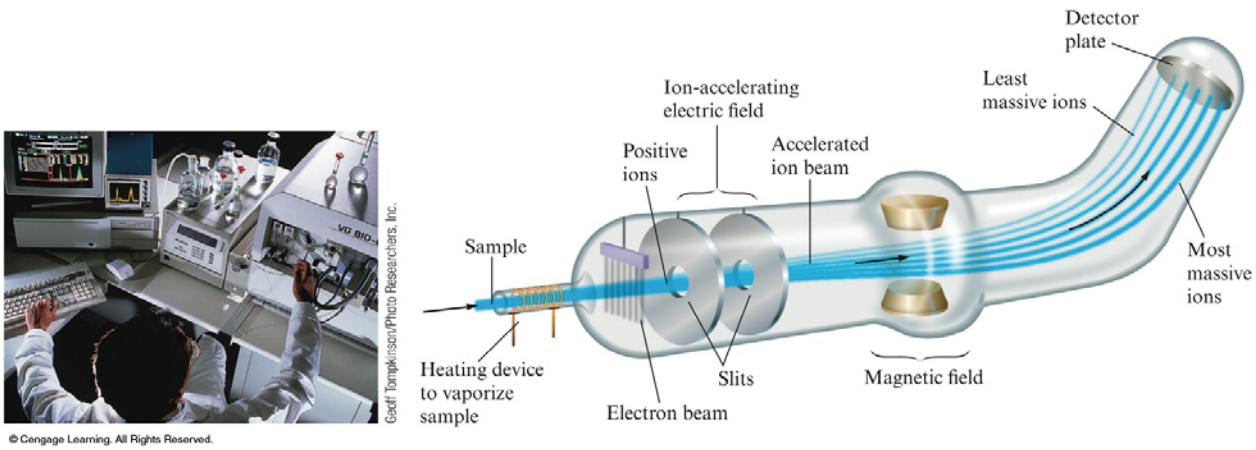
AP* Chemistry
Stoichiometry

ATOMIC MASSES

- **<u>^{12}C—Carbon 12</u>**—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.
- **<u>mass spectrometer</u>**—a device for measuring the mass of atoms or molecules



  - 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.

  - $\frac{\text{Mass } ^{13}\text{C}}{\text{Mass } ^{12}\text{C}} = 1.0836129 \therefore \text{Mass } ^{13}\text{C} = (1.0836129)(12 \text{ amu}) = 13.003355 \text{ amu}$

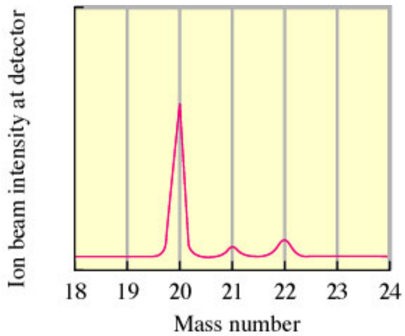
Exact by definition

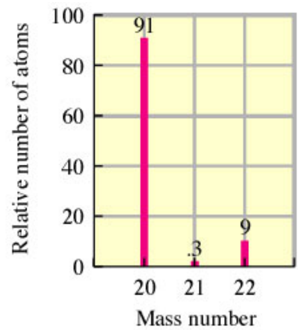
- **<u>average atomic masses</u>**—atoms have masses of whole numbers, HOWEVER samples of quadrillions of atoms have a few that are heavier or lighter [isotopes] due to different numbers of neutrons present
- **<u>percent abundance</u>**--percentage of atoms in a natural sample of the pure element represented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

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

- **<u>counting by mass</u>**—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.

