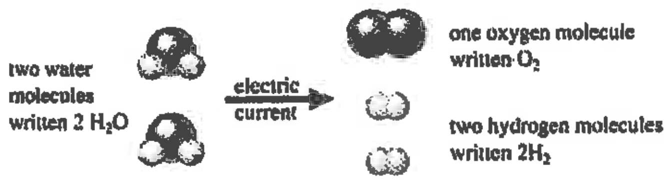
NATIONAL
MATH + SCIENCE
INITIATIVE

1

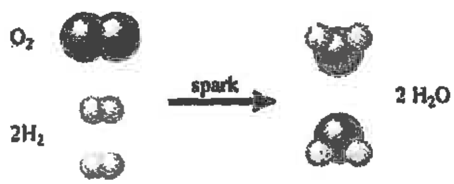
AP* Chemistry
Chemical Foundations

Chemistry: An Overview

- **Matter** – takes up space, has mass, exhibits inertia
  - composed of atoms only 100 or so different types
  - Water made up of one oxygen and two hydrogen atoms
  - Pass an electric current through it to separate the two types of atoms and they rearrange to become two different types of molecules



- reactions are reversible

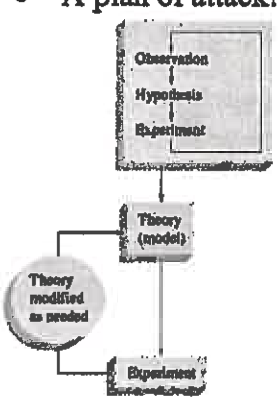


Chemistry – is defined as the study of matter and energy and more importantly, the changes between them

- **Why study chemistry?**
  - become a better problem solver in all areas of your life
  - safety – had the Roman's understood lead poisoning, their civilization would not have fallen
  - to better understand all areas of science

The Scientific Method

- A plan of attack!



The fundamental steps of the scientific method

Steps in the Scientific Method

1. *Making observations.* Observations may be *qualitative* (the sky is blue; water is a liquid) or *quantitative* (water boils at 100°C; a certain chemistry book weighs 2 kilograms). A qualitative observation does not involve a number. A quantitative observation (called a measurement) involves both a number and a unit.
2. *Formulating hypotheses.* A hypothesis is a *possible* explanation for an observation.
3. *Performing experiments.* An experiment is carried out to test a hypothesis. This involves gathering new information that enables a scientist to decide whether or not the hypothesis is valid—that is, whether it is supported by the new information learned from the experiment. Experiments always produce new observations, and this brings the process back to the beginning again.

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Units of Measure

A quantitative observation, or measurement, ALWAYS consists of two parts: a *number* and a *unit*.

Two major measurements systems exist: English (US and some of Africa) and Metric (the rest of the globe!)

- SI system – 1960 an international agreement was reached to set up a system of units so scientists everywhere could better communicate measurements. Le Système International in French; all based upon or derived from the metric system

Table 1.1 Fundamental SI Units

Physical Quantity	*Name of Unit*	*Abbreviation*
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

Table 1.3 Some Examples of Commonly Used Units

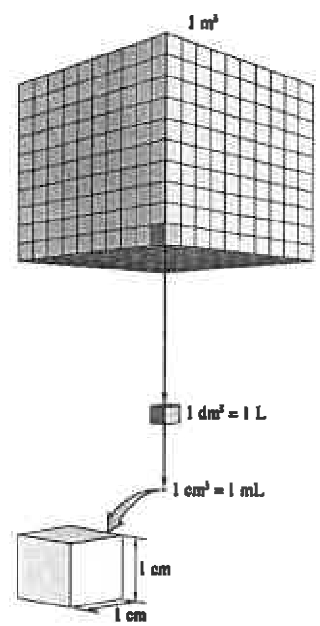
Length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
Mass	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
Volume	A 12-oz can of soda has a volume of about 360 mL.

Table 1.2 The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)

Prefix	*Symbol*	*Meaning*	*Exponential Notation**
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

*See Appendix 1.1 if you need a review of exponential notation.

KNOW THE UNITS AND PREFIXES shown in **BLUE**!!!



- Volume – derived from length; consider a cube 1m on each edge $\therefore$ 1.0 m^3
  - A decimeter is 1/10 of a meter so
    $(1m)^3 = (10\ dm)^3 = 10^3\ dm^3 = 1{,}000\ dm^3$
    $1dm^3$ = 1 liter (L) and is slightly larger than a quart also
    $1dm^3 = 1\ L = (10\ cm)^3 = 10^3 cm^3 = 1{,}000\ cm^3 = 1{,}000\ mL$
    AND $1\ cm^3 = 1\ mL = 1$ gram of H_2O (at 4°C if you want to be picky)

Mass vs. Weight – chemists are quite guilty of using these terms interchangeably.

- mass (g or kg) – a measure of the resistance of an object to a change in its state of motion (i.e. exhibits inertia); the quantity of matter present
- weight (a force $\therefore$ has units of Newtons) – the response of mass to gravity; since all of our measurements will be made here on Earth, we consider the acceleration due to gravity a constant so we'll use the terms interchangeably as well *although* it is technically incorrect! We "weigh" chemical quantities on a **balance NOT a scale!!**

Significant Figures and Calculations

Determining the Number of Significant Figures (or Digits) in a Measurement

- Nonzero digits are significant. (Easy enough to identify!)
- A zero is significant IF and ONLY IF it meets one of the conditions below:
  - The zero in question is "terminating AND right" of the decimal [must be both]
  - The zero in question is "sandwiched" between two significant figures
- Exact or counting numbers have an ∞ amount of significant figures as do fundamental constants (never to be confused with derived constants)

Exercise 2 **Significant Figures (SF)**
Give the number of significant figures for each of the following experimental results.

a. A student's extraction procedure on a sample of tea yields 0.0105 g of caffeine.

b. A chemist records a mass of 0.050080 g in an analysis.

c. In an experiment, a span of time is determined to be 8.050×10^{-3} s.

a. three; b. five; c. four

Reporting the Result of a Calculation to the Proper Number of Significant Figures

- When × and ÷ , the term with the **least** number of *significant figures* (∴ least accurate measurement) determines the number of **maximum** number of significant figures in the answer. (It's helpful to underline the digits in the least significant number as a reminder.)

$$4.56 \times \underline{1.4} = 6.38 \xrightarrow{\text{corrected}} \underline{6.4}$$

- When + and (−), the term with the least number of *decimal places* (∴ least accurate measurement) determines the number of significant figures in the final answer.

$$\begin{array}{ll} 12.11 & \\ 18.0 & \leftarrow \text{limiting term (only 1 decimal place)} \\ \underline{1.013} & \\ 31.123 \xrightarrow{\text{corrected}} 31.1 & \text{(limits the overall answer to only one decimal place)} \end{array}$$

- pH – the *number of significant figures in least accurate measurement* determines *number decimal places* on the reported pH (usually explained in the appendix of your text)

Rounding Guidelines for the AP Exam and This Course:

- Round ONLY at the end of all calculations (keep the numbers in your calculator)
- Examine the significant figure one place beyond your desired number of significant figures.
  IF > 5 round up; < 5 drop the remaining digits.
- Don't "double round"! Example: The number 7.348 rounded to 2 SF is reported as 7.3
  In other words, DO NOT look beyond the 4 after the decimal and think that the 8 rounds the 4 up to a five which in turn makes the final answer 7.4.
  [Even though you may have conned a teacher into rounding your final average this way before!]

Exercise 6
The speed limit on many highways in the United States is 55 mi/h. What number would be posted if expressed in kilometers per hour?

