of -33.3 °C, allow it to evaporate, and then continue warming to 0.0 °C, how much heat energy must you supply?



Pressure-Temperature phase diagram for CO_2.

4. Use the phase diagram for carbon dioxide given above to answer the following questions:
   a. In what phase is CO_2 found at room temperature and 1.0 atm pressure?
   b. If the pressure exerted on a sample is 0.75 atm and the temperature is -114 °C, in what phase does the substance exist?
   c. If you measure the vapor pressure of a liquid sample and find it to be about 10 atm, what is the temperature of the liquid phase?
   d. What is the vapor pressure of the solid at -120 °C?
   e. Which is the denser phase, solid or liquid? Explain.
5. The very dense metal iridium has a face-centered cubic unit cell and a density of 22.56 g/cm^3. Use this information to calculate the radius of an atom of the element.
6. Use the vapor pressure data (below) for octane, C_8H_{18}, and the Clausius-Clapeyron equation to calculate the molar enthalpy of vaporization of octane and its normal boiling point.

Temperature (°C)	**Vapor Pressure (mm Hg)**
25	13.6
50	45.3
75	127.2
100	310.8

7. Liquid methanol, CH_3OH, is placed in a glass tube. Predict whether the meniscus of the liquid is concave or convex.

8. Rationalize the observation that $CH_3CH_2CH_2OH$, 1-propanol, has a boiling point of 97.2 °C, whereas a compound with the same empirical formula, methyl ethyl ether ($CH_3CH_2OCH_3$) boils at 7.4 °C.
9. Fill in the blanks in the table. All solutions are aqueous.

Compound	**Molality**	**Weight Percent**	**Mole Fraction**
NaI	0.15	_____	_____
C_2H_5OH	_____	5	_____
$C_{12}H_{22}O_{11}$	_____	_____	0.0027

10. Hydrochloric acid is sold as a concentrated aqueous solution. If the molarity of commercial HCl is 12.0 and its density is 1.18 g/cm^3, calculate the following:
    a. the molality of the solution
    b. the weight percent of HCl in the solution
11. The average lithium ion concentration in sea water is 0.18 ppm. What is the molality of Li^+ in sea water?
12. An unopened soda can has an aqueous CO_2 concentration of 0.0506 M at 25 °C. What is the pressure of CO_2 gas in the can? ($k_H = 4.48 \times 10^{-5}$ M/mm Hg)
13. Pure iodine (105 g) is dissolved in 325 g of CCl_4 at 65 °C. Given that the vapor pressure of CCl_4 at this temperature is 531 mm Hg, what is the vapor pressure of the CCl_4–I_2 solution at 65 °C? (Assume that I_2 does not contribute to the vapor pressure.)
14. What is the boiling point of a solution composed of 15.0 g of $CHCl_3$ (which boils at 61.70 °C) and 0.515 g of the nonvolatile solute acenaphthene, $C_{12}H_{10}$, a component of coal tar? (K_{bp} = 3.63 °C/m)
15. Assume a bottle of wine consists of an 11 weight percent solution of ethanol (C_2H_5OH) in water. If the bottle of wine is chilled to -20 °C, will the solution begin to freeze? (K_{fp} = 1.86 °C/m)
16. Anthracene, a hydrocarbon obtained from coal, has an empirical formula of C_7H_5. To find its molecular formula you dissolve 0.500 g in 30.0 g of benzene (K_{bp} = 2.53 °C/m). The boiling point of the pure benzene is 80.10 °C, whereas the solution has a boiling point of 80.34 °C. What is the molecular formula of anthracene?
17. Phenylcarbinol is used in nasal sprays as a preservative. A solution of 0.52 g of the compound in 25.0 g of water (K_{fp} = -1.86 °C/m) has a melting point of -0.36 °C. What is the molar mass of phenylcarbinol?
18. An aqueous solution containing 1.00 g of bovine insulin (a protein, not ionized) per liter has an osmotic pressure of 3.1 mm Hg at 25 °C. Calculate the molar mass of bovine insulin.

Answers to the Practice Problem Set:

1. *Answers*:
   a. CS_2: 620 mm Hg CH_3NO_2: 80 mm Hg
   b. induced dipole/induced dipole; dipole-dipole
   c. 46 ºC; 100 ºC
   d. 39 ºC
   e. 34 ºC
2. -1.97 kJ evolved. +1.97 kJ absorbed for solid -> liquid.
3. $q_{\text{total}} = 9.4 \times 10^2$ kJ + 1.6×10^4 kJ + 8.8×10^2 kJ = $\mathbf{1.8 \times 10^4}$ **kJ**
4. a. gas b. solid c. between -40 and -10 d. 0.01 atm e. solid denser than liquid
5. 135.7 pm
6. ΔH_{vap} = 38.6 kJ/mol, and T = 128 ºC
7. The meniscus is concave since there are adhesive forces between the methanol and the silicate of the glass.
8. 1-propanol has stronger intermolecular forces (hydrogen bonding) than methyl ethyl ether (dipole-dipole) zinc
9. *Answers*:

Compound	Molality	Weight percent	Mole fraction
NaI	0.15	2.2	0.0027
C_2H_5OH	1.1	5.0	0.020
$C_{12}H_{22}O_{11}$	0.15	4.9	0.0027

10. a. 16.2 *m* b. 37.1%
11. 2.6 x 10^{-5} *m*
12. 1130 mm Hg
13. 444 mm Hg
14. 62.51 ºC
15. Solution will freeze beginning at -5.0 ºC
16. $C_{14}H_{10}$
17. 110 g/mol
18. 6.0 x 10^3 g/mol

CH 222 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 Twelve, Chapter Twenty-one** *and* **Chapter Guide Six**
Important Tables and/or Constants: **R = 8.3145 J mol^{-1} K^{-1}**, "**Reaction Mechanisms Guide**" *(Handout)*

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1. Give the relative rates of disappearance of reactants and formation of products for each of the following reactions.
   a. $2\ O_3(g) \rightarrow 3\ O_2(g)$
   b. $2\ HOF(g) \rightarrow 2\ HF(g) + O_2(g)$
2. In the reaction $2\ O_3(g) \rightarrow 3\ O_2(g)$, the rate of formation of O_2 is 1.5×10^{-3} mol/L • s. What is the rate of decomposition of O_3?
3. The reaction between ozone and nitrogen dioxide at 231 K is first order in both NO_2 and O_3:
   $2\ NO_2(g) + O_3(g) \rightarrow N_2O_5(s) + O_2(g)$
   a. Write the rate equation for the reaction.
   b. If the concentration of NO_2 is tripled, what is the change in the reaction rate?
   c. What is the effect on reaction rate if the concentration of O_3 is halved?
4. The data in the table are for the reaction of NO and O_2 at 660 K.
   $2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$

Reactant Concentration (M)		
[NO]	**[O_2]**	**Rate of Disappearance of NO (M s^{-1})**
0.01	0.01	2.5×10^{-5}
0.02	0.01	1.0×10^{-4}
0.01	0.02	5.0×10^{-5}

   a. Determine the order of the reaction for each reactant.
   b. Write the rate equation for the reaction.
   c. Calculate the rate constant.
   d. Calculate the rate (in mol/L • s) at the instant when [NO] = 0.015 M and [O_2] = 0.0050 M
   e. At the instant when NO is reacting at the rate 1.0×10^{-4} mol/L • s, what is the rate at which O_2 is reacting and NO_2 is forming?
5. The rate equation for the hydrolysis of sucrose to fructose and glucose:
   $$C_{12}H_{22}O_{11}(aq) + H_2O(l) \rightarrow 2\ C_6H_{12}O_6(aq)$$
   is "$-\Delta[\text{sucrose}]/\Delta t = k[C_{12}H_{22}O_{11}]$." After 2.57 h at 27 °C, the sucrose concentration decreased from 0.0146 M to 0.0132 M. Find the rate constant, *k*.
6. Ammonium cyanate, NH_4NCO, rearranges in water to give urea, $(NH_2)_2CO$:
   $$NH_4NCO(aq) \rightarrow (NH_2)_2CO(aq)$$
   The rate equation for this process is "Rate = $k[NH_4NCO]^2$" where k = 0.0113 L/mol • min. If the original concentration of NH_4NCO in solution is 0.229 mol/L, how long will it take for the concentration to decrease to 0.180 mol/L?

7. Hydrogen peroxide, $H_2O_2(aq)$, decomposes to $H_2O(l)$ and $O_2(g)$ in a reaction that is first order in H_2O_2 and has a rate constant $k = 1.06 \times 10^{-3}$ min^{-1}.
   a. How long will it take for 15% of a sample of H_2O_2 to decompose?
   b. How long will it take for 85% of the sample to decompose?
8. The compound $Xe(CF3)_2$ decomposes in a first-order reaction to elemental Xe with a half-life of 30. min. If you place 7.50 mg of $Xe(CF_3)_2$ in a flask, how long must you wait until only 0.25 mg of $Xe(CF_3)_2$ remains?
9. Gaseous NO_2 decomposes at 573 K:

$$2\ NO_2(g) \rightarrow 2\ NO(g) + O_2(g)$$

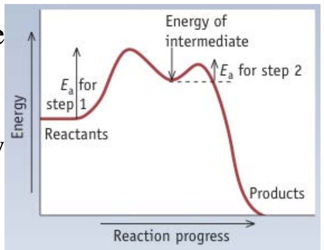
   The concentration of NO_2 was measured as a function of time. A graph of $1/[NO_2]$ versus time gives a straight line with a slope of 1.1 L/mol • s. What is the rate law for this reaction? What is the rate constant *k*?
10. Common sugar, sucrose, breaks down in dilute acid solution to form glucose and fructose. Both products have the same formula, $C_6H_{12}O_6$.

$$C_{12}H_{22}O_{11}(aq) + H_2O(l) \rightarrow 2\ C_6H_{12}O_6(aq)$$

   The rate of this reaction has been studied in acid solution, and the data in the table were obtained.

Time (min)	**$[C_{12}H_{22}O_{11}]$ (M)**
0	0.316
39	0.274
80	0.238
140	0.19
210	0.146

   a. Plot ln [sucrose] versus time and 1/[sucrose] versus time. What is the order of the reaction?
   b. Write the rate equation for the reaction, and calculate the rate constant, *k*.
   c. Estimate the concentration of sucrose after 175 min.
11. Answer the following questions based on the reaction coordinate diagram shown to the right.
    a. Is the reaction exothermic or endothermic?
    b. Does this reaction occur in more than one step? If so, how many?



12. What is the rate law for each of the following *elementary* reactions?
    a. $NO(g) + NO_3(g) \rightarrow 2\ NO_2(g)$
    b. $Cl(g) + H_2(g) \rightarrow HCl(g) + H(g)$
    c. $(CH_3)_3CBr(aq) \rightarrow (CH_3)_3C^+(aq) + Br^-(aq)$

13. Iodide ion is oxidized in acid solution by hydrogen peroxide:

$$H_2O_2(aq) + 2\,H^+(aq) + 2\,I^-(aq) \rightarrow I_2(aq) + 2\,H_2O(l)$$

A *proposed* mechanism is:

Step 1 *(slow)* $H_2O_2(aq) + I\text{-}(aq) \rightarrow H_2O(l) + OI^-(aq)$
Step 2 *(fast)* $H\text{+}(aq) + OI^-(aq) \rightarrow HOI(aq)$
Step 3 *(fast)* $HOI(aq) + H^+(aq) + I^-(aq) \rightarrow I_2(aq) + H_2O(l)$

a. Show that the three elementary steps add up to give the overall, stoichiometric equation.
b. What is the molecularity of each step?
c. For this mechanism to be consistent with kinetic data, what must be the experimental rate equation?
d. Identify any intermediates in the elementary steps in this reaction.

14. At temperatures below 500 K, the reaction between carbon monoxide and nitrogen dioxide

$$NO_2(g) + CO(g) \rightarrow CO_2(g) + NO(g)$$

has the following rate equation: **Rate = *k*[NO$_2$]^2** Which of the three mechanisms suggested below best agrees with the experimentally observed rate equation?

Mechanism 1 *Single, elementary step*
$NO_2 + CO \rightarrow CO_2 + NO$

Mechanism 2 *Two steps*
Slow $NO_2 + NO_2 \rightarrow NO_3 + NO$
Fast $NO_3 + CO \rightarrow NO_2 + CO_2$

Mechanism 3 *Two steps*
Slow $NO_2 \rightarrow NO + O$
Fast $CO + O \rightarrow CO_2$

T(K)	***k*(s^{-1})**
298	0.0409
308	0.0818
318	0.157

15. Data for the reaction **$[Mn(CO)_5(CH_3CN)]^+ + NC_5H_5 \rightarrow [Mn(CO)_5(NC_5H_5)] + CH_3CN$** are given in the table below. Calculate E_a from a plot of ln *k* versus 1/T.

16. Complete the following nuclear equations. Write the mass number and atomic number for the remaining particle, as well as its symbol.

a. ${}^{54}_{26}Fe + {}^{4}_{2}He \rightarrow 2\,{}^{1}_{0}n + ?$
b. ${}^{27}_{13}Al + {}^{4}_{2}He \rightarrow {}^{30}_{15}P + ?$
c. ${}^{32}_{16}S + {}^{1}_{0}n \rightarrow {}^{1}_{1}H + ?$
d. ${}^{96}_{42}Mo + {}^{2}_{1}H \rightarrow {}^{1}_{0}n + ?$
e. ${}^{98}_{42}Mo + {}^{1}_{0}n \rightarrow {}^{99}_{43}Tc + ?$
f. ${}^{18}_{9}F \rightarrow {}^{18}_{8}O + ?$

17. What particle is emitted in the following nuclear reactions? Write an equation for each reaction.
    a. Gold-198 decays to mercury-198.
    b. Radon-222 decays to polonium-218.
    c. Cesium-137 decays to barium-137.
    d. Indium-110 decays to cadmium-110.
18. Predict the probable mode of decay for each of the following radioactive isotopes, and write an equation to show the products of decay.
    a. Bromine-80m
    b. Californium-240
    c. Cobalt-61
    d. Carbon-11
19. Boron has two stable isotopes, ^{10}B and ^{11}B. Calculate the binding energies per nucleon of these two nuclei. The required masses (in grams per mole) are $^{1}_{1}H = 1.00783$, $^{1}_{0}n = 1.00867$, $^{10}B = 10.01294$, and $^{11}B = 11.00931$.
20. A three step mechanism for the reaction of $(CH_3)_3CBr$ and H_2O is proposed:

    Step 1 (slow) $(CH_3)_3CBr \rightarrow (CH_3)_3C^{+1} + Br^{-1}$

    Step 2 (fast) $(CH_3)_3C^{+1} + H_2O \rightarrow (CH_3)_3COH_2^{+1}$

    Step 3 (fast) $(CH_3)_3COH_2^{+1} + Br^{-1} \rightarrow (CH_3)_3COH + HBr$

    a. Write an equation for the overall reaction.
    b. Which step is rate determining?
    c. What rate law is expected for this reaction?
    d. What is the molecularity of each step?