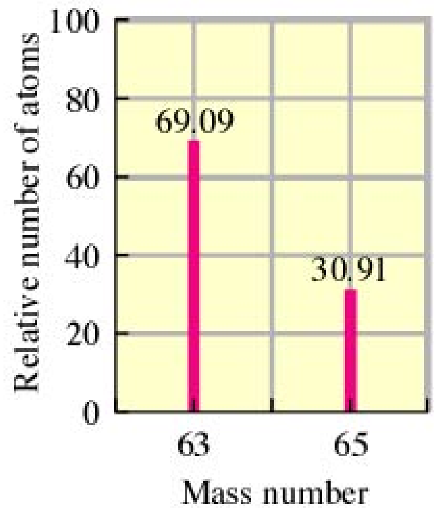
- **<u>mass spectrometer to determine isotopic composition</u>**—load in a pure sample of natural neon or other substance. The areas of the "peaks" or heights of the bars indicate the relative abundances of ${}^{20}_{10}Ne$, ${}^{21}_{10}Ne$, and ${}^{22}_{10}Ne$





Exercise 1 The Average Mass of an Element

When a sample of natural copper is vaporized and injected into a mass spectrometer, the results shown in the figure are obtained. Use these data to **calculate** the average mass of natural copper. (The mass values for ${}^{63}Cu$ and ${}^{65}Cu$ are 62.93 amu and 64.93 amu, respectively.)



63.55 amu/atom

THE MOLE

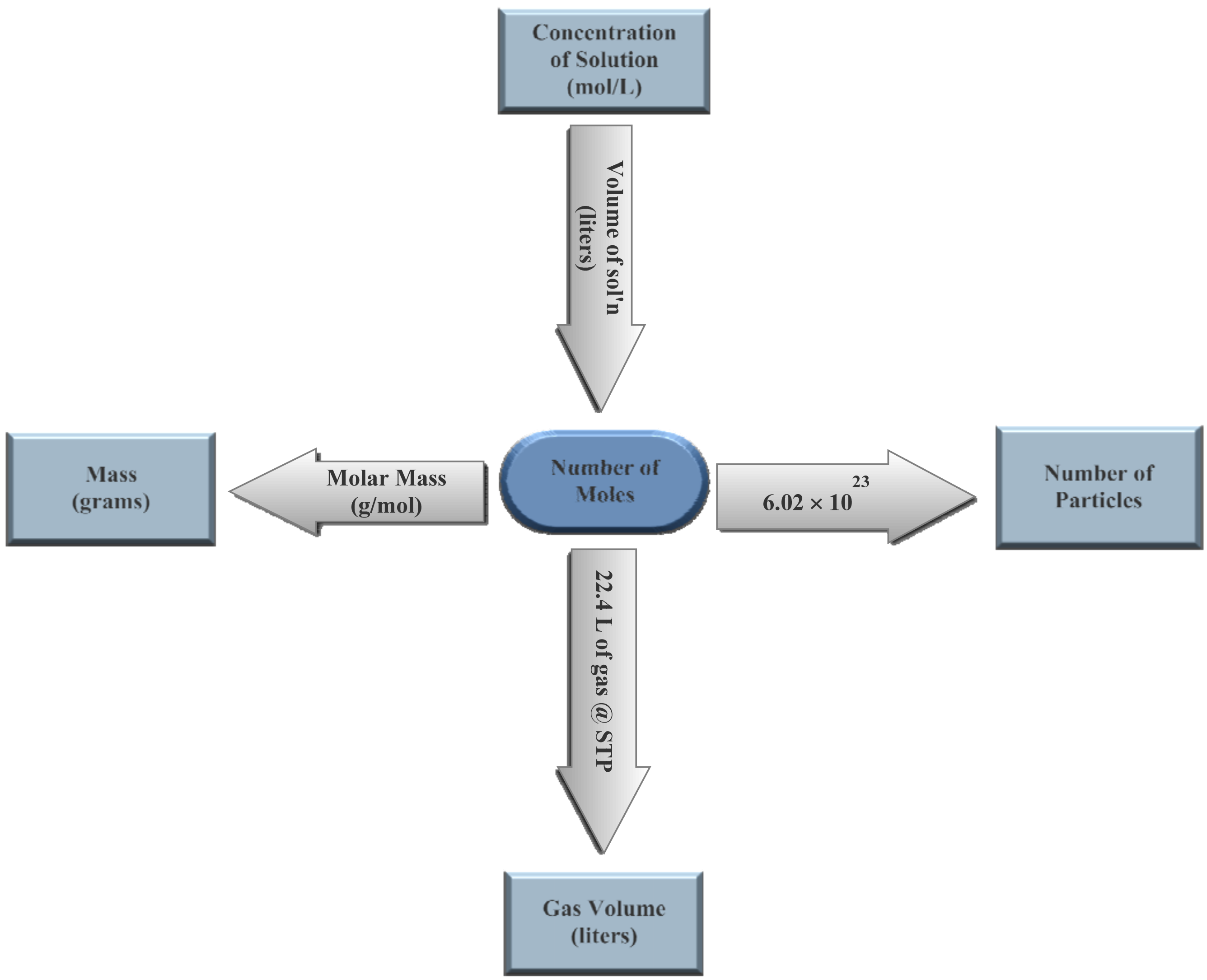
- **<u>mole</u>**—the number of C atoms in exactly 12.0 grams of ${}^{12}C$; also a number, 6.02×10^{23} just as the word "dozen" means 12 and "couple" means 2.
- **<u>Avogadro's number</u>**—6.02×10^{23}, the number of particles in a mole of anything