88 km/h

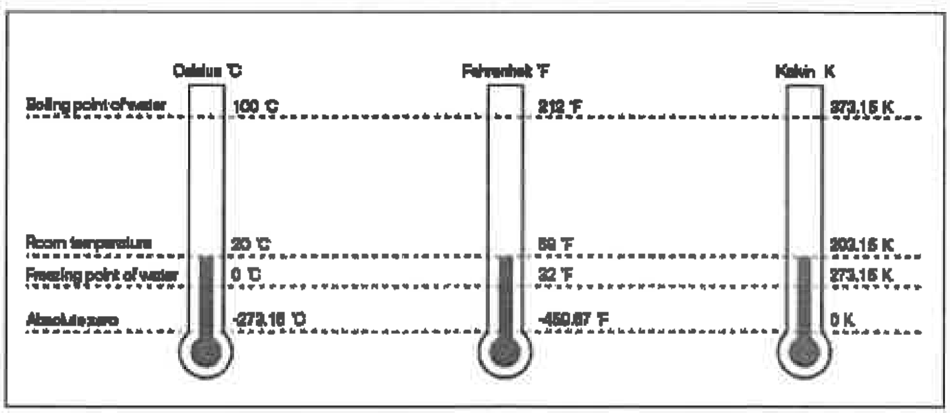
Exercise 7
A Japanese car is advertised as having a fuel economy of 15 km/L. Convert this rating to miles per gallon.

35 mi /gal

Temperature

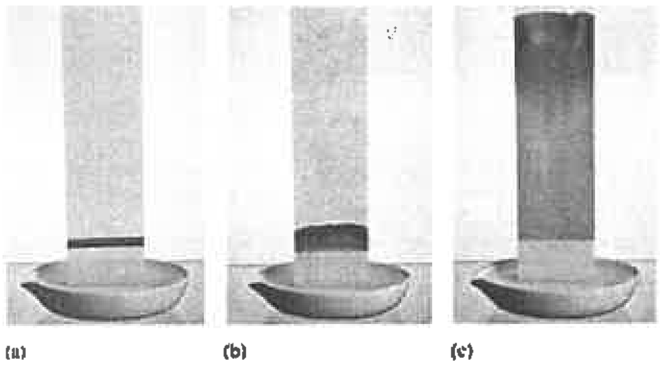
I suspect you are aware there are **three** temperature scales commonly in use today. A comparison follows:

Known Temperature				Required Temperature	Formulas
Celsius	°C	to	°F	Fahrenheit	°F = (1.8 X °C) + 32
Celsius	°C	to	K	Falvin	K = °C + 273.15
Fahrenheit	°F	to	°C	Celsius	°C = (°F - 32)/1.8
Fahrenheit	°F	to	K	Kelvin	K = °F + 459.67)/1.8
Kelvin	°K	to	°C	Celsius	°C = K-273.15
Kelvin	°K	to	°F	Fahrenheit	°F = (1.8 X K) - 459.67



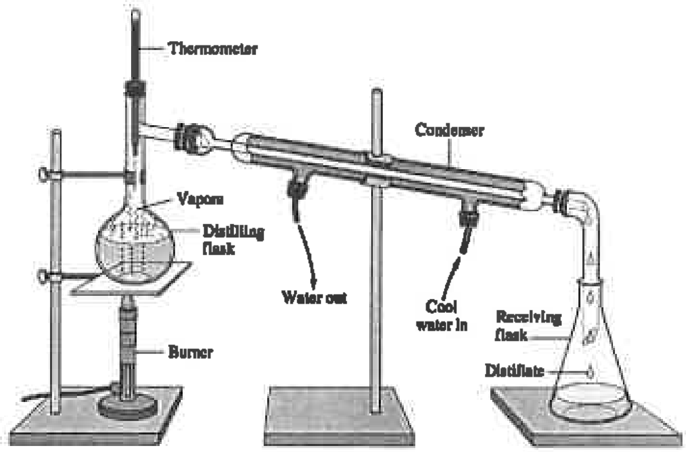
Notice a degree of temperature change on the Celsius scale represents the same quantity of change on the Kelvin scale.

Paper Chromatography:

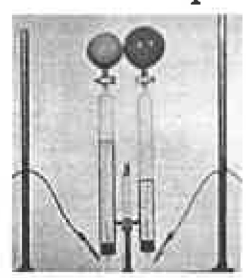


Paper chromatograph of ink. (a) A line of the mixture to be separate is placed at one end of a sheet of porous paper. (b) The paper acts as a wick to draw up the liquid. (c) The component with the weakest attraction for the paper travels faster than those that cling to the paper.

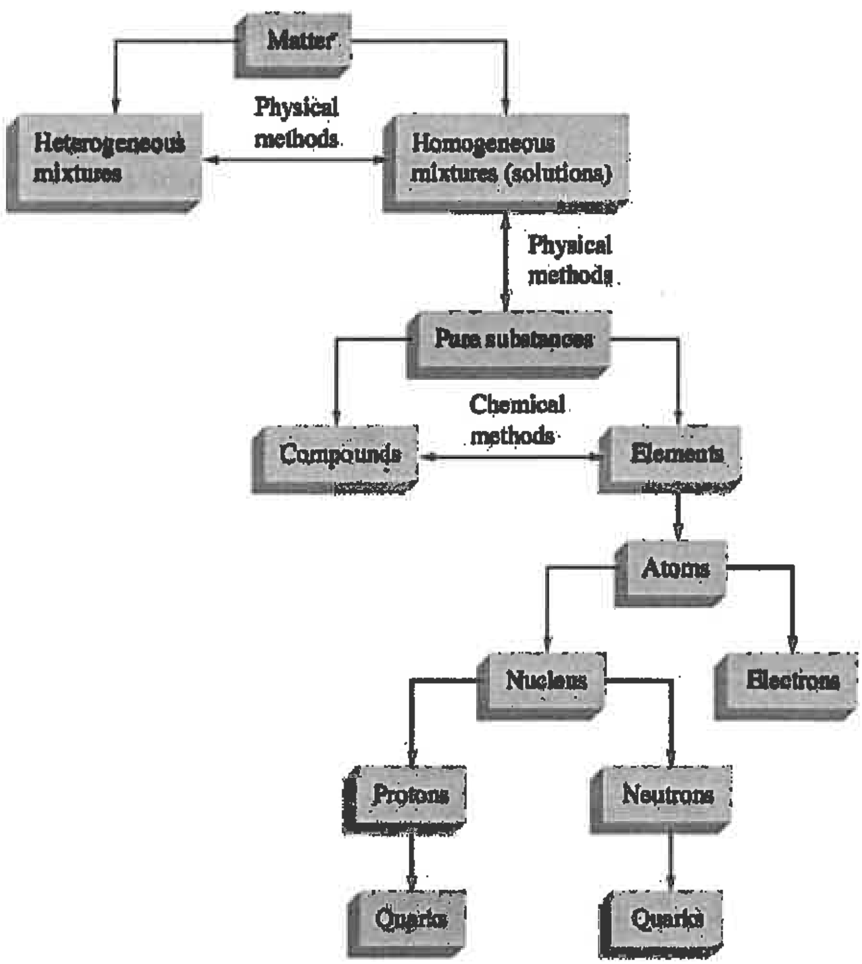
Distillation:



- **<u>Pure substances</u>** – compounds like water, carbon dioxide etc. <u>and</u> elements. Compounds can be separated into elements by **chemical** means
  - electrolysis is a common chemical method for separating compounds into elements
  - elements can be broken down into atoms which can be further broken down into
    - nuclei and electrons
    - p^+, n^0 and e^-
    - quarks



Electrolysis is an example of a chemical change. In this apparatus, water is decomposed to hydrogen gas (filling the red balloon) and Oxygen gas (filling the blue balloon).