Answers to the Practice Problem Set:

1. *Answers:*

   a. $$-\frac{1}{2}\left(\frac{\Delta[O_3]}{\Delta t}\right) = \frac{1}{3}\left(\frac{\Delta[O_2]}{\Delta t}\right)$$

   b. $$-\frac{1}{2}\left(\frac{\Delta[HOF]}{\Delta t}\right) = \frac{1}{2}\left(\frac{\Delta[HF]}{\Delta t}\right) = \frac{\Delta[O_2]}{\Delta t}$$

2. -1.0×10^{-3} mol/L • s
3. a. rate = $k[NO_2][O_3]$ b. rate triples c. rate halved
4. a. NO = 2nd order, O_2 = first order b. Rate = $k[NO]^2[O_2]$ c. $k = 13$ L^2/mol^2 • s d. 1.4×10^{-5} mol/L·s e. 5.0×10^{-5} mol/L·s (O_2), 1.0×10^{-4} mol/L·s (NO_2)
5. 0.0392 h^{-1}
6. 105 min (*note*: 2nd order kinetics!)
7. a. 153 min b. 1790 min
8. 150 min
9. rate = $k[NO_2]^2$ and $k = 1.1$ L/mol • s
10. a. first order b. $k = 0.0037$ min^{-1} c. 0.167 M
11. a. exothermic b. two steps
12. a) Rate = $k[NO][NO_3]$ b) Rate = $k[Cl][H_2]$ c) Rate = $k[(CH_3)_3CBr]$
13. a. (add equations together) b. bimolecular (steps 1 and 2) and termolecular (step 3) c. rate = $k[H_2O_2][I^{-1}]$ d. OI^- and HOI
14. Mechanism 2
15. y = -6373.3x + 18.19, r = -1 E_a = 53.0 kJ/mol
16. *Answers:*

    a. $^{56}_{28}Ni$ (d) $^{97}_{43}Tc$

    b. $^{1}_{0}n$ (e) $^{0}_{-1}\beta$

    c. $^{32}_{15}P$ (f) $^{0}_{+1}\beta$

17. *Answers:*

    a. $^{198}_{79}Au \rightarrow ^{198}_{80}Hg + ^{0}_{-1}\beta$

    b. $^{222}_{86}Rn \rightarrow ^{218}_{84}Po + ^{4}_{2}\alpha$

    c. $^{137}_{55}Cs \rightarrow ^{137}_{56}Ba + ^{0}_{-1}\beta$

    d. $^{110}_{49}In \rightarrow ^{110}_{48}Cd + ^{0}_{+1}\beta$

18. a. gamma decay b. alpha decay c. beta decay d. positron decay
19. For ^{10}B: 6.26×10^8 kJ/mol nucleons; for ^{11}B: 6.70×10^8 kJ/mol nucleons
20. a. $(CH_3)_3CBr + H_2O \rightarrow (CH_3)_3COH + HBr$ b. Step 1 c. Rate = $k[(CH_3)_3CBr]$ d. uni, bi, bi

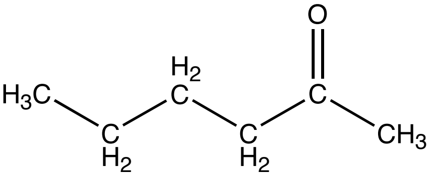
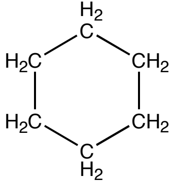
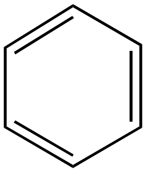
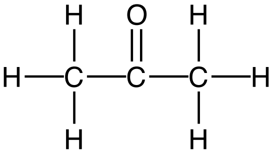
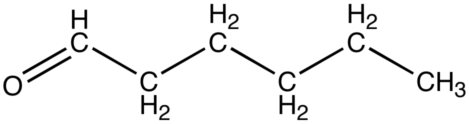
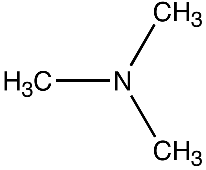
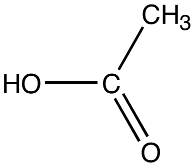
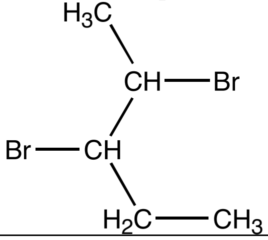
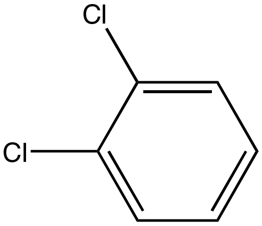
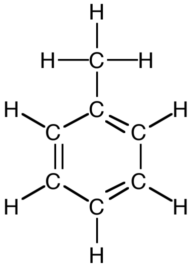
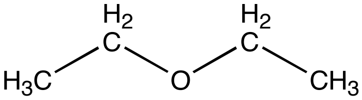
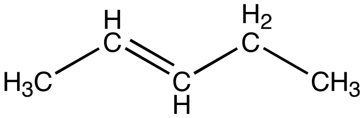
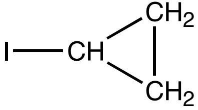
CH 222 Organic Chemistry
Self Quiz Name: ____ Lab Section: _______

This is a sample quiz for CH 222 providing examples of naming and nomenclature. Answers are provided on page 2. *Good luck!*

Provide names or structures for the following organic molecules:

n-pentane	2-hexanone	cyclohexane
formaldehyde	acetylene	dimethyl ether
2-butyne	benzene	acetone
hexanal	CH_4	CH_3OH
trimethylamine	ammonia	acetic acid
$CH_3CH_2CH_2CH_2OH$	2,3-dibromopentane	1,2-dichlorobenzene
$CH_3CH_2CH(OH)CH_3$	methylbenzene	ethoxy ethane
trans-2-pentene	iodoethane	cyclopropyl iodide

CH 222 Organic Chemistry *Self Quiz* - **ANSWERS**

n-pentane $\mathbf{CH_3CH_2CH_2CH_2CH_3}$	2-hexanone 	cyclohexane 
formaldehyde $\mathbf{H_2C{=}O}$	acetylene $\mathbf{HC{\equiv}CH}$	dimethyl ether $\mathbf{CH_3\text{-}O\text{-}CH_3}$
2-butyne $\mathbf{H_3C\text{-}C{\equiv}C\text{-}CH_3}$	benzene 	acetone 
hexanal 	CH_4 **methane**	CH_3OH **methanol**
trimethylamine 	ammonia $\mathbf{NH_3}$	acetic acid 
$CH_3CH_2CH_2CH_2OH$ **1-butanol**	2,3-dibromopentane 	1,2-dichlorobenzene 
$CH_3CH_2CH(OH)CH_3$ **2-butanol**	methylbenzene 	ethoxy ethane 
trans-2-pentene 	iodoethane $\mathbf{CH_3CH_2I}$	cyclopropyl iodide 

CH 222
Sample Organic Quiz Name: ______________________________ Lab Section: _______

Answers found at end.

Fill in the name or formula for the following open-chain hydrocarbons (no cycloalkanes):

1. C_3H_8 ______________

2. C_5H_8 ______________

3. C_2H_4 ______________

4. C_7H_{16} ______________

5. C_4H_{10} ______________

6. C_8H_{16} ______________

7. C_9H_{16} ______________

8. $C_{10}H_{22}$ ______________

9. ______________ butene

10. ______________ octyne

11. ______________ propene

12. ______________ pentane

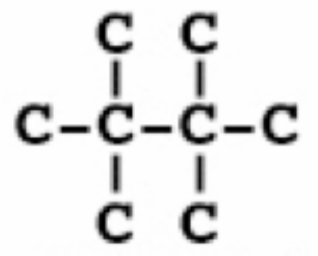
13. ______________ nonene

14. ______________ ethyne

15. ______________ hexyne

16. - 21. Draw and name the three isomers of pentane:

22. Name the following molecule:



23. Draw the molecule:

3-ethyl-2,2,5,5-tetramethyloctane

CH 222
Sample Organic Quiz Name: ______________________________ Lab Section: _______

*Answer key; sample quiz appears on the Chemistry 222 website. Answers appear in **bold**.*

Fill in the name or formula for the following open-chain hydrocarbons (no cycloalkanes):

1. C_3H_8 **propane**

2. C_5H_8 **pentyne**

3. C_2H_4 **ethene**

4. C_7H_{16} **heptane**

5. C_4H_{10} **butane**

6. C_8H_{16} **octene**

7. C_9H_{16} **nonyne**

8. $C_{10}H_{22}$ **decane**

9. **C_4H_8** butene

10. **C_8H_{14}** octyne

11. **C_3H_6** propene

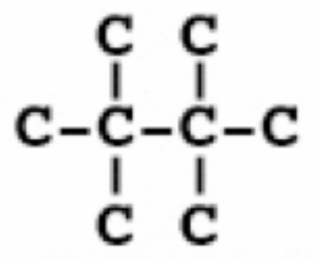
12. **C_5H_{12}** pentane

13. **C_9H_{18}** nonene

14. **C_2H_2** ethyne

15. **C_6H_{10}** hexyne

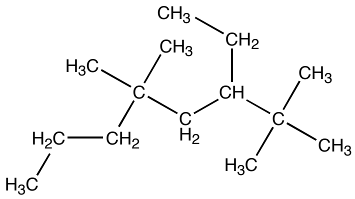
22. Name the following molecule:



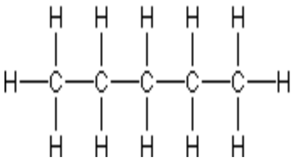
2,2,3,3-tetramethylbutane

23. Draw the molecule:

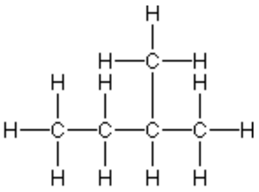
3-ethyl-2,2,5,5-tetramethyloctane



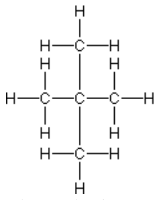
16. - 21. Draw and name the three isomers of pentane:



n-pentane

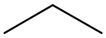


2-methylbutane



2,2-dimethylpropane

Organic Nomenclature Problems – *answers at end*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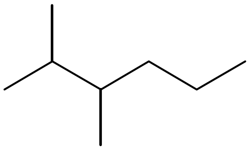
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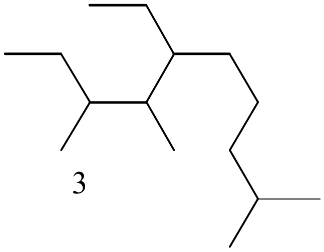
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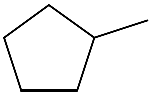
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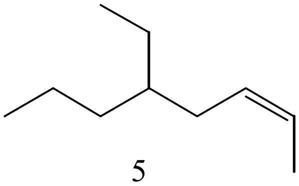
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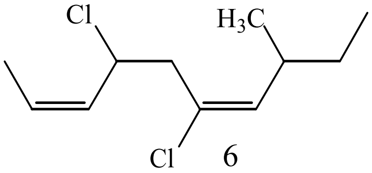

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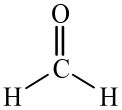

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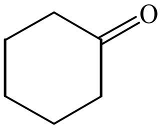

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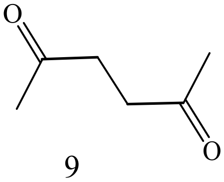

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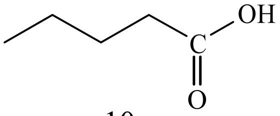

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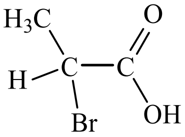

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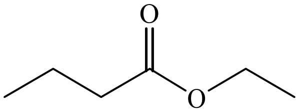

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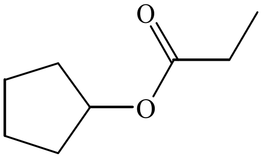

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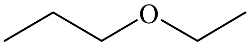

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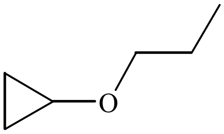

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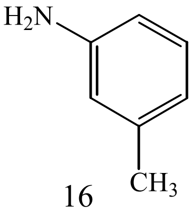

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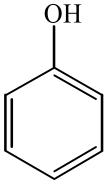

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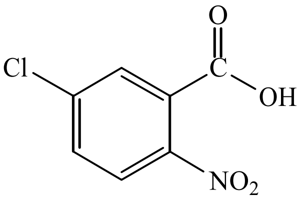


18

Answer Key

1. Propane
2. 2,3-dimethylhexane
3. 5-ethyl-3,4,9-trimethyldecane
4. methylcyclopentane
5. *cis*-5-ethyl-2-octene
6. *cis, cis*-4,6-dichloro-8-methyl-2,6-decadiene
7. methanal (formaldehyde)
8. cyclohexanone
9. 2,5-hexadione
10. pentanoic acid
11. 2-bromopropanoic acid
12. ethyl butanoate
13. cyclopentyl propanoate
14. ethyl propyl ether (1-ethoxypropane)
15. cyclopropyl propyl ether (1-cyclopropoxy propane)
16. 3-methylaniline
17. phenol
18. 5-chloro-2-nitrobenzoic acid

Organic Chemistry Practice Problems - *Answers at end*

Note: **Ar** = **benzene** (so is C_6H_5), **R** = any carbon chain and **R′** = carbons different from R.

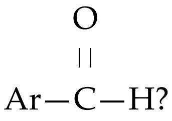
1. Which of the following is the general formula for an alcohol?
   (a) Ar—OH
   (b) $R—NH_2$
   (c) R—OH
   (d) R—O—R′
   (e) R—X

2. Which of the following is the general formula for a phenol (alcohol on a benzene ring)?
   (a) Ar—OH
   (b) $R—NH_2$
   (c) R—OH
   (d) R—O—R′
   (e) R—X

3. Which of the following is a general formula for ether?
   (a) Ar—OH
   (b) $R—NH_2$
   (c) R—OH
   (d) R—O—R′
   (e) R—X

4. Which of the following is a general formula for an amine?
   (a) Ar—OH
   (b) $R—NH_2$
   (c) R—OH
   (d) R—O—R′
   (e) R—X

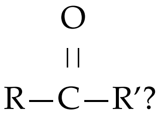
5. What class of compound has the following general formula:



$\mathrm{Ar{-}\overset{\overset{O}{\|}}{C}{-}H}$?

