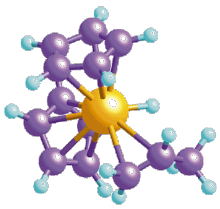
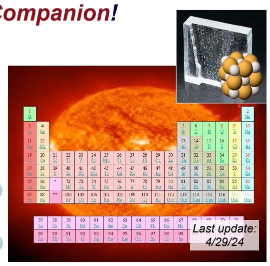
Welcome to Chemistry 221!

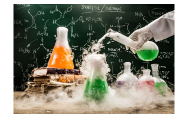
http://mhchem.org/221

Get the CH 221 Companion!





CH 221: Lectures and Labs



Lectures: **MWF** from **9 - 9:50 AM** in **AC 1303** *(this room)*
- Lectures recorded, available soon afterwards
- Lecture notes to print available (under "Problem Sets and Handouts", mhchem.org/221) and in ***Chemistry 221 Companion*** (get it!)

Labs (Section 01): **Mondays from 1:10 - 5 PM**
- **Start** in **room AC 2501**
- Move to **AC 2507** ("the lab") around **3 PM**
- *For first day*, **bring a printed copy of the "Eight Bottles" Lab** (mhchem.org/221) and your **calculator**
- Some labs will require safety glasses (Dollar store ok)

MAR *...more on Monday afternoon*

The Nature of Chemistry

What is Chemistry?
- "Keme" (earth)
- "Kehmeia" (transmutation)
- "Al-Khemia" (Arabic)
- "alchemy" (Europe's Dark Age)
- "chymistry" (Boyle's 1661 publication)
- "chemistry" (modern)

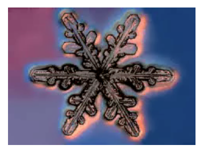
Khemeia - ancient Egyptian processes for embalming the dead, later extended to metallurgy

Khemeia (and later chemistry) seen as "occult" by laymen, extended to modern age

What is Matter?

M.

The Nature of Chemistry



How does Matter Change?

How does Matter Interact?

Why Study Chemistry?

MAR

The Art (?) of Chemistry

Chemistry and Art?!?

Dr. Roald Hoffman,
1981 Nobel Prize in Chemistry

Stick to the chemistry, Roald!

"There was no question that
the reaction worked
but transient colors were seen
in the slurry of sodium methoxide in
dichloromethane
and we got a whole lot of products
for which we can't sort out the
kinetics
the next slide show will show
the most important part
very rapidly
within two minutes
and I forgot to say on further
warming
we get in fact the ketone…"

MAR

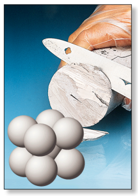
- **Organic - carbon, nitrogen, oxygen**
- **Inorganic - metals, everything "non-carbon"**
- **Analytical - Spectroscopy, "how much", "what kind"**
- **Physical - measurement, where physics meets chemistry**
- **Biochemical - the chemistry of life**
- ***many others!***

The Branches of Chemistry

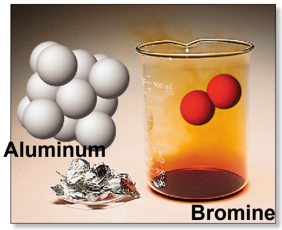
MAR

The Language of Chemistry

CHEMICAL ELEMENTS - pure substances that cannot be decomposed by ordinary means to other substances.



Sodium



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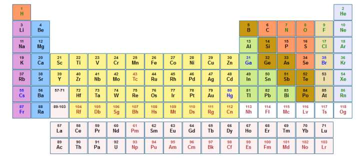
The Language of Chemistry

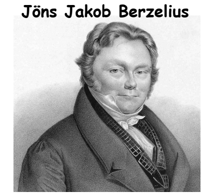
The elements, their names, and symbols are given on the PERIODIC TABLE

***Berzelius* - first to use letter symbols for atoms**

How many elements are there?