NATIONAL MATH + SCIENCE INITIATIVE

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AP* Chemistry
ATOMS, MOLECULES & IONS

This is the highest honor given by the American Chemical Society. Priestly discovered oxygen. Ben Franklin got him interested in electricity and he observed graphite conducts an electric current. Politics forced him out of England and he died in the US in 1804. The back side, pictured below was given to Linus Pauling in 1984. Pauling was the only person to win Nobel Prizes in TWO Different fields: Chemistry and Peace.

THE EARLY HISTORY OF CHEMISTRY

- 1,000 B.C.—processing of ores to produce metals for weapons and ornaments; use of embalming fluids
- 400 B.C.—Greeks—proposed all matter was make up of 4 "elements": fire, earth, water and air
- Democritus—first to use the term *atomos* to describe the ultimate, smallest particles of matter
- Next 2,000 years—*alchemy*—a pseudoscience where people sought to turn metals into gold. Much was learned from the plethora of mistakes alchemists made.
- 16th century—Georg Bauer, German , refined the process of extracting metals from ores & Paracelsus, Swiss, used minerals for medicinal applications
- Robert Boyle, English—first "chemist" to perform **quantitative** experiments of pressure versus volume. Developed a working definition for "elements".
- 17th & 18th Centuries—Georg Stahl, German—suggested "phlogiston" flowed OUT of burning material. An object stopped burning in a closed container since the air was "saturated with phlogiston"
- Joseph Priestley, English—discovered oxygen which was originally called "dephlogisticated air"

FUNDAMENTAL CHEMICAL LAWS

- late 18th Century—Combustion studied extensively
- CO_2, N_2, H_2 and O_2 discovered
- list of elements continued to grow
- Antione Lavoisier, French—explained the true nature of combustion—published the first modern chemistry textbook AND stated the Law of Conservation of Mass. The French Revolution broke out the same year his text was published. He once collected taxes for the government and was executed with a guillotine as an enemy of the people in 1794. He was the first to insist on *quantitative* experimentation.

THE LAW OF CONSERVATION OF MASS:

Mass is neither created nor destroyed.

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DALTON'S ATOMIC THEORY

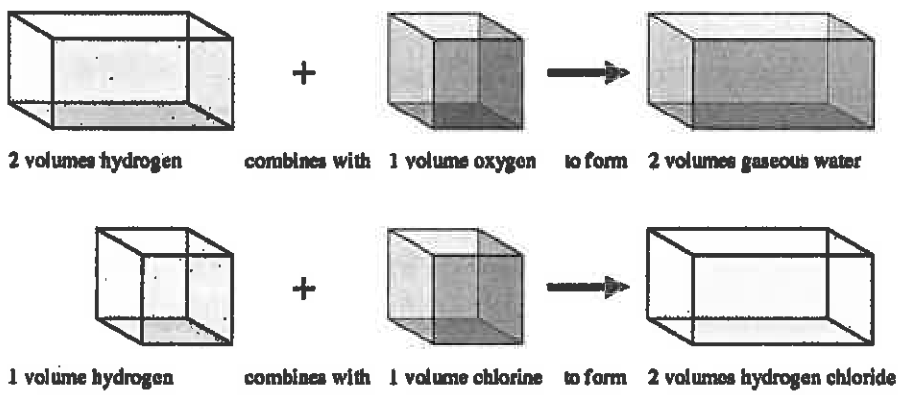
Postulates of Dalton's ATOMIC THEORY OF MATTER: (based on knowledge *at that time)*

1. All matter is made of **atoms**. These *indivisible and indestructible objects* are the ultimate chemical particles.
2. All the atoms of a given element are identical, in both weight and chemical properties. However, atoms of different elements have different weights and different chemical properties.
3. **Compounds** are formed by the combination of different atoms in the ratio of small whole numbers.
4. A **chemical reaction** involves only the combination, separation, or rearrangement of atoms; atoms are neither created nor destroyed in the course of ordinary chemical reactions.

**TWO MODIFICATIONS HAVE BEEN MADE TO DALTON'S THEORY:

1. *Subatomic particles were discovered. Bet you can name them!*
2. *Isotopes were discovered. Bet you can define "isotope" as well!*

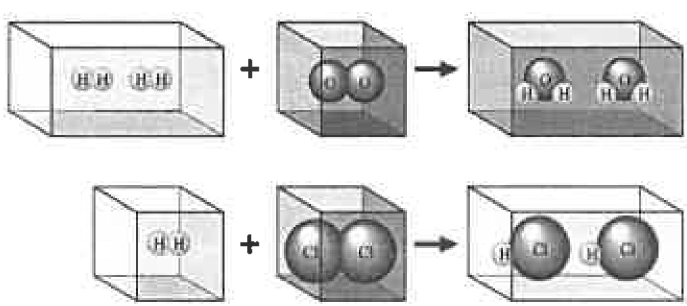
- 1809 Joseph Gay-Lussac, French—performed experiments [at constant temperature and pressure] and measured volumes of gases that reacted with each other.



- 1811 Avogadro, Italian—proposed his hypothesis regarding Gay-Lussac's work [and you thought he was just famous for 6.02×10^{23}] He was basically ignored, so 50 years of confusion followed.

AVOGADRO'S HYPOTHESIS:

At the same temperature and pressure, equal volumes of <u>different</u> gases contain the same number of particles.



RADIOACTIVITY

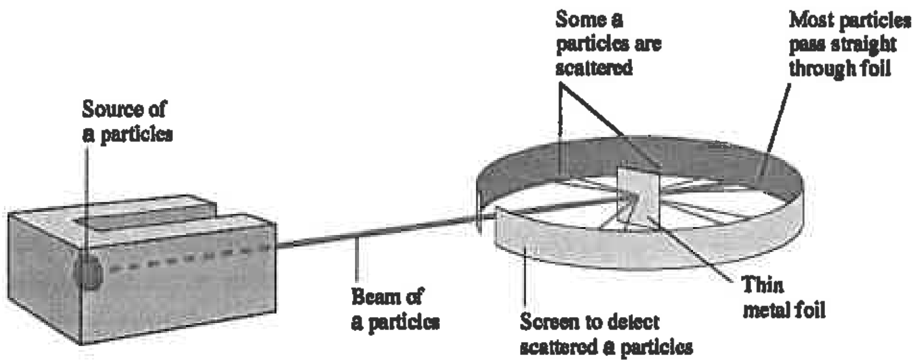
- Henri Becquerel, French—found out quite by accident [serendipity] that a piece of mineral containing uranium could produce its image on a photographic plate in the *absence* of light. He called this **radioactivity** and attributed it to a spontaneous emission of radiation by the uranium in the mineral sample.

- THREE types of radioactive emission:
  - **alpha, α**--equivalent to a helium nucleus; the largest particle radioactive particle emitted; 7300 times the mass of an electron. $^{4}_{2}He$ Since these are larger that the rest, early atomic studies often involved them.
  - **beta, β**--a high speed electron. $^{0}_{-1}\beta$ OR $^{0}_{-1}e$
  - **gamma, γ**--pure energy, no particles at all! Most penetrating, therefore, most dangerous.