(a) aldehyde
(b) carboxylic acid
(c) ester
(d) ketone
(e) phenol

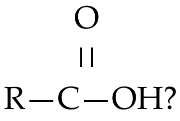
6. What class of compound has the following general formula:



$\mathrm{R{-}\overset{\overset{O}{\|}}{C}{-}R'}$?

(a) aldehyde
(b) alkene
(c) carboxylic acid
(d) ester
(e) ketone

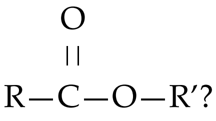
7. What class of compound has the following general formula:



$\mathrm{R{-}\overset{\overset{O}{\|}}{C}{-}OH}$?

(a) aldehyde
(b) carboxylic acid
(c) ester
(d) ketone
(e) phenol

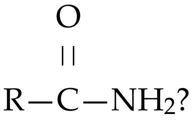
8. What class of compound has the following general formula:



$\mathrm{R{-}\overset{\overset{O}{\|}}{C}{-}O{-}R'}$?

(a) aldehyde
(b) carboxylic acid
(c) ester
(d) ketone
(e) phenol

9. What class of compound has the following general formula:



$\mathrm{R{-}\overset{\overset{O}{\|}}{C}{-}NH_2}$?

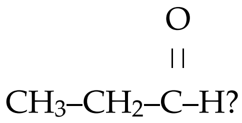
(a) aldehyde
(b) amide
(c) amine
(d) carboxylic acid
(e) ketone

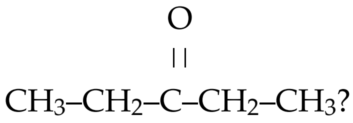
10. What class of compound is the following: $CH_3–CH_2–OH$?
    - (a) organic halide
    - (b) alcohol
    - (c) phenol
    - (d) ether
    - (e) amine

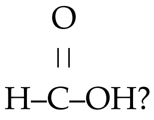
11. What class of compound is the following: $C_6H_5–OH$?
    - (a) organic halide
    - (b) alcohol
    - (c) phenol
    - (d) ether
    - (e) amine

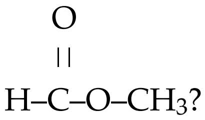
12. What class of compound is the following: $CH_3–CH_2–O–CH_2–CH_3$?
    - (a) organic halide
    - (b) alcohol
    - (c) phenol
    - (d) ether
    - (e) amine

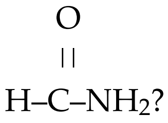
13. What class of compound is the following: $CH_3–CH_2–NH_2$?
    - (a) organic halide
    - (b) alcohol
    - (c) phenol
    - (d) ether
    - (e) amine

14. What class of compound is the following:  $CH_3–CH_2–\overset{\overset{\displaystyle O}{||}}{C}–H$?
    - (a) aldehyde
    - (b) ketone
    - (c) carboxylic acid
    - (d) ester
    - (e) amide

15. What class of compound is the following:  $CH_3–CH_2–\overset{\overset{O}{\|}}{C}–CH_2–CH_3$?
    (a) aldehyde
    (b) ketone
    (c) carboxylic acid
    (d) ester
    (e) amide

16. What class of compound is the following:  $H–\overset{\overset{O}{\|}}{C}–OH$?
    (a) aldehyde
    (b) ketone
    (c) carboxylic acid
    (d) ester
    (e) amide

17. What class of compound is the following:  $H–\overset{\overset{O}{\|}}{C}–O–CH_3$?
    (a) aldehyde
    (b) ketone
    (c) carboxylic acid
    (d) ester
    (e) amide

18. What class of compound is the following:  $H–\overset{\overset{O}{\|}}{C}–NH_2$?
    (a) aldehyde
    (b) ketone
    (c) carboxylic acid
    (d) ester
    (e) amide

19. What class of compound is methanol?
    (a) aldehyde
    (b) ketone
    (c) carboxylic acid
    (d) ester
    (e) alcohol

20. What class of compound is butanal?
    (a) aldehyde
    (b) ketone
    (c) carboxylic acid
    (d) ester
    (e) alcohol

21. What class of compound is methyl propanoate?
    (a) aldehyde
    (b) ketone
    (c) carboxylic acid
    (d) ester
    (e) alcohol

22. What class of compound is butanone?
    (a) aldehyde
    (b) ketone
    (c) carboxylic acid
    (d) ester
    (e) alcohol

23. What class of compound is ethanamide?
    (a) aldehyde
    (b) ketone
    (c) amide
    (d) ester
    (e) alcohol

24. What class of compound is dopamine?
    (a) aldehyde
    (b) ketone
    (c) carboxylic acid
    (d) amine
    (e) alcohol

Answer Key

1C, 2A, 3D, 4B, 5A, 6E, 7B, 8C, 9B, 10B, 11C, 12D, 13E, 14A, 15B, 16C, 17D, 18E, 19E, 20A, 21D, 22B, 23C, 24D

CH 222 Solution Calculations
Self Quiz Name: ____ Lab Section: _______

This is a sample quiz for CH 222 providing examples of solution calculations. Answers are provided on the next pages. *Good luck!*

Problem #1: Fill in the blanks in the table. All solutions are aqueous.

Solute	*solution density, g/cm³*	*molality (m)*	*weight percent solute*	*mole fraction solute (X)*	*Molarity (M)*
$NiCl_2$	0.99777			**0.00331**	
acetone	0.9163				**4.00 M**
NH_4NO_3	1.112	**3.54 m**			
HCl	1.47		**90.0%**		

Problem #2: Fill in the blanks in the table. All solutions are aqueous.

Solute	*solution density, g/cm³*	*molality (m)*	*weight percent solute*	*mole fraction solute (X)*	*Molarity (M)*
KBr	0.9977			**0.120**	
ethanol	0.9163				**0.801 M**
MgI_2	1.112	**1.17 m**			
H_2SO_4	1.84		**95.0%**		

Answers appear on the next pages

*Problem #1 Answers: (answers in **bold**)*

Solute	*solution density, g/cm^3*	*molality (m)*	*weight percent solute*	*mole fraction solute (X)*	*Molarity (M)*
$NiCl_2$	0.99777	**0.184 m**	**2.33%**	*0.00331*	**0.180 M**
acetone	0.9163	**5.85 m**	**25.3%**	**0.0943**	*4.00 M*
NH_4NO_3	1.112	*3.54 m*	**22.1%**	**0.0599**	**3.07 M**
HCl	1.47	**247 m**	*90.0%*	**0.817**	**36.3 M**

For $NiCl_2$: molar mass = 129.59 g/mol. Assume **1.00000 mol total**

Solute: **0.00331 mol $NiCl_2$** * 129.59 g/mol = **0.429 g $NiCl_2$**
Solvent: 1 – 0.00331 = **0.99669 mol H_2O** * 18.016 g/mol = **17.956 g H_2O**

molality = 0.00331 mol $NiCl_2$ / 0.017956 kg H_2O = **0.184 m**
wt% = (0.429 g $NiCl_2$) / (0.429 g $NiCl_2$ + 17.956 g H_2O) * 100% = **2.33%**

mass solution = (0.429 g $NiCl_2$ + 17.956 g H_2O) = **18.385 g solution**
volume solution = 18.385 g solution * (mL / 0.99777 g) = **18.426 mL solution**
molarity = 0.00331 mol $NiCl_2$ / 0.018426 L = **0.180 M**

For acetone (CH_3COCH_3): molar mass = 58.078 g/mol. Assume **1.00 L of solution**.

Solute: **4.00 mol acetone** * 58.078 g/mol = **232 g acetone**

mass solution = 1.00 L * (10^3 mL/L) * (0.9163 g/mL) = **916 g solution**
mass solvent = 916 g solution – 232 g solute = **684 g solvent (H_2O)**
moles solvent = 684 g H_2O * (mol / 18.016 g) = **38.0 mol H_2O**

molality = 4.00 mol acetone / 0.684 kg H_2O = **5.85 m**
wt% = (232 g acetone) / (232 g acetone + 684 g H_2O) * 100% = **25.3%**
X = (4.00 mol acetone) / (4.00 mol acetone + 38.0 mol H_2O) = **0.0943**

Answers continue on next page

For NH_4NO_3: molar mass = 80.052 g/mol. Assume **1.000 kg total solvent**

Solute: **3.54 mol NH_4NO_3** * 80.052 g/mol = **283 g NH_4NO_3**
Solvent: 1.000 kg * (10^3 g/kg) = **1000. g** * (1 mol / 18.016 g) = **55.51 mol H_2O**

X = (3.54 mol NH_4NO_3) / (3.54 mol NH_4NO_3 + 55.51 mol H_2O) = **0.0599**
wt% = (283 g NH_4NO_3) / (283 g NH_4NO_3+ 1000. g H_2O) * 100% = **22.1%**

mass solution = (283 g NH_4NO_3 + 1000. g H_2O) = **1283 g solution**
volume solution = 1283 g solution * (mL / 1.112 g) = **1154 mL solution**
molarity = 3.54 mol NH_4NO_3 / 1.154 L = **3.07 M**

For HCl: molar mass = 36.458 g/mol. Assume **100.0 g total solution**

Solute: **90.0 g HCl** * (mol / 36.458 g/mol) = **2.47 mol HCl**
Solvent: (100.0 – 90.0) = **10.0 g H_2O** * (1 mol / 18.016 g) = **.555 mol H_2O**

X = (2.47 mol **HCl**) / (2.47 mol **HCl** + 0.555 mol H_2O) = **0.817**
molality = 2.47 mol HCl / 0.0100 kg H_2O = **247 m**

mass solution = **100.0 g solution**
volume solution = 100.0 g solution * (mL / 1.47 g) = **68.0 mL solution**
molarity = 2.47 mol HCl / 0.0680 L = **36.3 M**

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*Problem #2 Answers: (answers in **bold**, try working the problems out on your own!)*

Solute	*solution density, g/cm^3*	*molality (m)*	*weight percent solute*	*mole fraction solute (X)*	*Molarity (M)*
KBr	0.9977	**7.55 m**	**47.4%**	*0.120*	**3.96 M**
ethanol	0.9163	**0.911 m**	**4.03%**	**0.0161**	*0.801 M*
MgI_2	1.112	*1.17 m*	**24.5%**	**0.0206**	**0.982 M**
H_2SO_4	1.84	**190 m**	*95.0%*	**0.777**	**17.8 M**

Colligative Properties and Phase Diagrams *- answers at end*

Colligative Properties

1) What is the boiling point of 0.10 M $CaCl_{2(}$(aq)? Would the actual boiling point be higher or lower than what you calculated? Why? $K_b = 0.512$

2) 10 grams of a non-ionic solid ($i = 1$) are dissolved into 100 mL of water. The freezing point of the water is depressed 3.32 °C. What is the molar mass of the solid? $K_f = 1.86$

3) Osmotic pressure can be used to determine the formula weight of a compound and is commonly used to calculate the mass of proteins. If 1.00 grams of a protein are dissolved in 100 mL of water, the osmotic pressure will be 92.92 torr. What is the molar mass of the protein? Assume $i = 1$

4) Ocean water has about the same amount of salt in it as 0.82 M NaCl. How much pressure must be exerted if you wanted to turn ocean water into salt water by reverse osmosis? Assume 25 °C.

5) The freezing point depression of a dilute solution of Ammonium Hydroxide (NH_4OH) was used measured and found to have an $i = 1.10$. Since this value should not be rounded and it was not a whole value, what can you conclude about the NH_4OH when it dissolves in water?

6) Why do road crews throw salt on the road when it snows? How does it do it?

Phase Diagrams

1a) Draw the phase diagram for naphthalene. It melts at 80.5 °C, boils at 218 °C, and the solid is more dense that its liquid. You do not have to be perfect. Just give the essential features of the diagram.

1b) Using the diagram just drawn, show how solid naphthalene is more dense than liquid naphthalene.

2) Using the phase diagram for water, explain why an ice skater can glide smoothly across the ice. What is the skater actually gliding on?

3) Sketch a phase diagram for benzene and locate these points: the triple point (5.5 °C and 35.8 torr), the boiling point (80.1 °C) and the critical point (288.5 °C and 47.7 atm). Solid benzene does not float on its liquid.

4) Sketch a phase diagram for dry ice (CO2) and locate these points: the triple point (-56.4 °C and 5.11 atm), the sublimation point (-78.5 °C) and the critical point (31.1 °C and 73 atm). Solid CO2 does not float on its liquid.

Colligative Properties

1) What is the boiling point of 0.10 M $CaCl_2$? Would the actual boiling point be higher or lower than what you calculated? Why? $K_b = 0.512$

> $\Delta T = 0.512\,(0.1)(3) = 0.1536$ °C + 100 °C = 100.1536 °C. The actual BP would be lower because *i* will not be exactly 3. This is because the ions are not completely separate in solution - they "clump" together a little, making the solution look less concentrated than it really is. We usually ignore this affect when we do our calculations.

2) 10 grams of a non-ionic solid (*i* = 1) are dissolved into 100 mL of water. The freezing point of the water is depressed 3.32 °C. What is the molar mass of the solid? $K_f = 1.86$

> $\Delta T = 1.86\,(m)(1) = 3.32$°C —> m = 1.785 M = (10g/MM)/0.100 L) so MM = 56.02 g/mol.

3) Osmotic pressure can be used to determine the formula weight of a compound and is commonly used to calculate the mass of proteins. If 1.00 gram of a protein is dissolved in 100 mL of water, the osmotic pressure will be 92.92 torr. What is the molar mass of the protein? Assume *i* = 1 and 25 ℃.

> (92.92 torr/760 torr/atm) = c(0.08206)(298) —> c = 0.005 M (1gram/MM)/(0.100L) =
>
> 0.005 M —> MM = 2000 g/mol

4) Ocean water has about the same amount of salt in it as 0.82 M NaCl. How much pressure must be exerted if you wanted to turn ocean water into salt water by reverse osmosis? Assume 25 °C.

> $\Pi = (2)(0.82\text{ M})(0.08206)(298\text{K}) = 40.1$ atm or more

5) The freezing point depression of a dilute solution of Ammonium Hydroxide (NH_4OH) was used measured and found to have an *i* = 1.10. Since this value should not be rounded and it was not a whole value, what can you conclude about the NH_4OH when it dissolves in water?