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. 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 themselves did, once they got to college, and made lots of new friends once word got out they had this "easy 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*!

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 2 **Determining the Mass of a Sample of Atoms**
Americium is an element that does not occur naturally. It can be made in very small amounts in a device known as a *particle accelerator*. **Calculate** the mass in grams of a sample of americium containing six atoms.

2.42×10^{-21} g

Exercise 3 **Determining Moles of Atoms**
Aluminum is a metal with a high strength-to-mass ratio and a high resistance to corrosion; thus it is often used for structural purposes. **Calculate** both the number of moles of atoms and the number of atoms in a 10.0-g sample of aluminum.

0.371 mol Al
2.23×10^{23} atoms

Exercise 4 **Calculating the Number of Moles and Mass**
Cobalt (Co) is a metal that is added to steel to improve its resistance to corrosion. **Calculate** both the number of moles in a sample of cobalt containing 5.00×10^{20} atoms and the mass of the sample.

8.31×10^{-4} mol Co
4.89×10^{-2} g Co

MOLAR MASS AND FORMULA WEIGHT

- **<u>molar mass, *MM*</u>**--the sum of all of the atomic masses in a given chemical formula in units of g/mol. It is also equal mass in grams of Avogadro's number of molecules; i.e. the mass of a mole
- **<u>empirical formula</u>**--the ratio in the network for an ionic substance
- **<u>formula weight</u>**--same as molecular weight, just a language problem ☞ "molecular" implies covalent bonding while "formula" implies ionic bonding {just consider this to be a giant conspiracy designed to keep the uneducated from *ever* understanding chemistry—kind of like the scoring scheme in tennis}. **Just use "molar mass" for all formula masses.**
- **A WORD ABOUT SIG. FIG.'s—**It is correct to "pull" from the periodic table the least number of sig. figs for your MM's as are in your problem—just stick with 2 decimal places for all MM's —much simpler!

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-}

Exercise 7 Molar Mass and Numbers of Molecules
Isopentyl acetate ($C_7H_{14}O_2$), the compound responsible for the scent of bananas, can be produced commercially. Interestingly, bees release about 1µg (1×10^{-6} g) of this compound when they sting. The resulting scent attracts other bees to join the attack.

(a) **Calculate** the number of molecules of isopentyl acetate released in a typical bee sting.

(b) **Calculate** the number of carbon atoms present.

5×10^{15} molecules
4×10^{16} carbon atoms

ELEMENTS THAT EXIST AS MOLECULES

Pure hydrogen, nitrogen, oxygen and the halogens exist as **DIATOMIC** molecules *under normal conditions.* MEMORIZE!!! Be sure you compute their molar masses as diatomics. We lovingly refer to them as the "gens", "Hydro**gen**, oxy**gen,** nitro**gen** & the halo**gen**s!"