THE NUCLEAR ATOM

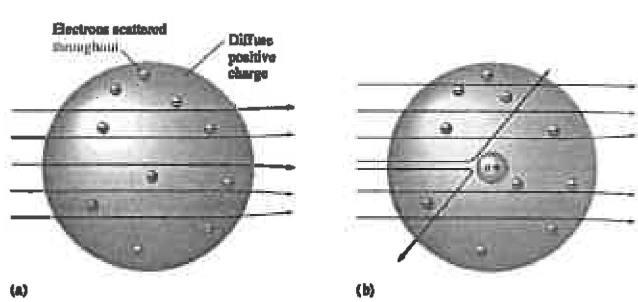
- 1911 Ernest Rutherford, England—A pioneer in radioactive studies, he carried out experiments to test Thomson's plum pudding model.
  - Directed α particles at a thin sheet of gold foil. He thought that if Thomson was correct, then the massive α particles would blast through the thin foil like "cannonballs through gauze". [He actually had a pair of graduate students Geiger & Marsden do the first rounds of experiments.] He expected the α particles to pass through with minor and occasional deflections.

  - The results were astounding [poor Geiger and Marsden first suffered Rutherford's wrath and were told to try again—since this couldn't be!].



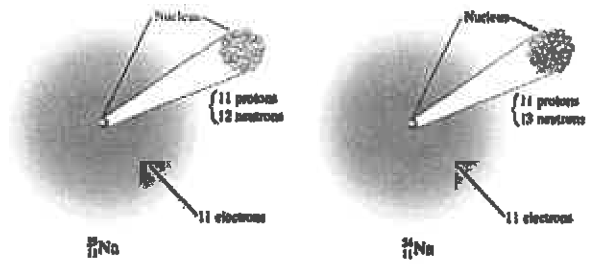
    - Most of the α particles did pass straight through, BUT many were deflected at LARGE angles and some even REFLECTED!
    - Rutherford stated that was like "shooting a howitzer at a piece of tissue paper and having the shell reflected back"
    - He knew the plum pudding model could not be correct!

    - Those particles with large deflection angles had a "close encounter" with the dense positive center of the atom
    - Those that were reflected had a "direct hit"
    - He conceived the **nuclear atom**; that with a dense (+) core or **nucleus.**
    - This center contains most of the mass of the atom while the remainder of the atom is empty space!



ISOTOPES

- **isotopes**--atoms having the same atomic number (# of p+) but a different number of neutrons
    - most elements have at least two stable isotopes, there are very few with only one stable isotope (Al, F, P)
    - Hydrogen's isotopes are so important they have special names:
        - 0 neutrons ☛ hydrogen
        - 1 neutron ☛ deuterium
        - 2 neutrons ☛ tritium

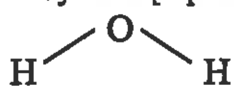


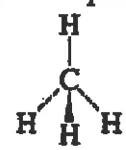
MOLECULES AND IONS

Electrons are the only subatomic particles involved in bonding and chemical reactivity.

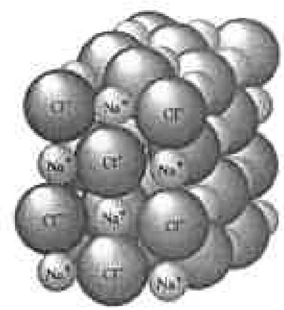
- **Chemical bonds**—forces that hold atoms together
- **Covalent bonds**—atoms share electrons and make molecules [independent units]; H_2, CO_2, H_2O, NH_3, O_2, CH_4 to name a few.
- **molecule**--smallest unit of a compound that retains the chem. characteristics of the compound; characteristics of the constituent elements are lost.
- **molecular formula**--uses symbols and subscripts to represent the composition of the molecule. (Strictest sense--covalently bonded)
- **structural formula**—bonds are shown by lines [representing shared e^- pairs]; may NOT indicate shape

H—O—H



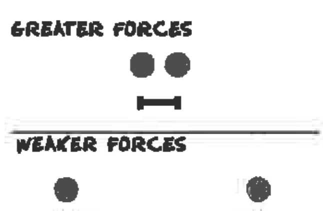


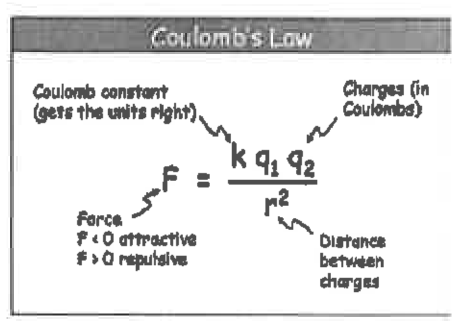
Methane



- **ions**--formed when electrons are lost or gained in ordinary chem. reactions; dramatic change in size (more about that shortly)
    - **cations**--(+) ions; often metals since metals *lose* electrons to become *positively* charged
    - **anions**--(–) ions; often nonmetals since nonmetals *gain* electrons to become *negatively* charged
    - **polyatomic ions**--units of atoms behaving as one entity ☛ It's worth memorizing 9 polyatomic ions + 3 patterns. (separate handout)
    - **ionic solids**—Electrostatic forces hold ions together. We can calculate the magnitude of them using Coulomb's Law. When these electrostatic attractions are strong, the ions are held together tightly and are close together ∴ solids. Additionally, the stronger the Coulombic force, F_c, the higher the melting point.

Coulomb's Law: $F_c \propto \dfrac{q_1 q_2}{d^2}$ or $F_c = k\dfrac{q_1 q_2}{d^2}$





NAMING SIMPLE COMPOUNDS

BINARY IONIC COMPOUNDS

Naming (+) ions: usually metals

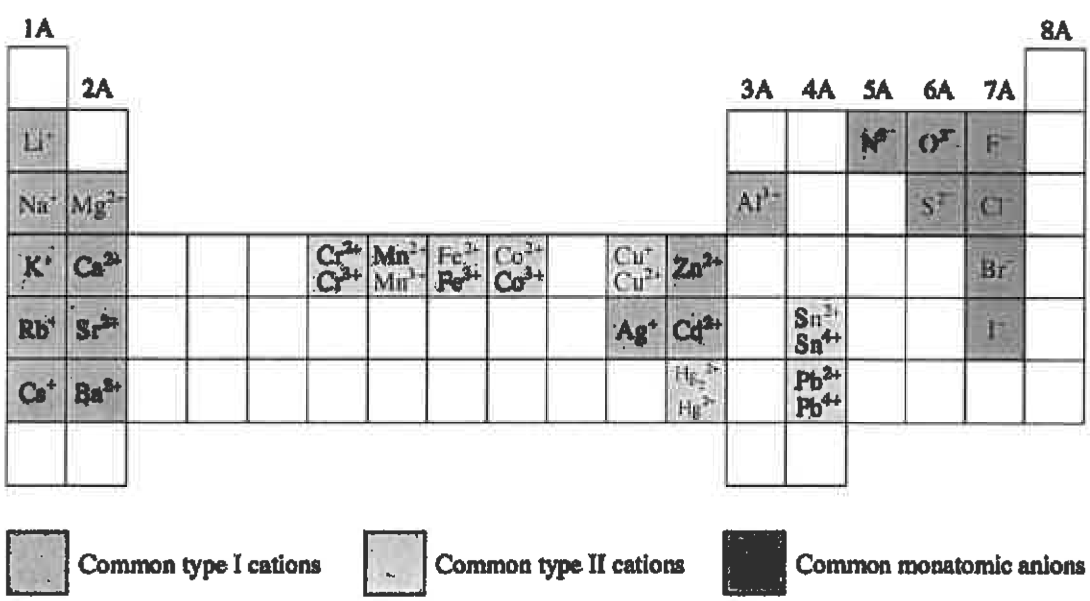
- monatomic, metal, cation → simply the name of the metal from which it is derived. **Al^{3+} is the aluminum ion;** transition metals form *more than one* ion; Roman Numerals (in) follow the ion's name, they are your friend—they tell you which charge is on that particular ion **Cu^{2+} is copper(II);** **Mercury(I)** ***is an exception ☞☐it is Hg_2^{2+} ∴ two Hg^+ associated together*** also, remember Hg is a metal that is a liquid at room temperature. (Yeah, the no space thing between the ion's name and (Roman Numeral) looks strange, but it is the correct way to do it. It's called the Stock system developed by the German chemist Alfred Stock and first published in 1916.)
- NH_4^+ is ammonium
- NO ROMAN NUMERAL IS USED WITH silver, cadmium and zinc. Why not? They only make one valence state.
  [Arrange their SYMBOLS in alphabetical order—first one is 1+ and the other two are 2+]