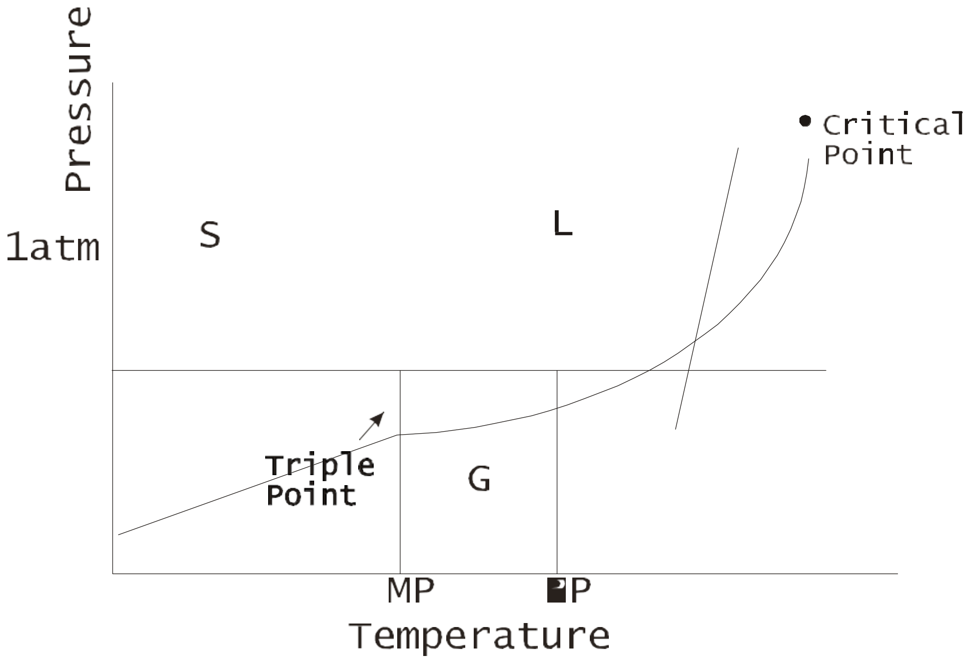
> It seems that NH_4OH does not form NH_4^+ and OH^- in water. It stays mostly as NH_4OH and only a small amount becomes NH_4^+ and OH^-.

6) Why do road crews throw salt on the road when it snows? How does it do it?

> The salt causes the ice to melt by lowering its freezing point.

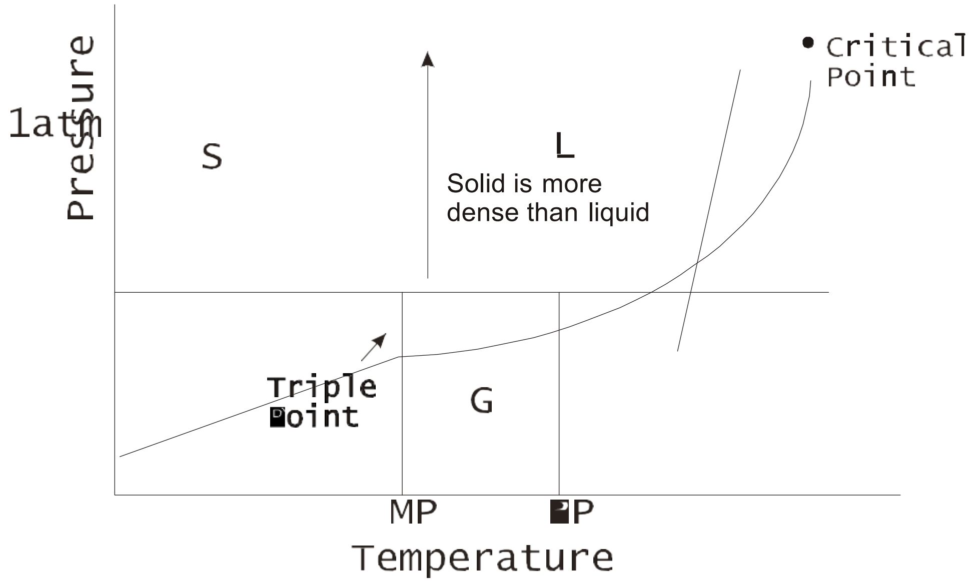
Phase Diagrams

1a) Draw the phase diagram for naphthalene. It melts at 80.5 °C, boils at 218 °C, and the solid is more dense that its liquid. You do not have to be perfect. Just give the essential features of the diagram.



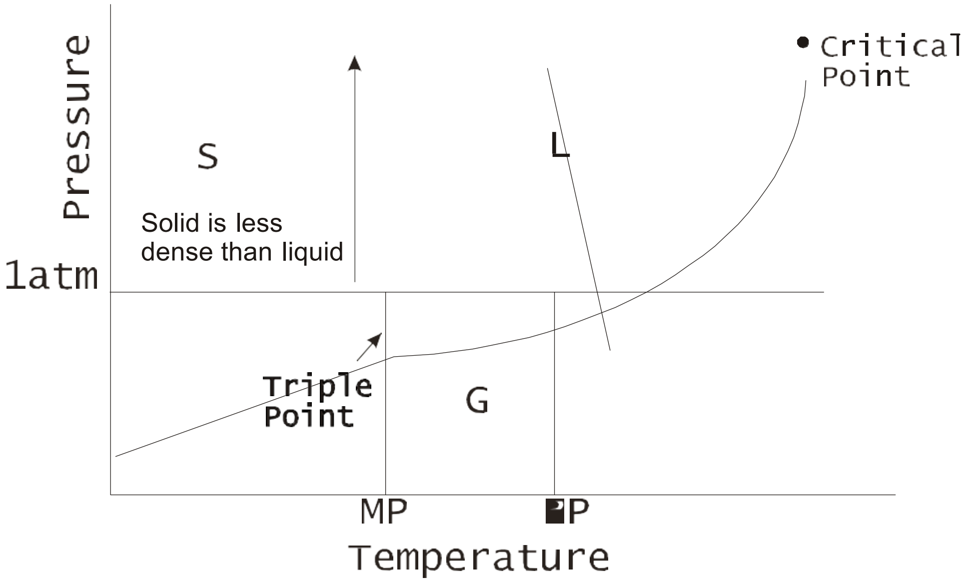
1b) Using the diagram just drawn, show how solid naphthalene is more dense than liquid naphthalene.

As pressure increases you move from the liquid into the solid phase. The more dense medium is the solid so it sinks on the liquid.



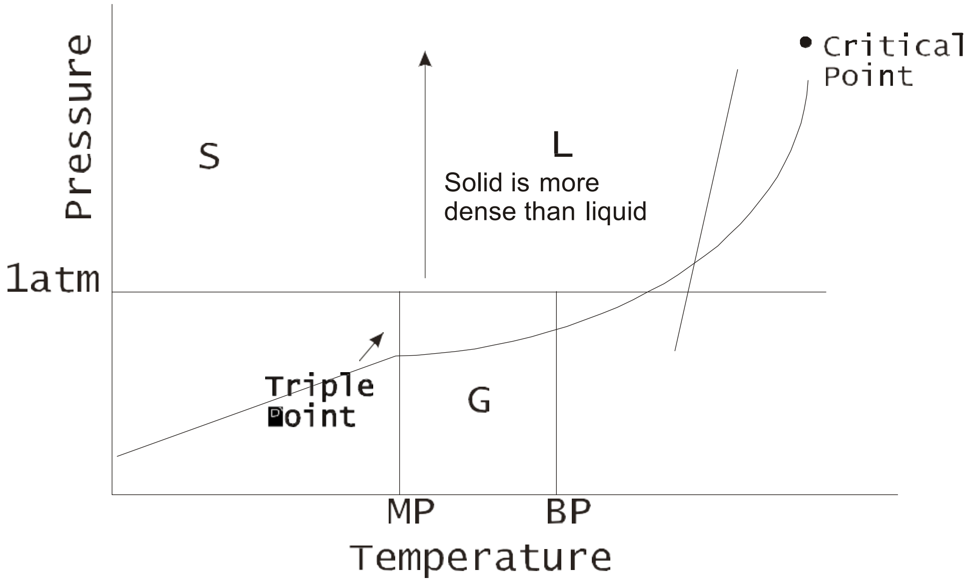
2) Using the phase diagram for water, explain why an ice skater can glide smoothly across the ice. What is the skater actually gliding on?

The pressure from the skate melts the ice so the skater actually skates on a thin layer of water.

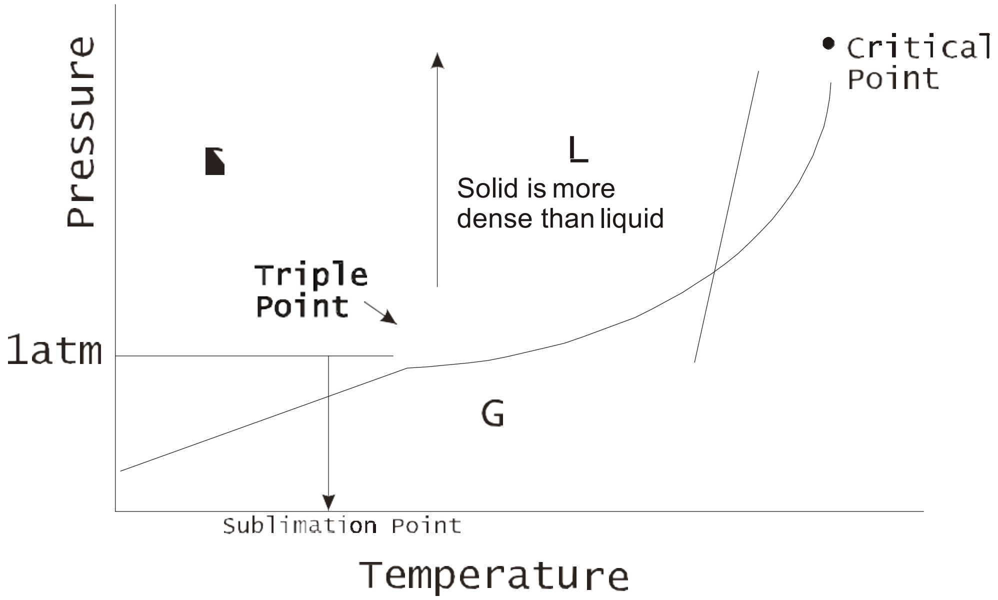


3) Sketch a phase diagram for benzene and locate these points: the triple point (5.5 °C and 35.8 torr), the boiling point (80.1 °C) and the critical point (288.5 °C and 47.7 atm).

The diagrams all look the same. The only thing that changes is the temperature and pressures where stuff happens.



4) Sketch a phase diagram for dry ice (CO_2) and locate these points: the triple point (-56.4 °C and 5.11 atm), the sublimation point (-78.5 °C) and the critical point (31.1 °C and 73 atm). Solid CO_2 does not float on its liquid



In this case, CO_2 does not have a boiling point or melting point. BP and MP are defined as occurring where the 1 atm pressure line crosses the solid/liquid line (melting point) and the liquid gas line (boiling point). The 1 atm line for CO_2 does not cross either of these line but crosses the solid/gas line which is sublimation.

Sample Chemistry Question (Ch. 9 - 11) - CH 222

Questions for Chapters Nine, Ten and Eleven:

1. When 7.00 grams of Helium and 14.0 grams of argon were mixed in a flask, the pressure was measured as 712 torr. What is the partial pressure of helium?

a. 593 torr
b. 356 torr
c. 833 torr
d. 1070 torr
e. 1420 torr

2. The empirical formula of a certain hydrocarbon is CH_2. When 0.125 moles of this hydrocarbon are completely combusted with excess oxygen, it is observed that 11.2 L of H_2O gas is produced at STP. What is the molecular formula of the unknown hydrocarbon?

a. C_2H_4
b. C_2H_3
c. C_3H_6
d. C_4H_8
e. C_6H_{12}

3. Lead (atomic mass 207.2 g/mol) crystallizes in a face-centered cubic arrangement. The radius of an atom of lead is $1.75*10^{-8}$ cm. What is the density of lead?

a. 9.85 g/cm^3
b. 11.4 g/cm^3
c. 13.2 g/cm^3
d. 14.7 g/cm^3
e. 19.7 g/cm^3

4. Rank the compounds NH_3, CH_4, and SiH_4 in order of increasing boiling point.

a. $NH_3 < CH_4 < SiH_4$
b. $CH_4 < NH_3 < SiH_4$
c. $NH_3 < SiH_4 < CH_4$
d. $CH_4 < SiH_4 < NH_3$
e. $SiH_4 < NH_3 < CH_4$

5. A 1.34 M $NiCl_2$ aqueous solution has a density of 1.12 g/cm^3. What is the molality of the solution?

a. 0.913 m
b. 1.42 m
c. 1.55 m
d. 2.55 m
e. 3.13 m

6. A solution is prepared by dissolving 0.500 g of non-dissociating solute in 12.0 g of cyclohexane. The freezing point depression of the solution is 8.94 °C. K_f for cyclohexane is 20.0 °C/m. Calculate the molar mass of the solute.

a. 93.3 g/mol
b. 112 g/mol
c. 128 g/mol
d. 182 g/mol
e. 205 g/mol

Here are the answers to the previous questions:

----

1. When 7.00 grams of Helium and 14.0 grams of argon were mixed in a flask, the pressure was measured as 712 torr. What is the partial pressure of helium?

a. 593 torr
b. 356 torr
c. 833 torr
d. 1070 torr
e. 1420 torr

Answer: Pressure is proportional to the quantity of moles present in a gas, and Dalton's Law of Partial Pressures says that the total pressure will be equal to the sum of the respective pressures. The mole fraction of the gas times the total pressure (712 torr) will give the respective pressures of the components. There are (7.00 g / 4.00 g/mol) = 1.75 mol He and (14.0 g / 39.9 g/mol) = 0.351 mol Ar. The mole fraction of He is 1.75 / (1.75 + 0.351) = 0.833, and the partial pressure of helium will be 0.833 * 712 torr = **593 torr**, answer **(a)**.

----

2. The empirical formula of a certain hydrocarbon is CH_2. When 0.125 moles of this hydrocarbon are completely combusted with excess oxygen, it is observed that 11.2 L of H_2O gas is produced at STP. What is the molecular formula of the unknown hydrocarbon?

a. C_2H_4
b. C_2H_3
c. C_3H_6
d. C_4H_8
e. C_6H_{12}

Answer: First, find the moles of H_2O produced, and then compare this number to 0.125 mol to find the ratio of "H_2" units in the hydrocarbon. Ex: ethane, C_2H_4, has two "H_2" units, so each mole of ethane will give two moles of water.

Also recall that STP implies a temperature of 273 K and 1 atm of pressure.