Jöns Jakob Berzelius



pre-Biblical elements: Au (sun), Ag (moon), Cu (Venus), Fe (Mars), Sn (Jupiter), Pb (Saturn), Hg (Mercury), S, C

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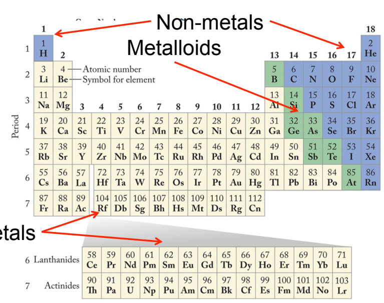
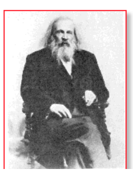
The Periodic Table

Periodic table originally organized by mass, now by atomic number



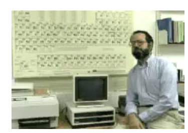


Dmitri Mendeleev (1834 - 1907)

Predicted Ga, Ge, Sc and Tc!

Dr. Frank DiSalvo (Cornell University)

"On the importance of the periodic table"



Number of compounds possible is virtually limitless!!!

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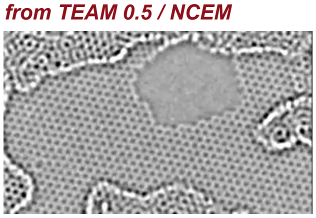
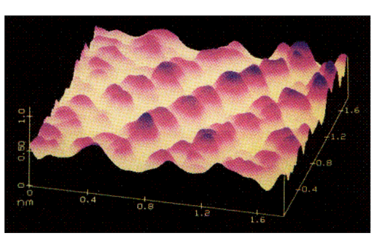
An atom is the smallest particle of an element that has the chemical properties of the element.

Real time carbon atoms from TEAM 0.5 / NCEM





Copper atoms on a silica surface
Distance across = 1.8 nanometer (1.8 x 10^{-9} m)

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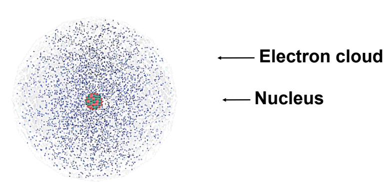
The Atom

An atom consists of a nucleus (of protons and neutrons) and electrons in space about the nucleus.



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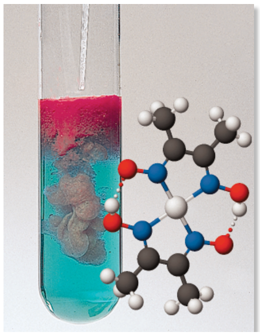


CHEMICAL COMPOUNDS are composed of atoms and so can be decomposed to those atoms.

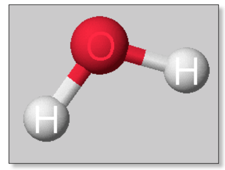
The red compound is composed of
- nickel (Ni) (silver)
- carbon (C) (black)
- hydrogen (H) (white)
- oxygen (O) (red)
- nitrogen (N) (blue)

This type of compound is an ionic compound - unshared electrons

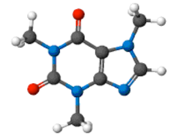
A **MOLECULE** is the smallest unit of a compound that retains the chemical characteristics of the compound.

Composition of molecules is given by a **MOLECULAR FORMULA**

H_2O



$C_8H_{10}N_4O_2$ - caffeine



Water and caffeine are examples of covalent compound - shared electrons

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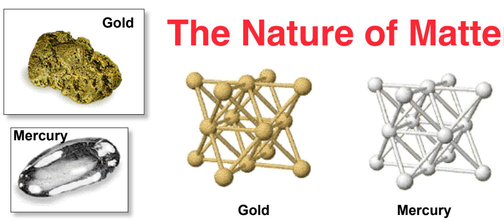
The Nature of Matter



Chemists are interested in the nature of matter and how this is related to its atoms and molecules.

We can explore the MACROSCOPIC world - what we can see - to understand the ATOMIC world - what we cannot see - using SYMBOLS.