Naming (–) ions: monatomic and polyatomic

- **MONATOMIC**--add the suffix ***-ide*** to the stem of the nonmetal's name. Halogens are called the ***halides***.
- **POLYATOMIC**--quite common; *oxyanions* are the PA's containing oxygen (Go, figure!)
  - *hypo*--"ate" the least oxygen
  - *-ite*--"ate" 1 more oxygen than hypo-
  - *-ate*--"ate" 1 more oxygen than -ite
  - *hyper*---ate--"ate" the most oxygen (often the "hy" is left off to read simply "per")

Example: hypochlorite ClO^-
Chlorite ClO_2^-
Chlorate ClO_3^-
Hyper or more commonly Perchlorate ClO_4^-
You can substitute any halogen in for the Cl.

NAMING IONIC COMPOUNDS: The + ion name is given *first* followed by the name of the negative ion.



Exercise 5 **Naming Binary Compounds**

Give the systematic name of each of the following compounds.

a. $CoBr_2$ b. $CaCl_2$ c. Al_2O_3 d. $CrCl_3$

a. Cobalt(II) bromide; b. Calcium chloride; c. Aluminum oxide; d. Chromium(III) chloride

Exercise 6 **Naming Compounds Containing Polyatomic Ions**

Give the systematic name of each of the following compounds.

a. Na_2SO_4 b. KH_2PO_4 c. $Fe(NO_3)_3$ d. $Mn(OH)_2$
e. Na_2SO_3 f. Na_2CO_3 g. $NaHCO_3$ h. $CsClO_4$
i. $NaOCl$ j. Na_2SeO_4 k. $KBrO_3$

a. Sodium sulfate; b. Potassium dihydrogen phosphate; c. Iron(III) nitrate; d. Manganese(II) hydroxide; e. Sodium sulfite; f. Sodium carbonate; g. Sodium hydrogen carbonate; h. Cesium perchlorate; i. Sodium hypochlorite; j. Sodium selenite; k. Potassium bromate

NAMING BINARY COVALENT COMPOUNDS: (covalently bonded)

- Use prefixes!!! Don't forget the –ide ending as well.

Exercise 7 **Naming Type III Binary Compounds**

Name each of the following compounds.

a. PCl_5 b. PCl_3 c. SF_6 d. SO_3 e. SO_2 f. CO_2

a. Phosphorus pentachloride; b. Phosphorus trichloride; c. Sulfur hexafluoride; d. Sulfur trioxide; e. Sulfur dioxide; f. Carbon dioxide

ACIDS

Naming acids is actually easy. The nomenclature follows quite an elegant pattern:

Hydrogen, if present, is listed first followed by a suffix and finally the word "acid".

If the negative ion's name ends in:

-ide → *hydro*[negative ion root]*ic* acid Ex: hydrosulfuric acid, H_2S
-ate → *-ic* acid Ex: sulfuric acid, H_2SO_4
-ite → *-ous* acid Ex: chlorous acid, H_2SO_3

Exercise 9 **Naming Various Types of Compounds**

Give the systematic name for each of the following compounds.

a. P_4O_{10} b. Nb_2O_5 c. Li_2O_2 d. $Ti(NO_3)_4$

a. Tetraphosphorus decoxide; b. Niobium(V) oxide; c. Lithium peroxide; d. Titanium(IV) nitrate

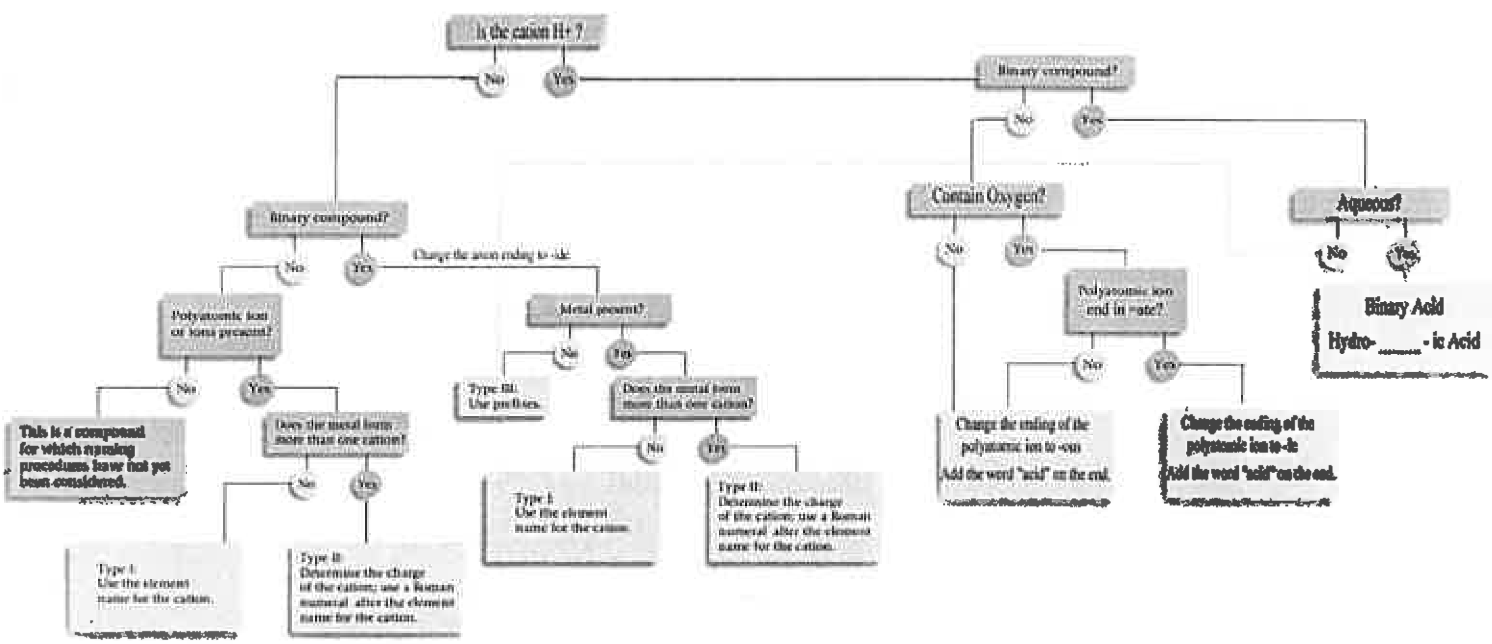
Exercise 10 **Writing Compound Formulas from Names**

Given the following systematic names, write the formula for each compound.

a. Vanadium(V) fluoride b. Dioxygen difluoride
c. Rubidium peroxide d. Gallium oxide

a. VF_5; b. O_2F_2; c. Rb_2O_2 d. Ga_2O_3

For the visual learners among you, here's a "Cheat Sheet". Practice, practice, practice!