To find the moles of water: n = PV/RT = 1 atm*11.2 L / 0.082057 * 273 K = 0.500 mol H_2O.

The ratio (0.500 / 0.125) = 4, and this is how many "H_2" units are present in the hydrocarbon. Since the empirical formula is CH_2, we can imply there are four carbons with the four "H_2" units.

The final answer is **(d), C_4H_8.**

----

3. Lead (atomic mass 207.2 g/mol) crystallizes in a face-centered cubic arrangement. The radius of an atom of lead is $1.75*10^{-8}$ cm. What is the density of lead?

a. 9.85 g/cm^3
b. 11.4 g/cm^3
c. 13.2 g/cm^3
d. 14.7 g/cm^3
e. 19.7 g/cm^3

Answer: Using unit analysis, we need an answer in g/cm^3. The numerator, g, comes from the molar mass of lead (207.2 g/mol), for the mol can be cancelled using Avogadro's number and by remembering that the face centered cubic requires four atoms per unit cell. The denominator, cm^3, can be obtained from the equation: Edge = $4*radius/(2)^{1/2}$ and remembering that volume (in cm^3) is equal to the cubed root of the edge.

So: numerator (g) = 207.2 g/mol * (mol / $6.022*10^{23}$ atoms) * (4 atoms / cell) = $1.38 *10^{-21}$ g
denominator (cm^3) = $(edge)^3 = (4*1.75*10^{-8}\ cm/2^{1/2})^3 = 1.21 * 10^{-22}\ cm^3$
Therefore, density = $1.38 *10^{-21}$ g / $1.21 * 10^{-22}\ cm^3$ = **11.4 g/cm^3**, answer **(b)**.

----

4. Rank the compounds NH_3, CH_4, and SiH_4 in order of increasing boiling point.

a. $NH_3 < CH_4 < SiH_4$
b. $CH_4 < NH_3 < SiH_4$
c. $NH_3 < SiH_4 < CH_4$
d. $CH_4 < SiH_4 < NH_3$
e. $SiH_4 < NH_3 < CH_4$

Answer: Two factors affect boiling point: 1) mass (molar mass), and 2) strength of intermolecular forces. The larger the molar mass or intermolecular force, the higher the boiling point.

In terms of molar mass, the order would be CH_4 (16 g/mol) < NH_3 (17 g/mol) < SiH_4 (32 g/mol). However, while CH_4 and SiH_4 have only induced dipole - induced dipole (ID-ID) forces, NH_3 experiences hydrogen bonding forces, which is much stronger than the ID-ID force. Because the mass difference is not too large between NH_3 and SiH_4, we would expect the order to be **$CH_4 < SiH_4 < NH_3$**, answer **(d)**.

----

5. A 1.34 M $NiCl_2$ aqueous solution has a density of 1.12 g/cm^3. What is the molality of the solution?

a. 0.913 m
b. 1.42 m
c. 1.55 m
d. 2.55 m
e. 3.13 m

Answer: Converting from molarity to molality (or weight percent or mole fraction) requires a solution density value. 1.34 M implies that there are 1.34 moles of $NiCl_2$ in 1 L of solution.

We wish to calculate molality, which is the moles of solute per kg of solvent. We have 1.34 mol of solute in 1 L of solution, so if we can calculate the kg of solvent, we'll be good to go.

1.34 mol * 129.6 g/mol = 174 g $NiCl_2$
1 L * (1000 mL / L) * (cm^3 / mL) * (1.12 g/cm^3) = 1120 g of solution (which equals the g of $NiCl_2$ and water)
g H_2O = 1120 g - 174 g = 950 g water * (1 kg / 1000 g) = 0.95 kg water

molality = mol solute / kg solvent = 1.34 mol / 0.95 kg = **1.4 m**, answer **(b)**.

----
6. A solution is prepared by dissolving 0.500 g of non-dissociating solute in 12.0 g of cyclohexane. The freezing point depression of the solution is 8.94 °C. K_f for cyclohexane is 20.0 °C/m. Calculate the molar mass of the solute.

a. 93.3 g/mol
b. 112 g/mol
c. 128 g/mol
d. 182 g/mol
e. 205 g/mol

Answer: molar mass = g/mol. We have 0.500 g for the mass, so we need to calculate the moles of solute present. We can use the freezing point depression equation, $\Delta T = k_f * m * i$.

Since the solute is non-dissociating, we can assume that i = 1.

Solve for m: $m = \Delta T / k_f$ = 8.94 °C / 20.0 °C/m = 0.447 m = 0.447 mol solute / kg solvent (cyclohexane)

Since the problem used 12.0 g of solvent,
{0.447 mol solute / kg solvent (cyclohexane)}* 0.0120 kg solvent = $5.36*10^{-3}$ mol solute

Molar mass = g solute / mol solute = 0.500 g / $5.36*10^{-3}$ mol = **93.3 g/mol**, answer **(a)**.

Nuclear Chemistry - Radioactive Decay *(answers at end)*

1. What particle is emitted when a Fr-210 nucleus decays to At-206?
    - (a) alpha
    - (b) beta
    - (c) neutron
    - (d) positron
    - (e) proton

2. What particle is emitted when a Ra-221 nucleus decays to Rn-217?
    - (a) alpha
    - (b) beta
    - (c) neutron
    - (d) positron
    - (e) proton

3. What particle is emitted when a Th-228 nucleus decays to Ra-224?
    - (a) alpha
    - (b) beta
    - (c) neutron
    - (d) positron
    - (e) proton

4. What particle is emitted when a F-20 nucleus decays to Ne-20?
    - (a) alpha
    - (b) beta
    - (c) neutron
    - (d) positron
    - (e) proton

5. What particle is emitted when an Ar-39 nucleus decays to K-39?
    - (a) alpha
    - (b) beta
    - (c) neutron
    - (d) positron
    - (e) proton

6. What particle is emitted when a Sr-90 nucleus decays radioactively to Y-90?
    - (a) alpha
    - (b) beta
    - (c) neutron
    - (d) positron
    - (e) proton

7. What particle is emitted when a carbon-11 nucleus decays to boron-11?
   (a) alpha
   (b) beta
   (c) neutron
   (d) positron
   (e) proton

8. What particle is emitted when a fluorine-17 nucleus decays to oxygen-17?
   (a) alpha
   (b) beta
   (c) neutron
   (d) positron
   (e) proton

9. What particle is emitted when a neon-19 nucleus decays to fluorine-19?
   (a) alpha
   (b) beta
   (c) neutron
   (d) positron
   (e) proton

10. What nuclide is produced when Pt-175 decays by alpha emission?
   (a) $^{171}_{76}\text{Os}$
   (b) $^{175}_{76}\text{Os}$
   (c) $^{171}_{78}\text{Pt}$
   (d) $^{171}_{79}\text{Au}$
   (e) $^{175}_{79}\text{Au}$

11. What nuclide is produced when U-235 decays by alpha emission?
   (a) $^{231}_{90}\text{Th}$
   (b) $^{235}_{90}\text{Th}$
   (c) $^{231}_{92}\text{U}$
   (d) $^{235}_{93}\text{Np}$
   (e) $^{231}_{93}\text{Np}$

12. What nuclide is produced when Ra-223 decays by alpha and gamma emission?
   (a) $^{219}_{86}\text{Rn}$
   (b) $^{227}_{86}\text{Rn}$
   (c) $^{227}_{88}\text{Ra}$
   (d) $^{219}_{90}\text{Th}$
   (e) $^{227}_{90}\text{Th}$

13. What radionuclide decays to Pb-210 by alpha emission?

(a) $^{206}_{80}\text{Hg}$

(b) $^{214}_{80}\text{Hg}$

(c) $^{206}_{82}\text{Pb}$

(d) $^{206}_{84}\text{Po}$

(e) $^{214}_{84}\text{Po}$

14. What nuclide is produced when K-43 decays by beta emission?

(a) $^{43}_{18}\text{Ar}$

(b) $^{42}_{19}\text{K}$

(c) $^{42}_{20}\text{Ca}$

(d) $^{43}_{20}\text{Ca}$

(e) $^{44}_{20}\text{Ca}$

15. What nuclide is produced when Pb-210 decays by beta emission?

(a) $^{210}_{81}\text{Tl}$

(b) $^{212}_{81}\text{Tl}$

(c) $^{211}_{82}\text{Pb}$

(d) $^{210}_{83}\text{Bi}$

(e) $^{211}_{83}\text{Bi}$

16. What nuclide is produced when Ar-39 decays by beta and gamma emission?

(a) $^{39}_{17}\text{Cl}$

(b) $^{40}_{17}\text{Cl}$

(c) $^{40}_{18}\text{Ar}$

(d) $^{39}_{19}\text{K}$

(e) $^{40}_{19}\text{K}$

17. What radionuclide decays to Fe-56 by beta emission?

(a) $^{55}_{25}\text{Mn}$

(b) $^{56}_{25}\text{Mn}$

(c) $^{55}_{26}\text{Fe}$

(d) $^{56}_{27}\text{Co}$

(e) $^{57}_{27}\text{Co}$

18. What nuclide is produced when N-13 decays by positron emission?

(a) $^{12}_{6}\mathrm{C}$

(b) $^{13}_{6}\mathrm{C}$

(c) $^{14}_{6}\mathrm{C}$

(d) $^{14}_{7}\mathrm{N}$

(e) $^{13}_{8}\mathrm{O}$

19. What nuclide is produced when O-15 decays by positron emission?

(a) $^{14}_{7}\mathrm{N}$

(b) $^{15}_{7}\mathrm{N}$

(c) $^{14}_{8}\mathrm{O}$

(d) $^{15}_{9}\mathrm{F}$

(e) $^{16}_{9}\mathrm{F}$

20. What nuclide is produced when K-40 decays by positron emission?

(a) $^{39}_{18}\mathrm{Ar}$

(b) $^{40}_{18}\mathrm{Ar}$

(c) $^{41}_{18}\mathrm{Ar}$

(d) $^{40}_{19}\mathrm{K}$

(e) $^{40}_{20}\mathrm{Ca}$

21. What radionuclide decays to Br-73 by positron emission?

(a) $^{72}_{34}\mathrm{Se}$

(b) $^{74}_{34}\mathrm{Se}$

(c) $^{72}_{35}\mathrm{Br}$

(d) $^{74}_{35}\mathrm{Br}$

(e) $^{73}_{36}\mathrm{Kr}$

22. What nuclide is produced when a Cs-129 nucleus decays by electron capture?

(a) $^{128}_{54}\mathrm{Xe}$

(b) $^{129}_{54}\mathrm{Xe}$

(c) $^{128}_{55}\mathrm{Cs}$

(d) $^{128}_{56}\mathrm{Ba}$

(e) $^{129}_{56}\mathrm{Ba}$

23. What nuclide is produced when a W-181 nucleus decays by electron capture?

- (a) $^{180}_{73}\text{Ta}$
- (b) $^{181}_{73}\text{Ta}$
- (c) $^{180}_{74}\text{W}$
- (d) $^{180}_{75}\text{Re}$
- (e) $^{181}_{75}\text{Re}$

24. What nuclide is produced when a Mn-52 nucleus decays by electron capture?

- (a) $^{52}_{24}\text{Cr}$
- (b) $^{53}_{24}\text{Cr}$
- (c) $^{53}_{25}\text{Mn}$
- (d) $^{52}_{26}\text{Fe}$
- (e) $^{53}_{26}\text{Fe}$

25. What radionuclide decays to Cs-133 by electron capture?

- (a) $^{132}_{54}\text{Xe}$
- (b) $^{133}_{54}\text{Xe}$
- (c) $^{134}_{55}\text{Cs}$
- (d) $^{133}_{56}\text{Ba}$
- (e) $^{134}_{56}\text{Ba}$

Radioactive Decay Series

26. In the final step of the uranium-238 disintegration series, the parent nuclide decays to lead-206 and an alpha particle. What is the parent nuclide?

- (a) $^{202}_{80}\text{Hg}$
- (b) $^{210}_{83}\text{Bi}$
- (c) $^{206}_{84}\text{Po}$
- (d) $^{210}_{84}\text{Po}$
- (e) none of the above

27. In the final step of the uranium-235 disintegration series, the parent nuclide decays to lead-207 and a beta particle. What is the parent nuclide?

- (a) $^{207}_{81}\text{Tl}$
- (b) $^{206}_{82}\text{Pb}$
- (c) $^{208}_{82}\text{Pb}$
- (d) $^{207}_{83}\text{Bi}$
- (e) none of the above

28. In the final step of the thorium-232 disintegration series, the parent nuclide decays to lead-208 and an alpha particle. What is the parent nuclide?

(a) $^{208}_{83}\text{Bi}$
(b) $^{212}_{83}\text{Bi}$
(c) $^{208}_{84}\text{Po}$
(d) $^{212}_{84}\text{Po}$
(e) none of the above

29. The uranium-238 decay series begins with the emission of an alpha particle. If the daughter decays by beta emission, what is the resulting nuclide?

(a) $^{234}_{89}\text{Ac}$
(b) $^{233}_{90}\text{Th}$
(c) $^{234}_{90}\text{Th}$
(d) $^{233}_{91}\text{Pa}$
(e) $^{234}_{91}\text{Pa}$

30. The uranium-235 decay series begins with the emission of an alpha particle. If the daughter decays by beta emission, what is the resulting nuclide?

(a) $^{231}_{89}\text{Ac}$
(b) $^{230}_{90}\text{Th}$
(c) $^{231}_{90}\text{Th}$
(d) $^{230}_{91}\text{Pa}$
(e) $^{231}_{91}\text{Pa}$

31. The thorium-232 decay series begins with the emission of an alpha particle. If the daughter decays by beta emission, what is the resulting nuclide?

(a) $^{228}_{87}\text{Fr}$
(b) $^{227}_{88}\text{Ra}$
(c) $^{228}_{88}\text{Ra}$
(d) $^{227}_{89}\text{Ac}$
(e) $^{228}_{89}\text{Ac}$

Answer Key

1A, 2A, 3A, 4B, 5B, 6B, 7D, 8D, 9D, 10A, 11A, 12A, 13E, 14D, 15D, 16D, 17B, 18B, 19B, 20B, 21E, 22B, , 23B, 24A, 25D, 26D, 27A, 28D, 29E, 30E, 31E

CH 222
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.