Others to be aware of, but not memorize:

- P_4—tetratomic form of elemental phosphorous; an allotrope
- S_8—sulfur's elemental form; also an allotrope
- Carbon—diamond and graphite ☞covalent networks of atoms

PERCENT COMPOSITION OF COMPOUNDS

There are two common ways of describing the composition of a compound: 1) in terms of the number of its constituent atoms and 2) in terms of the percentages (by mass) of its elements.

Percent Composition (by mass): The Law of Constant Composition states that *any sample of a pure compound always consists of the same elements combined in the same proportions by mass.* Remember, all percent calculations are simply $\frac{\text{part}}{\text{whole}} \times 100\%$

$$\%\text{ comp} = \frac{\text{mass of desired element}}{\text{total mass of compound}} \times 100\%$$

Consider ethanol, C_2H_5OH

$$\text{Mass of C} = 2\ \cancel{\text{mol}} \times 12.01\frac{g}{\cancel{mol}} = 24.02\text{ g}$$

$$\text{Mass of H} = 6\ \cancel{\text{mol}} \times 1.01\frac{g}{\cancel{mol}} = 6.06\text{ g}$$

$$\text{Mass of O} = 1\ \cancel{\text{mol}} \times 16.00\frac{g}{\cancel{mol}} = \underline{16.00\text{g}}$$

$$\therefore \text{ Mass of 1 mol of } C_2H_5OH = 46.08\text{ g}$$

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.

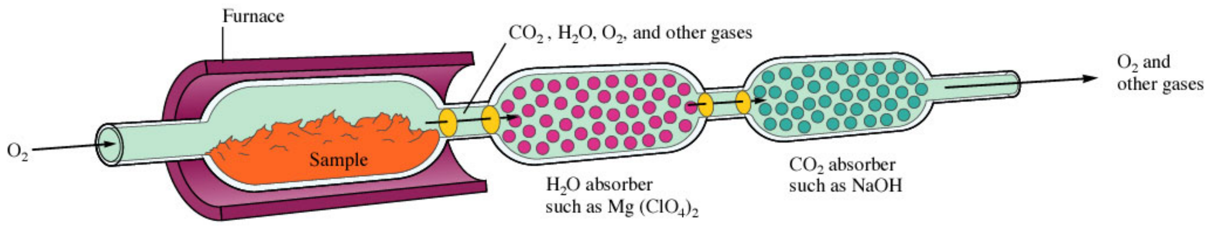
C = 79.96%
H = 9.394%
O = 10.65%

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%

DETERMINING THE FORMULA OF A COMPOUND



When faced with a hydrocarbon compound of "unknown" formula, one of the most common techniques is to combust it with oxygen to produce oxides of the nonmetals CO_2 and H_2O which are then collected and weighed.

- **Calculating empirical and molecular formulas:** empirical formulas represent the *simplest or smallest ratio of elements within a compound* while molecular formulas represent the *actual numbers of elements within a compound*. The empirical mass is the **least common multiple** of the molar mass.

  Example: CH_2O is the empirical for a carbohydrate—get it? "carbon waters".

  Anyway, glucose is a perfect example of a carbohydrate (a sugar to be exact) with an empirical molar mass of 12 + 2(1) + 16 = 30 g/mol and since glucose is 6 units of CH_2O which is equivalent to $(CH_2O)_6$ or $C_6H_{12}O_6$; the empirical mass of 30 is also multiplied by 6. Thus the MM of glucose is 180 g/mol.

- Make your problem solving life easy and ***assume a 100 gram sample if given %'s***—that way you can convert the percents given directly into grams and subsequently into moles in order to simplify your life!

Other twists and turns occurring when calculating molar masses involve:

- **hydrates**—waters of hydration or "dot waters". They count in the calculation of molar masses for hydrates and used to "cement" crystal structures together

- **anhydrous**—means *without* water—just to complete the story—just calculate the molar masses of anhydrous substances as you would any other substance

Example:

A compound is composed of carbon, nitrogen and hydrogen. When **0.1156 g of this compound** is reacted with oxygen [a.k.a. "burned in air" or "combusted"], 0.1638 g of carbon dioxide and 0.1676 g of water are collected. Determine the empirical formula of the compound.