NATIONAL MATH + SCIENCE INITIATIVE

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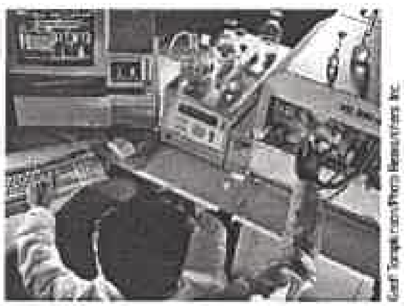
AP* Chemistry
Stoichiometry

ATOMIC MASSES

- **^{12}C—Carbon 12**—In 1961 it was agreed that this isotope of carbon would serve as the standard used to determine all other atomic masses and would be *defined* to have a mass of EXACTLY 12 atomic mass units (amu). All other atomic masses are measured ***relative*** to this.
- **mass spectrometer**—a device for measuring the mass of atoms or molecules



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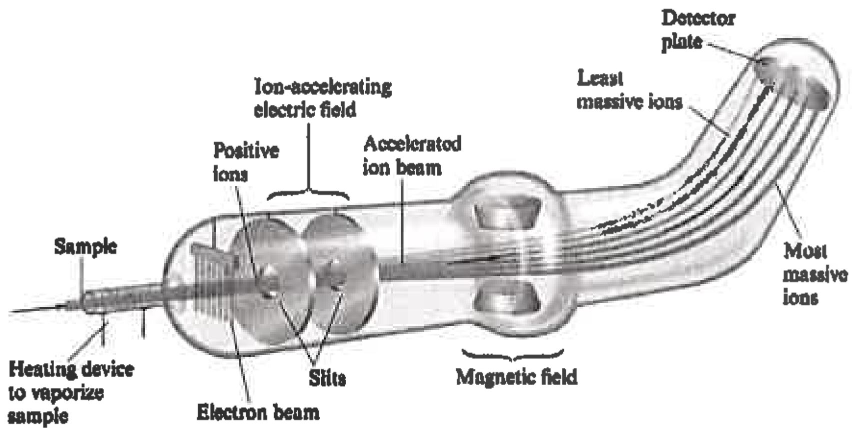


  - atoms or molecules are passed into a beam of high-speed electrons
  - this knocks electrons OFF the atoms or molecules transforming them into cations
  - apply an electric field
  - this accelerates the cations since they are repelled from the (+) pole and attracted toward the (−) pole
  - send the accelerated cations into a magnetic field
  - an accelerated cation creates it's OWN magnetic field which perturbs the original magnetic field
  - this perturbation changes the path of the cation
  - the amount of deflection is proportional to the mass; heavy cations deflect little
  - ions hit a detector plate where measurements can be obtained.

  - $\dfrac{\text{Mass}\,^{13}C}{\text{Mass}\,^{12}C} = 1.0836129 \therefore \text{Mass}\,^{13}C = (1.0836129)(\underbrace{12\ \text{amu}}_{\text{Exact by definition}}) = 13.003355\ \text{amu}$

- **average atomic masses**—atoms have masses of whole numbers, HOWEVER samples of quadrillions of atoms have a few that are heavier or lighter [isotopes] due to different numbers of neutrons present
- **percent abundance**--percentage of atoms in a natural sample of the pure element represented by a particular isotope

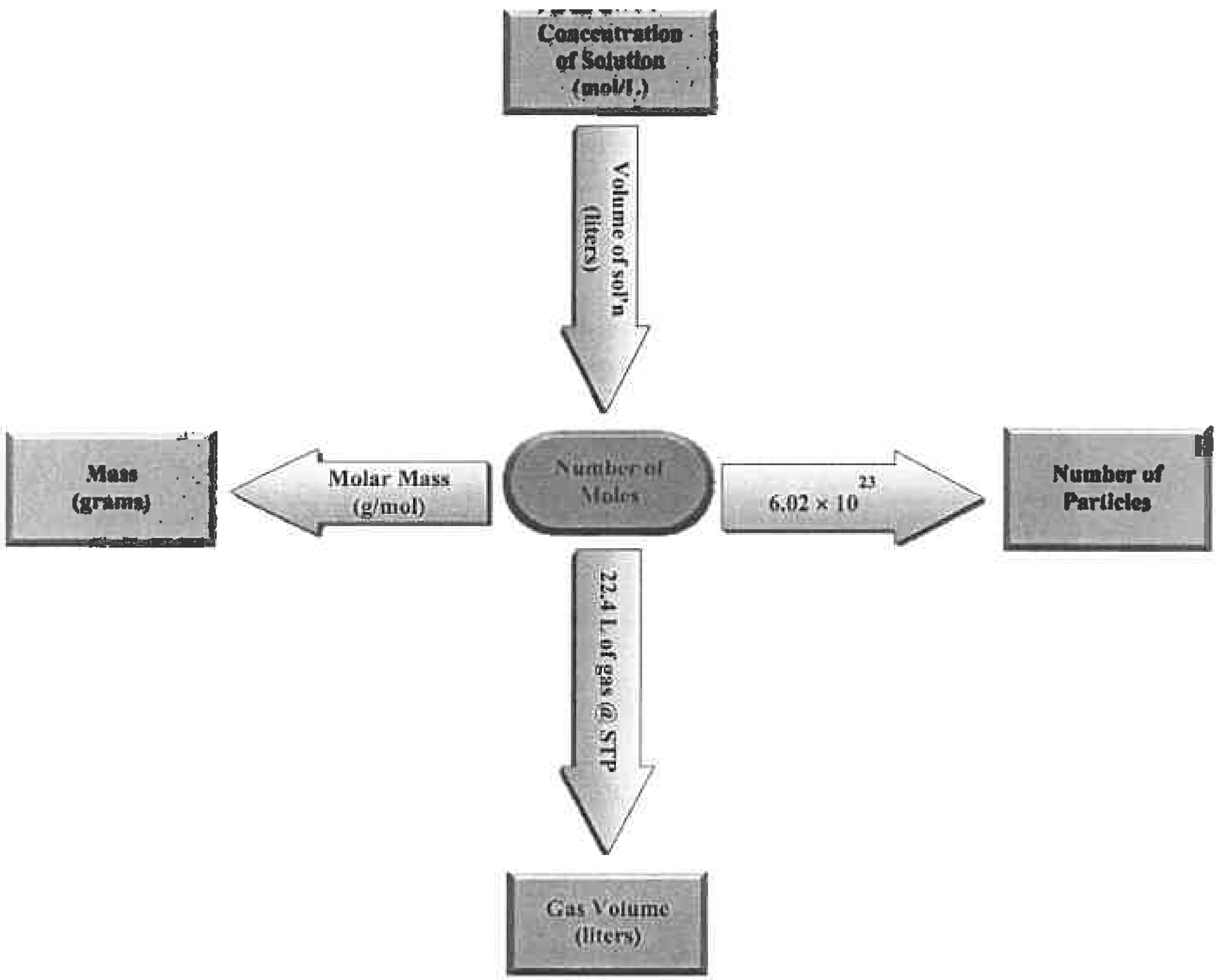
$$\text{percent abundance} = \frac{\text{number of atoms of a given isotope}}{\text{Total number of atoms of all isotopes of that element}} \times 100\%$$

- **counting by mass**—when particles are small this is a matter of convenience. Just as you buy 5 lbs of sugar rather than a number of sugar crystals, or a pound of peanuts rather than counting the individual peanuts….this concept works very well if your know an *average* mass.

ALTERNATE TECHNIQUE #1—USING THE MOLE MAP:

Simply reproduce this map on your scratch paper until you no longer need to since the image will be burned into your brain!