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1. Give the number of **core** *and* **valence** electrons in the following elements: Li, Te, and Ga. (6 pts)

2. Write Lewis Dot structures for the following molecules: BeI_2, CBr_2Cl_2, and AsI_3. (6 pts)

3. Draw and name the electron-pair geometry and molecular shape for AlF_3 and AlF_4^-. (4 pts)

4. Determine the formal charge on each atom in the molecule ClF_2^{-1}. (4 pts)

CH 222
Sample Quiz #1 Name: ____ Lab Section: _______
Answers

---

1. Give the number of **core** *and* **valence** electrons in the following elements: Li, Te, and Ga. (6 pts)

   Li: 2 core, 1 valence
   Te: 46 core, 6 valence
   Ga: 28 core, 3 valence

2. Write Lewis Dot structures for the following molecules: BeI_2, CBr_2Cl_2, and AsI_3. (6 pts)

   BeI_2: linear EPG and MG, no lone pairs
   CBr_2Cl_2: Tetrahedral for both EPG and MG, lone pairs around outer atoms
   AsI_3: tetrahedral EPG, trigonal pyramid MG, 1 lone pair on As

3. Draw and name the electron-pair geometry and molecular shape for AlF_3 and AlF_4^-. (4 pts)

   AlF_3: trigonal planar for both EPG and MG, no lone pairs
   AlF_4^{-1}: tetrahedral for both EPG and MG

4. Determine the formal charge on each atom in the molecule ClF_2^{-1}. (4 pts)

   ClF_2^{-1}: EPG is trigonal bipyramid, MG is linear, Cl has 3 lone pairs
   Cl: 7 - 6 - 2 = -1
   (both) F: 7 - 6 - 1 = 0

CH 222
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.

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1. Complete the following table: (12 points)

Molecule	**Lewis Structure**	**Electron Pair Geometry & Molecular Geometry**	**Bond Angle(s)**	**Polar or Nonpolar?**
ICl_2Br				
$SeBr_4$				
IF_5				

2. Complete the following table: (6 points)

Molecule	**Lewis Structure**	**Bond Order**	**Which structure has stronger bonds?**	**Which structure has longer bonds?**	**Does it exhibit resonance?**
PO_2^{-1}					
PO_2^{+1}					

3. Arrange the following bonds in order of increasing polarity: Na-Br, Br-Br, N-Br, As-Br. (2 pts)

CH 222
Sample Quiz #2 Name: ______________________________ Lab Section: _______
Answers

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1. Complete the following table: (12 points)

Molecule	**Lewis Structure**	**Electron Pair Geometry & Molecular Geometry**	**Bond Angle(s)**	**Polar or Nonpolar?**
ICl_2Br	**I in middle, 2 lone pairs on I, Cl-I-Cl has 180 angle**	**EPG = trigonal bipyramid MG = T-shaped**	**90, 180**	**polar**
$SeBr_4$	**Se in middle, 1 lone pair on Se, lone pair in "plane" of molecule**	**EPG = trigonal bipyramid MG = see-saw**	**90, 120, 180**	**polar**
IF_5	**I in middle, 1 lone pair on I,**	**EPG = octahedral MG = square pyramid**	**90, 180**	**polar**

2. Complete the following table: (6 points)

Molecule	**Lewis Structure**	**Bond Order**	**Which structure has stronger bonds?**	**Which structure has longer bonds?**	**Does it exhibit resonance?**
PO_2^{-1}	**P in middle, 1 lone pair on P, one double bond to O, one single bond to O**	**1.5**		**longer**	**yes**
PO_2^{+1}	**P in middle, 0 lone pair on P, two double bonds to O,**	**2**	**stronger**		**no**

3. Arrange the following bonds in order of increasing polarity: Na-Br, Br-Br, N-Br, As-Br. (2 pts)

Br-Br, As-Br, N-Br, Na-Br

CH 222
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.

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1. For the molecules B_2 and B_2^{2-}:
   a. Draw molecular orbital energy level diagrams. (6 points)
   b. For each molecule determine the bond order *and* magnetic properties (para-/dia-magnetic) (2 points).
   c. Which has the longer bond? (1 point)
   d. Which has the highest bond energy? (1 point)

2. Provide names or structures for the following organic molecules: (10 pts)

3-methylhexane	2-pentanone	cyclobutyl bromide
structure:	*structure:*	*structure:*
$CH_3CH_2CH_2OH$	C_6H_6	CH_3-O-$CH_2CH_2CH_3$
2-heptyne	toluene	acetone
structure:	*structure:*	*structure:*
ethanal		
structure:		

CH 222
Sample Quiz #3 Name: ____ Lab Section: _______
Answers

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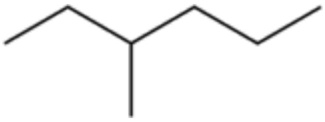
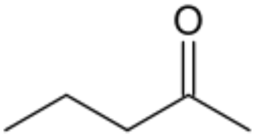
1. For the molecules B_2 and B_2^{2-}:
   a. Draw molecular orbital energy level diagrams. (6 points)
   b. For each molecule determine the bond order *and* magnetic properties (para-/dia-magnetic) (2 points).
   c. Which has the longer bond? (1 point)
   d. Which has the highest bond energy? (1 point)

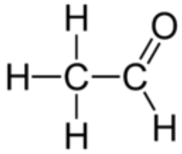
B_2: [core] $(\sigma_{2s})^2 (\sigma_{2s}^*)^2 (\pi_{2p})^2$ paramagnetic bo = 1 longest bond
B_2^{2-}: [core] $(\sigma_{2s})^2 (\sigma_{2s}^*)^2 (\pi_{2p})^4$ diamagnetic bo = 2 highest bond energy

2. Provide names or structures for the following organic molecules: (10 pts)

3-methylhexane	2-pentanone	cyclobutyl bromide
structure: 	*structure:* 	*structure:*

$CH_3CH_2CH_2OH$	C_6H_6	$CH_3\text{-}O\text{-}CH_2CH_2CH_3$
1-propanol	**benzene**	**1-methoxypropane**

2-heptyne	toluene	acetone
structure:	*structure:*	*structure:*

ethanal
structure: 

CH 222
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.

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1. Ammonia (NH_3) is synthesized through the combination of hydrogen and nitrogen gases. What mass of nitrogen gas must be reacted to create ammonia at 3.00 atm in a 2.50 L flask at 100. °C? *Hint:* write a balanced equation! (6 pts)

2. Calculate the density (g/L) of gaseous phosphorus pentachloride at STP to three significant figures. (5 pts)

3. What volume of O_2, measured at 44.3 °C and 766 mm Hg, will be produced by the decomposition of 4.27 g $NaClO_3$? (5 points)
   $2\ NaClO_3(s) \rightarrow 2\ NaCl(s) + 3\ O_2(g)$

4. The lid is tightly sealed on a rigid flask containing 3.50 L H_2 at 17.0 °C and 694 torr. If the flask is heated to 71 °C, what is the pressure in the flask in atm? (4 points)

CH 222
Sample Quiz #4 Name: ____ Lab Section: _______
Answers

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1. Ammonia (NH_3) is synthesized through the combination of hydrogen and nitrogen gases. What mass of nitrogen gas must be reacted to create ammonia at 3.00 atm in a 2.50 L flask at 100. °C? *Hint:* write a balanced equation! (6 pts)

   3.43 g

2. Calculate the density (g/L) of gaseous phosphorus pentachloride at STP to three significant figures. (5 pts)

   9.30 g/L

3. What volume of O_2, measured at 44.3 °C and 766 mm Hg, will be produced by the decomposition of 4.27 g $NaClO_3$? (5 points)
   $\mathbf{2\ NaClO_3(s) \rightarrow 2\ NaCl(s) + 3\ O_2(g)}$

   V = 1.56 L

4. The lid is tightly sealed on a rigid flask containing 3.50 L H_2 at 17.0 °C and 694 torr. If the flask is heated to 71 °C, what is the pressure in the flask in atm? (4 points)

   P_2 = 1.08 atm *also: 823 torr*

CH 222
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. Fill in the blanks using the following species: (1 point each)

   MgO, $LiCl$, NH_3, $CHCl_3$, $Kr(g)$, $Mg^{2+}(aq)$

   a. Which species will have no appreciable intermolecular forces? _____

   b. Which species will exhibit hydrogen bonding in the liquid state? _____

   c. Which species will have the highest melting point? _____

   d. Which has the lowest normal boiling point? _____

   e. Will species has a solvent and solute? _____

2. The molar enthalpy of vaporization for 2-pentanol is 46.26 kJ/mol at 71.0 ˚C, and the density is 0.9884 g/cm^3. How much energy is required to evaporate 2.25 L of 2-pentanol? (5 points)

3. A quantity (161 mmol, where 1 mmol = 10^{-3} mol) of a nonvolatile solute is dissolved in 0.101 kg of benzene (C_6H_6). The vapor pressure of pure benzene at 27 °C is 115.8 mm Hg. Find the mole fraction of the solute and the vapor pressure of the solution at 27 °C. (5 points)

4. Which will generate the higher osmotic pressure at 298 K: 1 L of a 0.250 M cesium sulfate solution or 1 L of a 0.220 M phosphoric acid? Explain your reasoning, calculate the osmotic pressure for both solutions and assume 100% dissociation into ions. (5 points)

CH 222
Sample Quiz #5 Name: ____ Lab Section: _______
Answers

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1. Fill in the blanks using the following species: (1 point each)

MgO, $LiCl$, NH_3, $CHCl_3$, $Kr(g)$, $Mg^{2+}(aq)$

a. Which species will have no appreciable intermolecular forces? **Kr**

b. Which species will exhibit hydrogen bonding in the liquid state? **NH_3**

c. Which species will have the highest melting point? **MgO**

d. Which has the lowest normal boiling point? **Kr**

e. Will species has a solvent and solute? **$Mg^{2+}(aq)$**

2. The molar enthalpy of vaporization for 2-pentanol is 46.26 kJ/mol at 71.0 ˚C, and the density is 0.9884 g/cm^3. How much energy is required to evaporate 2.25 L of 2-pentanol? (5 points)

1170 kJ

3. A quantity (161 mmol, where 1 mmol = 10^{-3} mol) of a nonvolatile solute is dissolved in 0.101 kg of benzene (C_6H_6). The vapor pressure of pure benzene at 27 °C is 115.8 mm Hg. Find the mole fraction of the solute and the vapor pressure of the solution at 27 °C. (5 points)

χ(solute) = 0.111
VP = 103 mm Hg

4. Which will generate the higher osmotic pressure at 298 K: 1 L of a 0.250 M cesium sulfate solution or 1 L of a 0.220 M phosphoric acid? Explain your reasoning, calculate the osmotic pressure for both solutions and assume 100% dissociation into ions. (5 points)

for Cs_2SO_4: **π = 18.3 atm**
for H_2PO_4: **π = 21.5 atm**
H_3PO_4 is higher!

CH 222

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. Consider the data (below) gathered for the following reaction: **A + B → C** (8 points)

[A] (M)	[B] (M)	Δ[C]/Δ*t* (initial) M/s
0.100	0.200	6.80×10^{-6}
0.100	0.400	2.72×10^{-5}
0.200	0.400	5.44×10^{-5}

a. What is the order of the reaction with respect to A: ______ B: ______ Overall order: ______

b. What is the numerical value for the rate constant?

c. Write the rate law for the reaction.

d. What is the value of the rate when [A] = 0.337 M and [B] = 0.122 M?

2. In basic solution, $(CH_3)_3CCl$ reacts according to the equation below.

$$(CH_3)_3CCl + OH^- \rightarrow (CH_3)_3COH + Cl^-$$

The accepted mechanism for the reaction is

$$(CH_3)_3CCl \rightarrow (CH_3)_3C^+ + Cl^- \quad \text{(slow)}$$
$$(CH_3)_3C^+ + OH^- \rightarrow (CH_3)_3COH \quad \text{(fast)}$$

a. What is a rate law that is consistent with the mechanism for this reaction? (2 points)

b. Are intermediates present in the reaction? If so, list them. (2 points)

3. For a chemical reaction, the activation energy for the forward reaction is +187 kJ and the activation energy for the backward reaction is +112 kJ. What is the overall energy change for the forward reaction? (4 points)

4. What is the half-life of a first order reaction with a rate constant of 0.457 s^{-1}? (4 points)

CH 222
Sample Quiz #6 Name: ____ Lab Section: _______
Answers

---

1. Consider the data (below) gathered for the following reaction: **A + B → C** (8 points)

[A] (M)	[B] (M)	Δ[C]/Δ*t* (initial) M/s
0.100	0.200	6.80×10^{-6}
0.100	0.400	2.72×10^{-5}
0.200	0.400	5.44×10^{-5}

a. What is the order of the reaction with respect to A: _**1**_ B: _**2**_ Overall order: _**3**_

b. What is the numerical value for the rate constant?

k = 1.70 x 10^{-3}

c. Write the rate law for the reaction. **rate = k[A][B]^2**

d. What is the value of the rate when [A] = 0.337 M and [B] = 0.122 M?

rate = 8.53 x 10^{-6}

2. In basic solution, $(CH_3)_3CCl$ reacts according to the equation below.

$$(CH_3)_3CCl + OH^- \rightarrow (CH_3)_3COH + Cl^-$$

The accepted mechanism for the reaction is

$(CH_3)_3CCl \rightarrow (CH_3)_3C^+ + Cl^-$ (slow)
$(CH_3)_3C^+ + OH^- \rightarrow (CH_3)_3COH$ (fast)

a. What is a rate law that is consistent with the mechanism for this reaction? (2 points)

rate = k [$(CH_3)_3CCl$] ***slow step***

b. Are intermediates present in the reaction? If so, list them. (2 points)

yes, $(CH_3)_3C^+$

3. For a chemical reaction, the activation energy for the forward reaction is +187 kJ and the activation energy for the backward reaction is +112 kJ. What is the overall energy change for the forward reaction? (4 points)

forward reaction is endothermic (reverse reaction has lower activation energy)
$\Delta H = 187 - 112 = +75$ kJ

4. What is the half-life of a first order reaction with a rate constant of 0.457 s^{-1}? (4 points)

$t_{1/2} = 1.52$ s

CH 222 Sample Exam
Exam I Name: ____ Lab Section: ______

Part I: Multiple Choice Questions (100 Points) There is *only* one best answer for each question.