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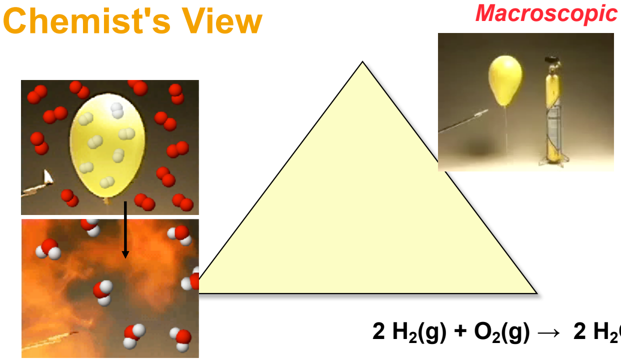
A Chemist's View

Macroscopic



$2 H_2(g) + O_2(g) \rightarrow 2 H_2O(g)$

Atomic *Symbolic*

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STATES OF MATTER

SOLIDS: rigid shape, fixed volume, reasonably well understood.

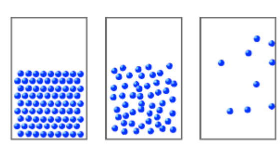
LIQUIDS: no fixed shape, may not fill a container completely, not well understood.

GASES: expand to fill their container, good theoretical understanding.

also PLASMA - more in CH 222!

KINETIC NATURE OF MATTER

Matter consists of atoms and molecules in motion.



Kinetic Molecular Theory describes solids, liquids and gases

~~Test Monkeys?~~ *Er, sorry,* Student volunteers?!?

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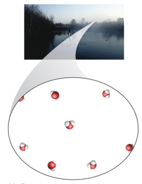
STATES OF MATTER & ENERGY

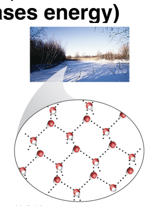
Energy of transitions *important to scientists:*

solid -> liquid -> gas: **endothermic** *(takes energy)*

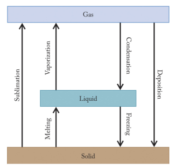
gas -> liquid -> solid: **exothermic** **(releases energy)**

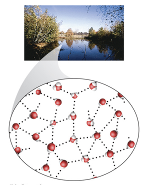


(c) Gas



(a) Solid





(b) Liquid

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Physical Properties

Physical properties can be observed and measured without changing the composition of a substance. They include:

- **color**
- **melting and boiling point**
- **odor**

Some **physical changes** would be

- boiling of a liquid
- melting of a solid
- dissolving a solid in a liquid to give a homogeneous mixture - a **SOLUTION**.

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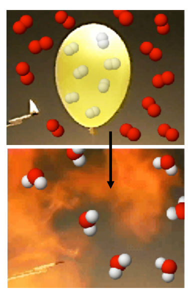
Chemical Properties and Chemical Change

Burning hydrogen (H_2) in oxygen (O_2) gives H_2O.

Chemical change or chemical reaction involves the transformation of one or more atoms or molecules into one or more different molecules.



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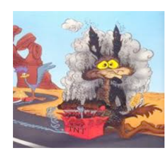
Chemical Properties and Physical Properties

Physical properties do not change the composition of the substance

Chemical properties change the composition of the substance



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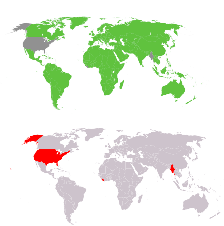
Physical Properties

Physical properties useful in separating compounds and elements

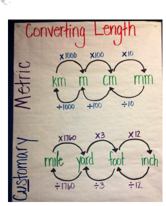
- **density**
- **melting and boiling point**
- **magnetism**



Physical and chemical properties require units - need

METRIC SYSTEM!

See the Metric Guide



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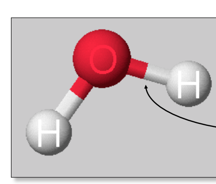
Units of Length / Conversions

1 kilometer (km)	=	**10^3 meters (m)**
1 centimeter (cm)	=	**10^{-2} meters (m)**
1 millimeter (mm)	=	**10^{-3} meters (m)**
1 micrometers (μm)	=	**10^{-6} meters (m)**
1 nanometer (nm)	=	**10^{-9} meters (m)**

Know these five metric conversions!