MULTIPLY [by the conversion factor on the arrow] when "traveling" IN THE DIRECTION OF THE ARROW and obviously, divide when "traveling" against an arrow.



Exercise 5 **Calculating Molar Mass I**

Juglone, a dye known for centuries, is produced from the husks of black walnuts. It is also a natural herbicide (weed killer) that kills off competitive plants around the black walnut tree but does not affect grass and other noncompetitive plants [a concept called *allelopathy*]. The formula for juglone is $C_{10}H_6O_3$.

(a) **Calculate** the molar mass of juglone.

(b) A sample of 1.56×10^{-2} g of pure juglone was extracted from black walnut husks. **Calculate** the number of moles of juglone present in this sample.

a. 174.1 g
b. 8.96×10^{-5} mol juglone

Exercise 6 **Calculating Molar Mass II**

Calcium carbonate ($CaCO_3$), also called *calcite*, is the principal mineral found in limestone, marble, chalk, pearls, and the shells of marine animals such as clams.

(a) **Calculate** the molar mass of calcium carbonate.

(b) A certain sample of calcium carbonate contains 4.86 moles. **Calculate** the mass in grams of this sample. **Calculate** the mass of the CO_3^{2-} ions present.

a. 100 g/mol
b. 486 g; 292g CO_3^{2-}

NEXT, THE MASS PERCENT CAN BE CALCULATED:

$$\text{Mass percent of C} = \frac{24.02 \text{ g C}}{46.08 \text{ g}} \times 100\% = 52.14\%$$

Repeat for the H and O present.

Exercise 8 Calculating Mass Percent I

Carvone is a substance that occurs in two forms having different arrangements of the atoms but the same molecular formula ($C_{10}H_{14}O$) and mass. One type of carvone gives caraway seeds their characteristic smell, and the other type is responsible for the smell of spearmint oil. **Calculate** the mass percent of each element in carvone.

C = 79.96%
H = 9.394%
O = 10.65%

Exercise 9 Calculating Mass Percent II

Penicillin, the first of a now large number of antibiotics (antibacterial agents), was discovered accidentally by the Scottish bacteriologist Alexander Fleming in 1928, but he was never able to isolate it as a pure compound. This and similar antibiotics have saved millions of lives that might have been lost to infections. Penicillin F has the formula $C_{14}H_{20}N_2SO_4$. Calculate the mass percent of each element.

C = 53.82%
H = 6.47%
N = 8.97%
S = 10.26%
O = 20.49%

Next, you can see that all of the hydrogen ended up in H_2O, so….calculate THE NUMBER OF MOLES!!

So, 0.1676 g H_2O ÷ 18.02 g/mol H_2O = 0 .009301 moles of H_2O, BUT there are 2 moles of H for each mole of water [Think "organ bank" one heart per body, one C per molecule of carbon dioxide while there are 2 lungs per body, 2 atoms H in water and so on…] thus, DOUBLE THE NUMBER OF MOLES of H_2O GIVES THE NUMBER OF MOLES OF HYDROGEN!! moles H = 2 × 0 .009301 moles of H_2O = 0.01860 moles of H

Therefore, the remaining mass must be nitrogen, BUT we only have mass data for the sample so convert your moles of C and H to grams:

grams C = 0.003781 moles C × 12.01 $\frac{\text{g}}{\text{mol}}$ = 0.04540 grams C

grams H = 0.01860 moles H × 1.01 $\frac{\text{g}}{\text{mol}}$ = <u>0.01879 grams H</u>

Total grams : **0.06419 total grams accounted for thus far**

What to do next? SUBTRACT!

0.1156 g sample – **0.06419 total grams accounted for thus far** = grams N left = 0.05141 g N so….

0.05141 g N ÷ 14.01 $\frac{\text{g}}{\text{mol}}$ = **0.003670 moles N**

Next, realize that chemical formulas represent **mole to mole ratios**, so…divide the number of moles of each by the smallest # of moles for any one of them to get a guaranteed ONE in your ratios…multiply by 2, then 3, etc to get to a ratio of small whole numbers. Clear as mud? WATCH THE SCREENCAST!!

Element	# moles	ALL Divided by the smallest (0.003670 moles)
C	**0.003781**	**1**
H	**0.01860**	**5**
N	**0.003670**	**1**

Therefore, the correct EMPIRICAL formula based on the data given is CH_5N.

Finally (this is drumroll worthy), IF we are told that the MM of the original substance is 31.06 g/mol, then simply use this relationship:

(Empirical mass) × n = MM

(12.01 + 5.05 + 14.01) × n = 31.07 g/mol ∴ n = 0.999678

This is mighty close to 1.0! Thus, the empirical formula and the molecular formula are one and the same.

Exercise 12
Caffeine, a stimulant found in coffee, tea, and chocolate, contains 49.48% carbon, 5.15% hydrogen, 28.87% nitrogen, and 16.49% oxygen by mass and has a molar mass of 194.2 g/mol. Determine the molecular formula of caffeine.

Molecular formula = $C_8H_{10}N_4O_2$

BALANCING CHEMICAL EQUATIONS

Chemical reactions are the result of a chemical change where atoms are reorganized into one or more new arrangements. Bonds are broken [*requires energy*] and new ones are formed [*releases energy*]. A chemical reaction transforms elements and compounds into new substances. A *balanced chemical equation* shows the relative amounts of reactants [on the left] and products [on the right] by molecule or by mole.

Subtle details:

- *s, l, g, aq*—state symbols that correspond to solid, liquid, gas, aqueous solution
- NO ENERGY or TIME is alluded to
- Antoine Lavoisier (1743-1794)—The Law of Conservation of Matter: ***matter can be neither created nor destroyed*** ☛ this means you having to "balance equations" is entirely his fault!!

State	*Symbol*
Solid	(*s*)
Liquid	(*l*)
Gas	(*g*)
Dissolved in water (in aqueous solution)	(*aq*)

BALANCING CHEMICAL EQUATIONS

- Begin with the most complicated-looking thing (often the scariest, too).
- Save the elemental thing for last.
- If you get stuck, double the most complicated-looking thing.
- MEMORIZE THE FOLLOWING:
- metals + halogens → M_aX_b
- CH and/or O + O_2 →# $CO_2(g) + H_2O(g)$
- H_2CO_3 [any time formed!] → $CO_2 + H_2O$; in other words, never write carbonic acid as a product, it spontaneously decomposes [in an open container] to become carbon dioxide and water
- metal carbonates → metal OXIDES + CO_2

STOICHIOMETRIC CALCULATIONS: AMOUNTS OF REACTANTS AND PRODUCTS

Stoichiometry – The study of quantities of materials consumed and produced in chemical reactions. Stoichiometry is the most important thing you can learn as you embark upon AP Chemistry! Get good at this and you will do well all year. This NEVER goes away!