---

1. Which of the following statements are CORRECT?
    1. Ionic bonds form when one or more valence electrons are transferred from one atom to another.
    2. Covalent bonds involve sharing of electrons between atoms.
    3. Ionic bond formation is always exothermic; covalent bond formation is always endothermic.

    a. 1 only
    b. 2 only
    c. 3 only
    d. 1 and 2
    e. 1, 2, and 3

2. Which of the following compounds is expected to have the strongest ionic bonds?

    a. MgO
    b. KBr
    c. NaI
    d. SrO
    e. CaS

3. What is the total number of valence electrons in a carbonate ion?

    a. 20
    b. 22
    c. 24
    d. 26
    e. 30

4. Which of the following species will have a Lewis structure most like that of the hydronium ion, H_3O^+?

    a. NO_3^-
    b. NH_3
    c. SO_3
    d. CO_3^{2-}
    e. H_2CO

5. How many resonance structures can be drawn for the thiocyanate ion, SCN^{-1}? The carbon atom is in the center of this ion.

    a. 1
    b. 2
    c. 3
    d. 4
    e. 5

6. The central atom in XeF_4 is surrounded by

    a. 3 single bonds, 1 double bond, and no lone pairs of electrons.
    b. 2 single bonds, 2 double bonds, and no lone pairs of electrons.
    c. 3 single bonds, 1 double bond, and 1 lone pair of electrons.
    d. 4 single bonds, no double bonds, and no lone pairs of electrons.
    e. 4 single bonds, no double bonds, and 2 lone pairs of electrons.

7. Use VSEPR theory to predict the electron-pair geometry and the molecular geometry of iodine trichloride, ICl_3.

   a. The e--pair geometry is trigonal-planar, the molecular geometry is trigonal-planar.
   b. The e--pair geometry is tetrahedral, the molecular geometry is trigonal-pyramidal.
   c. The e--pair geometry is tetrahedral, the molecular geometry is trigonal-planar.
   d. The e--pair geometry is trigonal-bipyramidal, the molecular geometry is T-shaped.
   e. The e--pair geometry is trigonal-bipyramidal, the molecular geometry is trigonal-planar.

8. What is the formal charge on each atom in a hypobromite ion, OBr^{-1}?

   a. O = -2, Br = -1
   b. O = -2, Br = +1
   c. O = -1, Br = +1
   d. O = -1, Br = 0
   e. O = 0, Br = -1

9. When heated, azomethane decomposes into nitrogen and ethane gas: $\mathbf{CH_3N{=}NCH_3(g) \rightarrow N_2(g) + C_2H_6(g)}$ Using this equation and the table of bond enthalpies below, calculate the enthalpy of reaction.

Bond	Bond Energy (kJ/mol)	Bond	Bond Energy (kJ/mol)
C-H	413	N-N	163
C-N	305	N=N	418
C-C	346	N ≡N	945

   a. -611 kJ
   b. -527 kJ
   c. -429 kJ
   d. -313 kJ
   e. -263 kJ

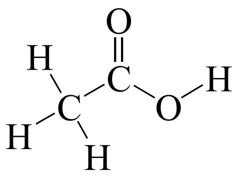
10. Based on bond order, predict which molecule has the shortest carbon-oxygen bond length.

   a. CO
   b. CO_2
   c. CH_3OH
   d. H_2CO
   e. O_3

11. All of the following statements concerning valence bond (VB) and molecular orbital (MO) bond theories are correct EXCEPT

   a. MO theory predicts that electrons are delocalized over the molecule.
   b. in VB theory, bonding electrons are localized between pairs of atoms.
   c. VB theory describes a molecular bond as the overlap between two atomic or hybrid orbitals.
   d. MO theory can describe molecular bonding in excited states.
   e. VB theory is used to predict the colors of compounds.

12. How many sigma (σ) bonds and pi (π) bonds are in acetic acid?



a. six σ and one π
b. six σ and two π
c. seven σ and one π
d. eight σ and zero π
e. eight σ and one π

13. What is the hybridization of the sulfur atom in SF_4?

a. sp
b. sp^2
c. sp^3
d. sp^3d
e. sp^3d^2

14. In which of the following molecules and ions does the central carbon atom have sp hybridization: Cl_2CO, CH_2Br_2, CO_2, and OCN^{-1}?

a. Cl_2CO only
b. Cl_2CO and CH_2Br_2
c. CH_2Br_2 and CO_2
d. CH_2Br_2 and OCN^{-1}
e. CO_2 and OCN^{-1}

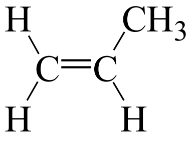
15. What is the molecular geometry around a central atom that is sp^2 hybridized, has three sigma bonds, and has one pi bond?

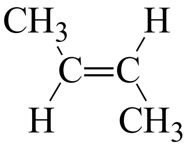
a. trigonal-planar
b. trigonal-pyramidal
c. trigonal-bipyramidal
d. linear
e. tetrahedral

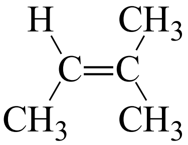
16. Carbon dioxide reacts with an aqueous solution of sodium hydroxide to form carbonate ion. What change in the hybridization of carbon occurs in this reaction?

a. sp to sp^2
b. sp^2 to sp^3
c. sp^3 to sp^3d
d. sp^3 to sp^3d^2
e. no change

17. For which of the following compounds is it possible for cis and trans isomers to exist?


(1)


(2)


(3)

a. 1 only
b. 2 only
c. 3 only
d. 1 and 2
e. 1, 2, and 3

18. Atomic orbitals combine most effectively to form molecular orbitals when

a. electrons in the orbitals have no spins.
b. electrons in the orbitals have the same spin.
c. the atomic orbitals are hybridized.
d. the atomic orbitals have similar energies.
e. metals combine with nonmetals.

19. According to molecular orbital theory, which of the following species is the most likely to exist?

a. H_2^{2-}
b. He_2
c. Li_2
d. Li_2^{2-}
e. Be_2

20. According to molecular orbital theory, what is the bond order of superoxide, O_2^{-1}?

a. 1
b. 3/2
c. 2
d. 5/2
e. 3

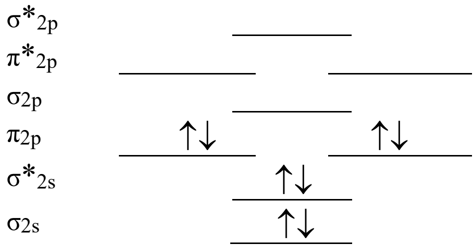
21. Use molecular orbital theory to predict which ion is diamagnetic.

a. C_2^{-2}
b. O_2
c. NO
d. N_2^{-1}
e. B_2^{-1}

22. What is the molecular orbital configuration of CO?

a. [core electrons] $(\sigma_{2s})^2 (\sigma^*_{2s})^2 (\pi_{2p})^4$
b. [core electrons] $(\sigma_{2s})^2 (\sigma^*_{2s})^2 (\pi_{2p})^2 (\sigma_{2p})^2 (\pi^*_{2p})^2$
c. [core electrons] $(\sigma_{2s})^2 (\sigma^*_{2s})^2 (\pi_{2p})^4 (\sigma_{2p})^2$
d. [core electrons] $(\sigma_{2s})^2 (\sigma^*_{2s})^2 (\sigma_{2p})^2 (\pi_{2p})^4$
e. [core electrons] $(\sigma_{2s})^2 (\sigma^*_{2s})^2 (\pi_{2p})^2$

23. Which molecule will have the following valence molecular orbital level energy diagram?



   a. Li_2
   b. Be_2
   c. B_2
   d. C_2
   e. N_2

24. Which of the following molecules may be a cycloalkane?

   a. C_3H_8
   b. C_4H_6
   c. C_5H_{12}
   d. C_6H_{12}
   e. C_7H_{16}

25. What is the molecular formula for heptane?

   a. C_6H_{12}
   b. C_6H_{14}
   c. C_7H_{14}
   d. C_7H_{16}
   e. C_8H_{14}

Part II: Short Answer / Calculation. *Show all work!*

1. For each of the following molecules or ions,
   - i) draw the Lewis structure
   - ii) give the hybridization of the central atom
   - iii) predict the electron pair *and* molecular shape of the molecule, and
   - iv) state if the molecule is polar or nonpolar

a) ICl_3 **b)** $TeBr_2$ **c)** XeF_4 **d)** BrF_2^{-1} **e)** I_3^{-1}

2. Draw molecular orbital energy diagrams for N_2, N_2^{1+} and N_2^{1-}. Determine the bond order and indicate if each molecule is paramagnetic or diamagnetic. Indicate which of the molecules will have the shortest bond length.

3. Provide the correct name for each of the following compounds.

C_2H_6 ______________________________

C_2H_4 ______________________________

C_2H_2 ______________________________

CH_3CH_2OH ______________________________

CH_3OCH_3 ______________________________

CH_3COCH_3 ______________________________

CH_3CHO ______________________________

C_6H_6 ______________________________

CH_3OH ______________________________

NH_3 ______________________________

CH 222 Sample Exam
Exam I Name: ____ Lab Section: ______
Answers

Part I: Multiple Choice Questions

1. D
2. A
3. C
4. B
5. C
6. E

7. D
8. D
9. E
10. A
11. E

12. C
13. D
14. E
15. A
16. A

17. B
18. D
19. C
20. B
21. A
22. D

23. D
24. D
25. D

Part II: Short Answer / Calculation.

1. Lewis structures:
   - **a. ICl_3: trigonal bipyramid EPG, T-shape MG, *dsp^3*, polar**
   - **b. $TeBr_2$: tetrahedral EPG, bent MG, *sp^3*, polar**
   - **c. XeF_4: octahedral EPG, square planar MG, *d^2sp^3*, nonpolar**
   - **d. BrF_2^{-1}: trigonal bipyramid EPG, linear MG, *dsp^3*, nonpolar**
   - **e. I_3^{-1}: trigonal bipyramid EPG, linear MG, *dsp^3*, nonpolar**

2. Molecular orbitals:
   - **N_2: [core electrons] $(\sigma_{2s})^2$ $(\sigma^*_{2s})^2$ $(\pi_{2p})^4$ $(\sigma_{2p})^2$ *bond order* = 3, diamagnetic, *shortest bond length***
   - **N_2^{+1}: [core electrons] $(\sigma_{2s})^2$ $(\sigma^*_{2s})^2$ $(\pi_{2p})^4$ $(\sigma_{2p})^1$ *bond order* = 2.5, paramagnetic**
   - **N_2^{-1}: [core electrons] $(\sigma_{2s})^2$ $(\sigma^*_{2s})^2$ $(\pi_{2p})^4$ $(\sigma_{2p})^2$ $(\pi^*_{2p})^1$ *bond order* = 2.5, paramagnetic**
3. *names:*
   - a. ethane
   - b. ethene
   - c. ethyne or acetylene
   - d. ethanol
   - e. methoxy methane or dimethyl ether
   - f. propanone or acetone
   - g. ethanal
   - h. benzene
   - i. methanol
   - j. ammonia

CH 222 Sample Exam
Exam II Name: ____ Lab Section: ______

Part I: Multiple Choice Questions (100 Points) There is *only* one best answer for each question.

---

1. A sample of gas (24.2 g) initially at 4.00 atm was compressed from 8.00 L to 2.00 L at constant temperature. After the compression, the gas pressure was __________ atm.

   a. 4.00
   b. 2.00
   c. 1.00
   d. 8.00
   e. 16.0

2. A balloon originally had a volume of 4.39 L at 44 °C and a pressure of 729 torr. The balloon must be cooled to __________°C to reduce its volume to 3.78 L (at constant pressure).

   a. 38
   b. 0
   c. 72.9
   d. 273
   e. 546

3. If 50.75 g of a gas occupies 10.0 L at STP, 129.3 g of the gas will occupy __________ L at STP.

   a. 3.92
   b. 50.8
   c. 12.9
   d. 25.5
   e. 5.08

4. The reaction of 50 mL of Cl_2 gas with 50 mL of CH_4 gas via the equation below will produce a total of _____ mL of products (assume pressure and temperature are kept constant.)

$$Cl_2(g) + CH_4(g) \rightarrow HCl(g) + CH_3Cl(g)$$

   a. 100
   b. 50
   c. 200
   d. 150
   e. 250

5. The pressure of a sample of CH_4 gas (6.022 g) in a 30.0 L vessel at 402 K is __________ atm.

   a. 2.42
   b. 6.62
   c. 0.413
   d. 12.4
   e. 22.4

6. The density of N_2O at 1.53 atm and 45.2 °C is __________ g/L.

   a. 18.2
   b. 1.76
   c. 0.388
   d. 9.99
   e. 2.58

7. Automobile air bags use the decomposition of sodium azide as their source of gas for rapid inflation per the reaction below. What mass (g) of NaN_3 is required to provide 40.0 L of N_2 at 25.0 °C and 763 torr?

$$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

   a. 1.64
   b. 1.09
   c. 160
   d. 71.1
   e. 107

8. A vessel contained N_2, Ar, He, and Ne. The total pressure in the vessel was 987 torr. The partial pressures of nitrogen, argon, and helium were 44.0, 486, and 218 torr, respectively. The partial pressure of neon in the vessel was __________ torr.

   a. 42.4
   b. 521
   c. 19.4
   d. 239
   e. 760

9. Of the following, __________ has the highest boiling point.

   a. N_2
   b. Br_2
   c. H_2
   d. Cl_2
   e. O_2

10. Of the following, __________ is an exothermic process.

   a. melting
   b. subliming
   c. freezing
   d. boiling
   e. All of the above are exothermic.

11. The heat of fusion of water is 333 J/g. The conversion of 50.0 g of ice at 0.0 °C to liquid water at 22.0 °C requires how many kilojoules (kJ) of heat?

   a. 3.8×10^2
   b. 21.3
   c. 17.2
   d. 0.469
   e. Insufficient data are given.