O-H distance =
9.4×10^{-11} m
9.4×10^{-9} cm
9.4×10^{-5} μm
0.094 nm

UNITS OF MEASUREMENT

We make QUALITATIVE observations of reactions - changes in color and physical state.

We also make QUANTITATIVE MEASUREMENTS, which involve numbers and amount.

Use SI units - based on the metric system

length	(meter, m)
mass	(kilogram, kg)
time	(second, s)

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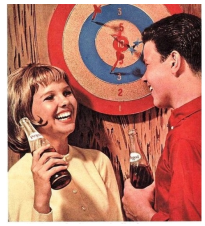
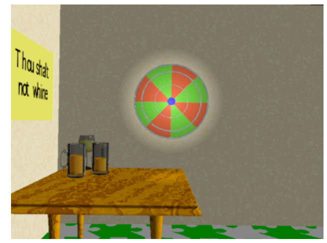
Accuracy and Precision

Measurements affected by accuracy and precision.



Don't stand next to the dart board, especially with poor precision!



Accurate *and* Precise

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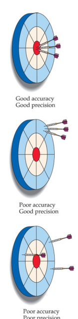
Accuracy versus Precision

Accuracy refers to the proximity of a measurement to the true value of a quantity

Accuracy determined by **% error**

Precision refers to the proximity (reproducibility) of several measurements to each other.

Determined by average deviation or **parts per thousand**



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Experimental Error

Average deviation:

Step 1: find the absolute value of the difference between each measurement and the average.

Step 2: find the summation of all the deviations and divide by the total number of measurements.

Standard deviation (not used in CH221):

$$\text{Standard deviation} = \sqrt{\frac{\text{sum of squares of deviations}}{(\text{\# of deviations} - 1)}}$$

ppt (parts per thousand):

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

Percent error:

$$\%\ \text{error} = \frac{\text{experimental value - accepted value}}{\text{accepted value}} \times 100$$

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Experimental Error - *Example*

Trial #	Boiling Point (°C)	Average (°C)	Deviations (°C)	Ave. Dev. (°C)
1	11.23	11.19	0.04	0.06
2	11.09		0.10	
3	11.27		0.08	
4	11.16		0.03	

Average Deviation = 0.06 °C (11.19 ± 0.06 °C)

ppt = (0.06 °C / 11.19 °C) x 1000 = 5 ppt

If the literature (accepted) value was 11.25 °C,

% error = (11.19 °C – 11.25 °C / 11.25 °C) x 100 = -0.5%

sometimes %error is absolute value (always positive)

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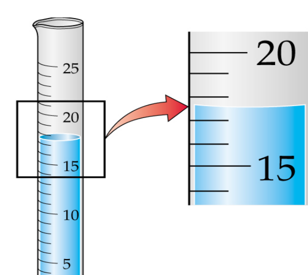
Measurement and Significant Figures

Every experimental measurement, *no matter how precise*, has a degree of uncertainty because there is a limit to the number of digits that can be determined.

Need mathematical system - SIGNIFICANT FIGURES - *very important*, see Chapter One in text and Handout



? mL of water

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Measurement and Significant Figures

To indicate the precision, recorded values should use all the digits known with certainty *plus* one additional estimated digit

Estimated ("doubtful") digit usually considered *uncertain* by plus or minus 1 (± 1)

The total number of digits used to express such a measurement is called the **number of *significant figures (sig figs)***.

Ex: **65.07 g - four sig figs, 7 "doubtful"**

Ex: **54.70318 g - seven sig figs, 8 "doubtful"**

MAR

Rules for Determining Significant Figures

1. **Zeroes in the middle of a number are significant. 69.08 has four sig figs.**
2. **Zeroes at the beginning of a number are not significant. 0.0089 has two sig figs (8 and 9).**
3. **Zeroes at the end of a number and after the decimal point are significant. 2.50 has three sig figs. 25.00 has four sig figs.**
4. **Zeroes at the end of a number and before the decimal point will be significant *only* with a decimal placeholder (period). 1500 has two sig figs, but 1500. has four sig figs.**
5. **Exact conversions (Definitions) have *infinite* sig figs (*ex:* 60 s/1 min, 10 mm/1 cm).**
6. **STUDY! PRACTICE! IMPORTANT!**

MAR

Scientific Notation

Know how your calculator displays scientific notation (and also "regular" notation!)