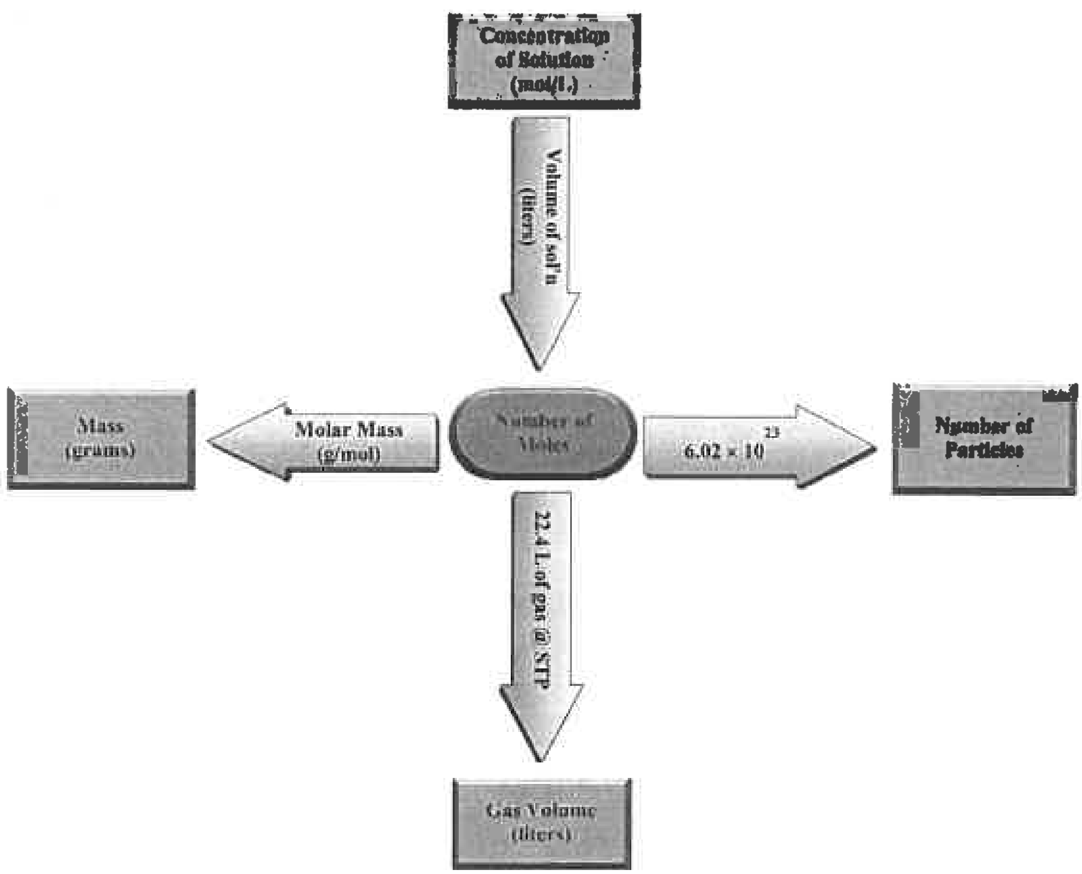
It's time to repeat my dimensional analysis disclaimer.

> **DIMENSIONAL ANALYSIS DISCLAIMER:** I will show you some alternatives to dimensional analysis. WHY? First, some of these techniques are faster and well-suited to the multi-step problems you will face on the AP Exam. Secondly, these techniques better prepare you to work the complex equilibrium problems you will face later in this course. The first problem you must solve in the free response section of the AP Exam will be an equilibrium problem and you will need to be able to work them quickly. Lastly, I used to teach both methods. Generations of successful students have encouraged me to share these techniques with as many students as possible. They did, once they got to college, and made lots of new friends once word got out they had this "cool way" to solve stoichiometry problems—not to mention their good grades! Give this a try. It doesn't matter which method you use, I encourage you to use the method that works best for you and lets you solve problems *accurately and quickly*!

First you have to be proficient at the following no matter which method you choose!:

- Writing CORRECT formulas—this requires knowledge of your polyatomic ions and being able to use the periodic table to deduce what you have not had to memorize. Review section 2.8 in your Chapter 2 notes or your text.
- Calculate CORRECT molar masses from a correctly written formula
- Balance a chemical equation
- Use the mole map to calculate the number of moles or anything else!

Remember the mole map? It will come in mighty handy as well!



Now, humor me…What if part (b) asked for liters of CO_2 at STP [1 atm, 273K]?

Use the mole map. Start in the middle with 6.53 moles × [in direction of arrow] 22.4 L/mol = 146 L

Molar Mass:	(44.11)	(32.00)	(44.01)	(18.02)
Balanced Eq'n	C_3H_8	+ 5 O_2	→ 3 CO_2	+ 4 H_2O
mole:mole	**1**	**5**	**3**	**4**
# moles	2.18	10.9	6.53	8.71
amount	96.1 grams	349 g	146 L	

What if part (c) asked you to calculate how many water molecules are produced?

Use the mole map , start in the middle with 8.71 mol water × $6.02 \times 10^{23} \frac{\text{molecules}}{\text{mol}} = 5.24 \times 10^{24}$ molecules of water.

Try these two exercises with whichever method you like best!

Exercise 15

Solid lithium hydroxide is used in space vehicles to remove exhaled carbon dioxide from the living environment by forming solid lithium carbonate and liquid water. What mass of gaseous carbon dioxide can be absorbed by 1.00 kg of lithium hydroxide?

920. g

Plan of attack: First, you must realize that you even ***need*** a plan of attack! **IF ever you are faced with TWO starting amounts of matter reacting, you have entered "The Land of Limiting Reactant".**

When faced with this situation …calculate the number of moles of everything you are given. Set up your table like before, only now you'll have TWO amounts and thus TWO # 's of moles to get you started.

Cover one set of moles up (pretending you only had one amount to work from) and ask yourself, "What if all of these moles reacted?" "How many moles of the other reactants would I need to use up all of these moles?" Next, do the calculation of how many moles of the "other" amount(s) you would need. Do you have enough? If so, the reactant you began with IS the limiting reactant. If not repeat this process with the "other" reactant amount you were given.

It doesn't matter where you start the "What if?" game….you get there either way.

Clear as mud? Read on…(and consider listening to the SCREENCAST!)

Let's revisit the Haber process:

Molar Mass:	(28.02)		(2.02)		(17.04)
Balanced Eq'n	N_2	+	$3\ H_2$	$\rightarrow$	$2\ NH_3$
mole:mole	**1**		**3**		**2**
# moles					
amount					

Suppose 25.0 kg of nitrogen reacts with 5.00 kg of hydrogen to form ammonia. What mass of ammonia can be produced? Which reactant is the limiting reactant? What is the mass of the reactant that is in excess?

*****Insert the masses in the amount row and find the number of moles of BOTH!***

Molar Mass:	(28.02)		(2.02)		(17.04)
Balanced Eq'n	N_2	+	$3\ H_2$	$\rightarrow$	$2\ NH_3$
mole:mole	**1**		**3**		**2**
# moles	892 moles		2,475 moles		
amount	25,000 g		5,000 g		

WHAT IF I used up all the moles of hydrogen? I'd need **1/3** × 2,475 moles = 825 moles of nitrogen. Clearly I have EXCESS moles of nitrogen!! Therefore, hydrogen limits me.

OR

WHAT IF I used up all the moles of nitrogen? I'd need **3** × **892** moles = 2,676 moles of hydrogen. Clearly I don't have enough hydrogen, so it limits me!! Therefore nitrogen is in excess.

Continued on next page.

Theoretical Yield: The amount of product formed when a limiting reactant is completely consumed. This assumes perfect conditions and gives a maximum amount!! Not likely!

Actual yield: That which is realistic.

Percent yield: The ratio of actual to theoretical yield.

$$\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\% = \text{Percent yield}$$

Exercise 19

Methanol (CH_3OH), also called *methyl alcohol*, is the simplest alcohol. It is used as a fuel in race cars and is a potential replacement for gasoline. Methanol can be manufactured by combination of gaseous carbon monoxide and hydrogen. Suppose 68.5 kg $CO(g)$ is reacted with 8.60 kg $H_2(g)$. Calculate the theoretical yield of methanol. If 3.57×10^4 g CH_3OH is actually produced, what is the percent yield of methanol?

Theoretical yield is 6.82×10^4 g
Percent yield is 52.3%