12. Of the following, __________ is the most volatile.

   a. CBr_4
   b. CCl_4
   c. CF_4
   d. CH_4
   e. C_6H_{14}

13. Potassium metal crystallizes in a body-centered cubic structure with a unit cell edge length of 5.31 Å. The radius of a potassium atom is __________ Å.

    a. 1.33
    b. 1.88
    c. 2.30
    d. 2.66
    e. 5.31

14. As a solid element melts, the atoms become __________ and they have __________ attraction for one another.

    a. more separated, more
    b. more separated, less
    c. closer together, more
    d. closer together, less
    e. larger, greater

15. Which one of the following exhibits dipole-dipole attraction between molecules?

    a. XeF_4
    b. AsH_3
    c. CO_2
    d. BCl_3
    e. Cl_2

16. Based on the following information, which compound has the strongest intermolecular forces?

Substance	**ΔH_{vap} (kJ/mol)**
Argon (Ar)	6.3
Benzene (C_6H_6)	31.0
Ethanol (C_2H_5OH)	39.3
Water (H_2O)	40.8
Methane (CH_4)	9.2

    a. Argon
    b. Benzene
    c. Ethanol
    d. Water
    e. Methane

17. The vapor pressure of any substance at its normal boiling point is

    a. 1 bar
    b. 1 torr
    c. 1 atm
    d. equal to atmospheric pressure
    e. equal to the vapor pressure of water

18. The solubility of oxygen gas in water at 25 °C and 1.0 atm pressure of oxygen is 0.041 g/L. The solubility of oxygen in water at 3.0 atm and 25 °C is __________ g/L.

    a. 0.041
    b. 0.014
    c. 0.31
    d. 0.12
    e. 3.0

19. A sample of potassium nitrate (49.0 g) is dissolved in 101 g of water at 100 °C with precautions taken to avoid evaporation of any water. The solution is cooled to 30.0 °C and a small amount of precipitate is observed. This solution is __________.

    a. hydrated
    b. placated
    c. saturated
    d. unsaturated
    e. supersaturated

20. The concentration of urea in a solution prepared by dissolving 16 g of urea in 39 g of H_2O is __________% by mass. The molar mass of urea is 60.0 g/mol.

    a. 29
    b. 41
    c. 0.29
    d. 0.41
    e. 0.48

21. The concentration of KBr in a solution prepared by dissolving 2.21 g of KBr in 897 g of water is __________ molal.

    a. 2.46
    b. 0.0167
    c. 0.0207
    d. 2.07×10^{-5}
    e. 0.0186

22. A solution is prepared by dissolving 15.0 g of NH_3 in 250.0 g of water. The density of the resulting solution is 0.974 g/mL. The molarity of NH_3 in the solution is __________.

    a. 0.00353
    b. 0.882
    c. 60.0
    d. 3.24
    e. 3.53

23. The concentration of sodium chloride in an aqueous solution that is 2.23 M and that has a density of 1.01 g/mL is __________% by mass.

    a. 2.21
    b. 7.83
    c. 45.3
    d. 12.9
    e. 10.1

24. The vapor pressure of pure water at 25 °C is 23.8 torr. What is the vapor pressure (torr) of water above a solution prepared by dissolving 18.0 g of glucose (a nonelectrolyte, MW = 180.0 g/mol) in 95.0 g of water?

    a. 24.3
    b. 23.4
    c. 0.451
    d. 0.443
    e. 23.8

25. Determine the freezing point (°C) of a 0.015 molal aqueous solution of $MgSO_4$. The molal freezing-point-depression constant of water is 1.86 °C/m. *Note:* Check your van't Hoff factor!

    a. -0.056
    b. -0.028
    c. -0.17
    d. -0.084
    e. 0.000

Part II: Short Answer / Calculation. *Show all work!*

1. A 1.44-g sample of an unknown pure elemental gas occupies a volume of 0.335 L at a pressure of 1.00 atm and a temperature of 100.0 °C. Use this information to determine the identity of the unknown gas. (10 points)

Part II: Short Answer / Calculation *(continued) Show all work!*

2. The fluorocarbon $C_2Cl_3F_3$ has a normal boiling point of 47.6 °C. The specific heats of $C_2Cl_3F_3(l)$ and $C_2Cl_3F_3(g)$ are 0.910 J/g-K and 0.670 J/g-K, respectively. The heat of vaporization of the compound is 27.49 kJ/mol. Calculate the heat required to convert 50.0 g of the compound from the liquid at 5.0 °C to the gas at 80.0 °C in **kilojoules** (kJ). (10 points)

3. What is the molar mass of a nonelectrolyte if 6.02 grams dissolved in 30.0 grams of benzene freezes at –1.55 °C? The freezing point of pure benzene is 5.50 °C and the freezing point depression constant, K_{fp}, is –5.12 °C/m. (10 points)

Extra Credit Question: The nonelectrolyte compound in problem #3 is found to have the following: **49.31%C, 6.90 %H, and 43.79 %O**. Determine the **empirical and molecular formulas** for the unknown compound. (5 points)

CH 222 Sample Exam
Exam II		Name: ____	Lab Section: ______
Answers

Part I: Multiple Choice Questions

1. E
2. B
3. D
4. A
5. C
6. E

7. D
8. D
9. B
10. C
11. B
12. D

13. C
14. B
15. B
16. D
17. C
18. D

19. C
20. A
21. C
22. D
23. D
24. B

25. A

Part II: Short Answer / Calculation.

1. xenon gas

2. 10.36 kJ

3. 146 g/mol

4. *(bonus question)* $C_3H_5O_2$ (EF), $C_6H_{10}O_4$ (MF)

CH 222 Sample Final Exam Name: ____ Lab Section: ______

Multiple Choice Questions (160 Points) There is *only* one best answer for each question. Good luck!

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1. Atomic number describes the number of ___________ in an atom.

    a. protons
    b. neutrons
    c. electrons
    d. a and b
    e. all of the above

2. Which of the following symbols represents an alpha particle?

    a. ${}^{4}_{2}\mathrm{He}$
    b. ${}^{2}_{4}\mathrm{He}$
    c. ${}^{0}_{+1}\mathrm{e}$
    d. ${}^{0}_{-1}\mathrm{e}$
    e. ${}^{1}_{0}\mathrm{n}$

3. The most penetrating type of radiation is a(n)

    a. alpha particle
    b. beta particle
    c. gamma ray
    d. positron
    e. cathode ray

4. If plutonium-244 decays by successive **α, β, β, α** emissions, what nucleus is produced?

    a. ${}^{236}_{88}\mathrm{Ra}$
    b. ${}^{236}_{89}\mathrm{Ac}$
    c. ${}^{236}_{90}\mathrm{Th}$
    d. ${}^{240}_{90}\mathrm{Th}$
    e. ${}^{236}_{92}\mathrm{U}$

5. If Ag-106 decays by electron capture, what is the product?

    a. ${}^{105}_{46}\mathrm{Pd}$
    b. ${}^{106}_{46}\mathrm{Pd}$
    c. ${}^{105}_{47}\mathrm{Ag}$
    d. ${}^{106}_{48}\mathrm{Cd}$
    e. ${}^{107}_{47}\mathrm{Ag}$

6. By what (single step) process does polonium-211 decay to lead-207?

   a. α particle emission
   b. β particle emission
   c. positron emission
   d. electron capture
   e. neutron capture

7. The decay of radioactive elements is a first-order process. The half-life of carbon-14 is 5730 years. How many years will it take for 5.0 g of carbon-14 to decay to 1.0 mg?

   a. 5730 years
   b. 17,200 years
   c. 24,900 years
   d. 57,300 years
   e. 70,400 years

8. Given the initial rate data for the reaction A + B → C, determine the rate expression for the reaction.

[A], M	[B], M	Δ[C]/Δ*t* (initial) M/s
0.334	0.134	4.11×10^{-9}
0.334	0.187	8.00×10^{-9}
0.668	0.134	4.11×10^{-9}

   a. $\frac{\Delta[\mathrm{C}]}{\Delta t} = 2.75 \times 10^{-7}\ \mathrm{M^{-2}s^{-1}[A]^2[B]}$
   b. $\frac{\Delta[\mathrm{C}]}{\Delta t} = 3.07 \times 10^{-8}\ \mathrm{s^{-1}[B]}$
   c. $\frac{\Delta[\mathrm{C}]}{\Delta t} = 2.29 \times 10^{-7}\ \mathrm{M^{-1}s^{-1}[B]^2}$
   d. $\frac{\Delta[\mathrm{C}]}{\Delta t} = 6.85 \times 10^{-7}\ \mathrm{M^{-2}s^{-1}[A][B]^2}$
   e. $\frac{\Delta[\mathrm{C}]}{\Delta t} = 1.23 \times 10^{-8}\ \mathrm{s^{-1}[A]}$

9. For a zero order reaction, which of the following (if plotted versus time) should give a straight line?

   a. ln [A]
   b. ln k
   c. ln [1/A]
   d. 1/[A]
   e. [A]

10. In basic solution, $(CH_3)_3CCl$ reacts according to the equation:

    $(CH_3)_3CCl + OH^- \rightarrow (CH_3)_3COH + Cl^-$

    The accepted mechanism for the reaction is

    $(CH_3)_3CCl \rightarrow (CH_3)_3C^+ + Cl^-$ (slow)
    $(CH_3)_3C^+ + OH^- \rightarrow (CH_3)_3COH$ (fast)

    What is a rate law that is consistent with the mechanism for this reaction?

    a. rate = $k[(CH_3)_3CCl]$
    b. rate = $k[(CH_3)_3CCl][OH^-]$
    c. rate = $k[(CH_3)_3C^+][OH^-]$
    d. rate = $k[(CH_3)_3CCl][OH^-]/[Cl^-]$
    e. rate = $k[(CH_3)_3CCl][OH^-]/[Cl^-]$

11. The elementary steps for the catalyzed decomposition of dinitrogen monoxide are shown below.

    $2\ N_2O(g) + 2\ NO(g) \rightarrow 2\ N_2(g) + 2\ NO_2(g)$
    $2\ NO_2(g) \rightarrow 2\ NO(g) + O_2(g)$

    Which of the following statement(s) is/are CORRECT?
    1. The overall balanced reaction is $2\ N_2O(g) \rightarrow 2\ N_2(g) + O_2(g)$.
    2. $NO(g)$ is a catalyst for the reaction.
    3. $N_2(g)$ is a reaction intermediate.

    a. 1 only
    b. 2 only
    c. 3 only
    d. 1 and 2
    e. 1, 2, and 3

12. What is the half-life for a first-order reaction with a rate constant of $0.291\ s^{-1}$?

    a. 0.420 s
    b. 1.93 s
    c. 2.38 s
    d. 6.87 s
    e. 13.1 s

13. The effect of adding a catalyst to a reaction is to

    a. increase the number of collisions between reactants
    b. increase the energy of the products
    c. increase the equilibrium constant of a reaction
    d. lower the activation energy of a reaction
    e. decrease the enthalpy change of a reaction

14. Which of the following elements is most likely to form compounds with an expanded valence shell?

    a. P
    b. Ne
    c. F
    d. Li
    e. N

15. Which of the following combinations is most likely to produce an ionic bond?

    a. Cl and Br
    b. P and S
    c. N and O
    d. B and O
    e. Li and F

16. Which of the following aqueous solutions would have the highest vapor pressure at 25 °C?

    a. pure water
    b. 1 m glucose ($C_6H_{12}O_6$)
    c. 1 m $NaNO_3$
    d. 1 m $MgCl_2$
    e. 1 M $(NH_4)_2SO_4$

17. What is the formal charge on each atom in a hypobromite ion, OBr^{-1}?

    a. O = -2, Br = -1
    b. O = -2, Br = +1
    c. O = -1, Br = +1
    d. O = -1, Br = 0
    e. O = 0, Br = -1

18. Use VSEPR theory to predict the molecular geometry of HCN.

    a. bent
    b. linear
    c. trigonal planar
    d. tetrahedral
    e. octahedral

19. How many sigma (σ) bonds and pi (π) bonds are in acetylene, C_2H_2?

    a. one σ, one π
    b. two σ, two π
    c. three σ, one π
    d. three σ, two π
    e. four σ, one π

20. One product of the combustion of ethylene, C_2H_4, is carbon dioxide. What change in hybridization of the carbon occurs in this reaction?

    a. sp^3 to sp^2
    b. sp^3 to sp
    c. sp^3 to sp^3d
    d. sp^2 to sp^3d^2
    e. sp^2 to sp

21. Use molecular orbital theory to predict which ion is paramagnetic.

    a. C_2^{2-}
    b. O_2^{2-}
    c. O_2^{2+}
    d. N_2^{2-}
    e. B_2^{2-}

22. Which of the following molecules may be a cycloalkane?

    a. C_3H_8
    b. C_4H_6
    c. C_5H_{12}
    d. C_6H_{12}
    e. C_7H_{16}

23. The empirical formula of a certain hydrocarbon is CH_2. When 0.120 mole of the hydrocarbon is completely combusted with excess oxygen, 17.7 L CO_2 gas is produced at 27 °C and 1.00 atm. What is the molecular formula of the hydrocarbon?

    a. C_2H_2
    b. C_2H_4
    c. C_3H_6
    d. C_5H_{10}
    e. C_6H_{12}

24. An unknown gaseous hydrocarbon contains 85.63 % C. Its density is 0.426 g/L at 0.465 atm and 373 K. What is the molecular formula of the gas?

    a. C_2H_4
    b. C_3H_6
    c. C_4H_8
    d. C_5H_{10}
    e. C_6H_{12}

25. What intermolecular force or bond is primarily responsible for the solubility of H_2S in water?

    a. ion-dipole force
    b. dipole-dipole force
    c. ionic bonding
    d. covalent bonding
    e. hydrogen bonding

26. What is the solute mole fraction of 1.98 m $Fe(NO_3)_3(aq)$? The molar mass of $Fe(NO_3)_3$ is 241.9 g/mol and the molar mass of water is 18.02 g/mol.

    a. 0.0345
    b. 0.0641
    c. 0.324
    d. 0.479
    e. 0.863

27. Concentrated hydrofluoric acid is 28.9 M and has a density of 1.18 g/mL. What is the weight percent of concentrated HF?

    a. 24.5%
    b. 49.0%
    c. 51.0%
    d. 68.2%
    e. 75.5%

28. The Henry's law constant for N_2 in water at 37 °C is 8.2×10^{-7} M/mm Hg. What is the equilibrium concentration of N_2 in water when the partial pressure of N_2 is 634 mm Hg?

    a. 1.3×10^{-9} M
    b. 5.2×10^{-4} M
    c. 1.9×10^{-2} M
    d. 1.9×10^{3} M
    e. 7.7×10^{8} M

29. Which of the following species will have a Lewis structure most like that of a sulfate ion, SO_4^{2-}? Assume that the Lewis structure has no double bonds.
    a. NH_3
    b. CBr_4
    c. SO_3
    d. H_2CO
    e. H_2O

30. Use Lewis structures to predict the bond order for a nitrogen-oxygen bond in the nitrite ion, NO_2^{1-}.

    a. 1/2
    b. 1
    c. 4/3
    d. 3/2
    e. 3

31. For $NH_4NO_3(aq)$, the solvent is

    a. NH_4NO_3
    b. NH_4^+
    c. NO_3^{1-}
    d. water
    e. Duff beer

32. Determine which of the following species is paramagnetic.

    a. NO^{+1}
    b. CO
    c. CN^{-1}
    d. OF^{-1}
    e. NO

CH 222 Sample Final Exam
Name: ____ Lab Section: ______
Answers

<u>Part I</u>: Multiple Choice Questions

1. A
2. A
3. C
4. E
5. B

6. A
7. E
8. C
9. E

10. A
11. D
12. C
13. D
14. A
15. E

16. A
17. D
18. B
19. D
20. E
21. D
22. D

23. E
24. A
25. B
26. A
27. B
28. B
29. B

30. D
31. D
32. E