Always use proper scientific notation when reporting answers in lab, quizzes, etc.

Scientific Notation is a convenient way to write very small or large numbers

Written as a product of a number between 1 and 10, times the number 10 raised to a power. *Examples:*

$$215. = 2.15 \times 10^2$$

Decimal point is moved two places to the left, so exponent is 2.

$$1.56 \times 10^{-8} = 0.000\,000\,015\,6$$

Negative exponent of –8,
so decimal point is moved to the left eight places.

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See: Scientific Notation Handout & Scientific Notation Handout #2

Calculators, Rounding and Sig Figs

Calculators produce large numbers in calculations but the reportable sig figs is usually *much less*! The calculator's large number must be **rounded off** to a smaller number *keeping only significant figures*.

Once you decide how many numbers to keep *(next slide)*, look at the **first** digit to be dropped:

If the first digit you remove is <u>between 0 and 4</u>, drop it and all remaining digits.

If the first digit you remove is <u>between 5 and 9</u>, round the number **up** by adding 1 to the digit to the left of the one you drop

Example: 2.4271 becomes 2.4 when rounded to two significant figures

Example: 4.5816 becomes 4.6 when rounded to two significant figures

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Rules for Rounding off Numbers

For multiplication and division: The answer cannot have more significant figures than either of the original numbers.

Three significant figures

Three significant figures

$$\frac{278\text{ mi}}{11.70\text{ gal}} = 23.8\text{ mi/gal}$$

Four significant figures

Actual value: 23.76068…..

MAR

Rules for Rounding off Numbers

For addition and subtraction: The final number must stop at the ***largest doubtful digit***.

Volume of water at start	3.18? ?? L
Volume of water added	+ 0.013 15 L
Total volume of water	3.19? ?? L

3.18: 8 is the **"*doubtful digit*"**, it stops at the **hundredths** spot

0.01315: 5 is the **"*doubtful digit*"**, it stops at the **hundredth thousandths** spot

Final answer is 3.19 L:
Answer stops at largest "doubtful digit" *(hundredths vs. hundredth thousandths)*

Actual value: 3.19315

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Density: *the* **ratio** *of a substance's* **mass** *(grams)* **to its volume** *(mL, cm³)*

Substances layer themselves according to their density: **superposition**

Density used to separate materials

Gold

d = 19.3 g/cm³

Mercury

d = 13.6 g/cm³

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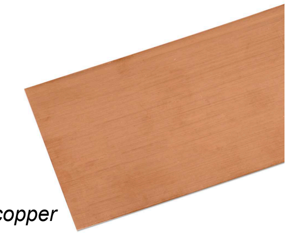
Density Problem

Problem: A piece of copper has a mass of 57.54 g. It is 9.36 cm long, 7.23 cm wide, and 0.95 mm thick. Calculate density (g/cm³).

$$\text{Density} = \frac{\text{mass (g)}}{\text{volume (cm}^3)}$$



Sheet of copper

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Density Problem

Problem: A piece of copper has a mass of 57.54 g. It is 9.36 cm long, 7.23 cm wide, and 0.95 mm thick. Calculate density (g/cm³).

SOLUTION

1. **Get dimensions in common units.**

$$0.95 \text{ mm} \cdot \frac{1 \text{ cm}}{10 \text{ mm}} = 0.095 \text{ cm}$$

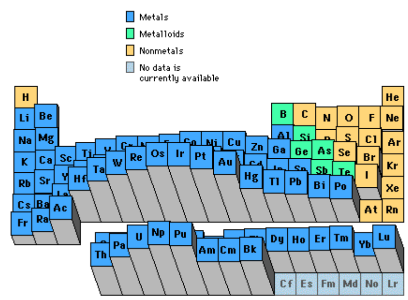
2. **Calculate volume in cubic centimeters.**

$$(9.36 \text{ cm})(7.23 \text{ cm})(0.095 \text{ cm}) = \mathbf{6.4\ cm^3}$$

6.42891...

3. **Calculate the density.**

Relative Densities of the Elements



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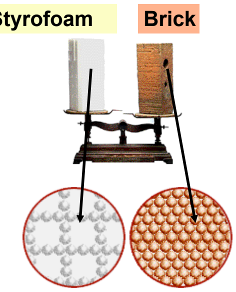
DENSITY

Notice that density is an INTENSIVE property of matter

INTENSIVE - does not depend on quantity of matter (density, boiling point, etc.)

Contrast with EXTENSIVE - depends on quantity of matter. Examples include mass and volume

Styrofoam **Brick**



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PROBLEM: Mercury (Hg) has a density of 13.6 g/cm³. What is the mass of 95 mL of Hg? (454 g = 1 lb)

Solve the problem using DIMENSIONAL ANALYSIS - *see the Dimensional Analysis and Factor Label handouts on the Web*

MAR

PROBLEM: Mercury (Hg) has a density of 13.6 g/cm³. What is the mass of 95 mL of Hg? (454 g = 1 lb)

First, note that **1 cm³ = 1 mL**

Then, use dimensional analysis to calculate mass.

$$95\ cm^3 \cdot \frac{13.6\ g}{1\ cm^3} = 1.3 \times 10^3\ g$$

1292

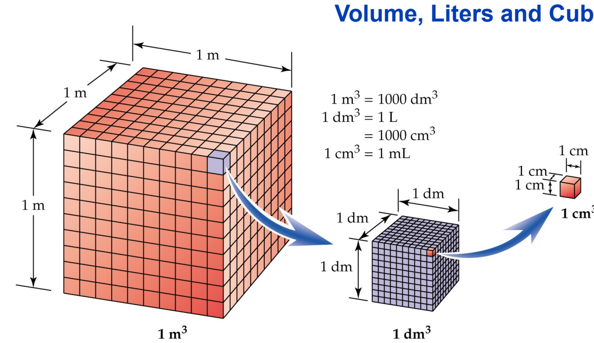
What is the mass in pounds?

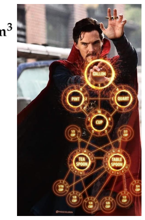
$$1.3 \times 10^3\ g \cdot \frac{1\ pound}{454\ g} = 2.9\ lb$$

2.863436...

MAR

Volume, Liters and Cubed Lengths





cm³ = cm * cm * cm, *etc.*

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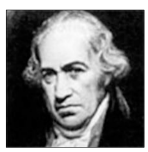
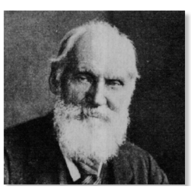
Temperature Scales

- **Fahrenheit (°F)**
- **Celsius (°C)**
- **Kelvin (K)**

Anders Celsius
1701-1744

Daniel Fahrenheit
1686-1736





Lord Kelvin
(William Thomson)
1824-1907

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Temperature Scales

	Fahrenheit	Celsius	Kelvin
Boiling Point of water	212	100	373.15
	180°	100°	100°
Freezing Point of water	32	0	273.15

Notice that 1 Kelvin degree = 1 degree Celsius
***Difference** between Celsius temperatures and Kelvin temperatures the same!*

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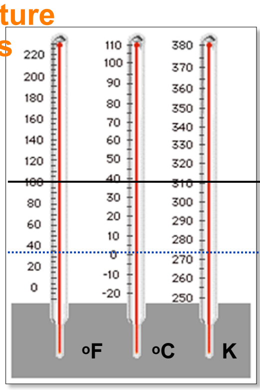
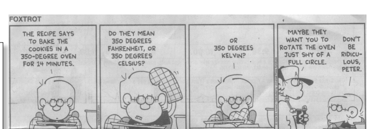
Temperature Scales





100 ˚F = 38 ˚C = 311 K

32 ˚F = 0 ˚C = 273 K

Note:
no negative K!
no degree K symbol!

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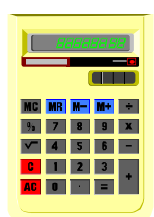

Calculations Using Temperature

Chemistry experiments recorded in Celcius (°C) but calculations *generally* require temperatures in Kelvins (K)

$$T\ (K) = T\ (˚C) + 273.15$$

Body temp = 37.0 ˚C + 273.15 = 310.2 K

310.15

Liquid nitrogen = 77 K - 273.15 = -196 ˚C

-196.15

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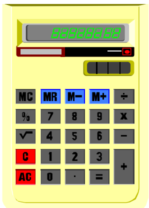
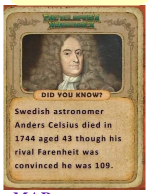
Memorize 273.15!

Calculations Using Temperature





Occasionally need Fahrenheit (F) values

Convert using Celsius scale

T (˚F) = $^9/_5$ T (˚C) + 32.00

Liquid He = 4.2 K - 273.15 = -269.0 ˚C *-268.95*

T (˚F) = $^9/_5$ (-269.0 ˚C) + 32.00 = -452.2 ˚F *-452.2*

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Mass Percentages in Chemistry

Often see “30% lead, 70% oxygen”

This means that in 100 grams of the substance

30 grams will be lead

70 grams will be oxygen

In one gram of the substance,

0.30 grams will be lead

0.70 grams will be oxygen

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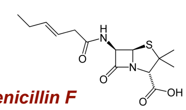
Mass Percentages in Chemistry

Example: **Penicillin F is 53.829% carbon. How much carbon in 75 g of Penicillin F?**

Solution

75 g Penicillin F* (53.829 g C / 100 g Penicillin F) =
40. g carbon *(40.37175)*

***Note* that volume percentages occasionally seen, but not often in our class**



Penicillin F

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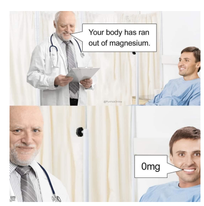
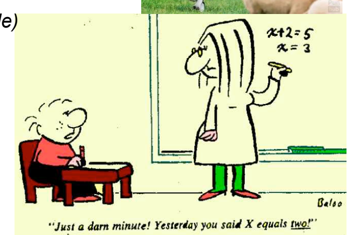
End of Chapter One

See also:

- *Chapter One Study Guide*
- *Chapter One Concept Guide*
- *Math ("Chapter Guide Zero") Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*





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Important Equations, Constants, and Handouts from this Chapter:

metric prefixes:
nano (n) = 10^{-9}
micro (μ) = 10^{-6}
milli (m) = 10^{-9}
centi (c) = 10^{-9}
kilo (k) = 10^{-9}

$$\text{Density} = \frac{\text{mass (g)}}{\text{volume (cm3)}}$$

T (K) = T (˚C) + 273.15

1 cm^3 = 1 mL

significant figures!!!

mass percentages

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End of Chapter Problems: *Test Yourself*

1. 32.32 - 23.2 = _____
2. 32.4 * 37.31 = _____
3. 4.311 / 0.07 = _____
4. Convert 37.0 C to K.
5. Convert 253.6 mL to cm^3.
6. Convert 24 m^3 to cm^3.
7. 235.05 + 19.6 + 2 = _____
8. 58.925 - 19 = _____
9. 2.19 x 4.2 = _____
10. 4.311 ÷ 0.07 = _____
11. The platinum-containing cancer drug cisplatin contains 65.0% platinum. If you have 1.53 g of the compound, what mass of platinum (in grams) is contained in this sample?
12. The anesthetic procaine hydrochloride is often used to deaden pain during dental surgery. The compound is packaged as a 10.% solution (by mass; d = 1.0 g/mL) in water. If your dentist injects 0.50 mL of the solution, what mass of procaine hydrochloride (in milligrams) is injected?

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End of Chapter Problems: *Answers*

1. 9.1
2. 1210
3. 60
4. 310.2 K
5. 253.6 cm^3
6. 2.4 x 10^7 cm^3
7. 257
8. 40.
9. 9.2
10. 60
11. 0.995 g Pt
12. 50. mg

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