So, Compound + O_2 $\rightarrow$ oxides of what is burned. In this case Compound + O_2 $\rightarrow CO_2 + H_2O + N_2$
(clearly not balanced)

You can see that all of the carbon ended up in CO_2 so…when in doubt, calculate THE NUMBER OF MOLES!!

0.1638 g CO_2 ÷ 44.01 g/mol CO_2 = 0.003781 moles of CO_2 = 0.003781 moles of C (why?)

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 $\times 12.01 \frac{\text{g}}{\text{mol}}$ = 0.04540 grams C

grams H = 0.01860 moles H $\times 1.01 \frac{\text{g}}{\text{mol}}$ = 0.01879 grams H

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 $\div 14.01 \frac{\text{g}}{\text{mol}}$ = **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.

Empirical Formula Determination

- Since mass percentage gives the number of grams of a particular element per 100 grams of compound, base the calculation on 100 grams of compound. Each percent will then represent the mass in grams of that element.
- Determine the number of moles of each element present in 100 grams of compound using the atomic masses of the elements present.
- Divide each value of the number of moles by the smallest of the values. If each resulting number is a whole number (after appropriate rounding), these numbers represent the subscripts of the elements in the empirical formula.
- If the numbers obtained in the previous step are not whole numbers, multiply each number by an integer so that the results are all whole numbers.

Exercise 10

Determine the empirical *and* molecular formulas for a compound that gives the following analysis in mass percents:

71.65% C1 24.27% C 4.07% H

The molar mass is known to be 98.96 g/mol.

Empirical formula = CH_2 C1
Molecular formula = $C_2H_4C1_2$

Exercise 11

A white powder is analyzed and found to contain 43.64% phosphorus and 56.36% oxygen by mass. The compound has a molar mass of 283.88 g/mol. What are the compound's empirical and molecular formulas?

Empirical formula = P_2O_5
Molecular formula = $(P_2O_5)_2$ or P_4O_{10}

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.

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

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!!

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

Table 3.2 Information Conveyed by the Balanced Equation for the Combustion of Methane

Reactants	$\longrightarrow$	*Products*
$CH_4(g) + 2O_2(g)$	$\longrightarrow$	$CO_2(g) + 2H_2O(g)$
1 molecule CH_4 + 2 molecules O_2	$\longrightarrow$	1 molecule CO_2 + 2 molecules H_2O
1 mol CH_4 molecules + 2 mol O_2 molecules	$\longrightarrow$	1 mol CO_2 molecules + 2 mol H_2O molecules
6.022×10^{23} CH_4 molecules + $2(6.022 \times 10^{23})$ O_2 molecules	$\longrightarrow$	6.022×10^{23} CO_2 molecules + $2(6.022 \times 10^{23})$ H_2O molecules
16 g CH_4 + 2(32 g) O_2	$\longrightarrow$	44 g CO_2 + 2(18 g) H_2O
80 g reactants	$\longrightarrow$	80 g products

Exercise 13

Chromium compounds exhibit a variety of bright colors. When solid ammonium dichromate, $(NH_4)_2Cr_2O_7$, a vivid orange compound, is ignited, a spectacular reaction occurs. Although the reaction is actually somewhat more complex, let's assume here that the products are solid chromium(III) oxide, nitrogen gas (consisting of N_2 molecules), and water vapor. Balance the equation for this reaction.

$$\mathbf{(NH_4)_2Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + 4H_2O(g)}$$

(4 × 2) H **(4 ×2) H**

http://www.youtube.com/watch?v=CW4hN0dYnkM

Exercise 14

At 1000ºC, ammonia gas, $NH_3(g)$, reacts with oxygen gas to form gaseous nitric oxide, $NO(g)$, and water vapor. This reaction is the first step in the commercial production of nitric acid by the Ostwald process. Balance the equation for this reaction.

$$\mathbf{4\,NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)}$$

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!

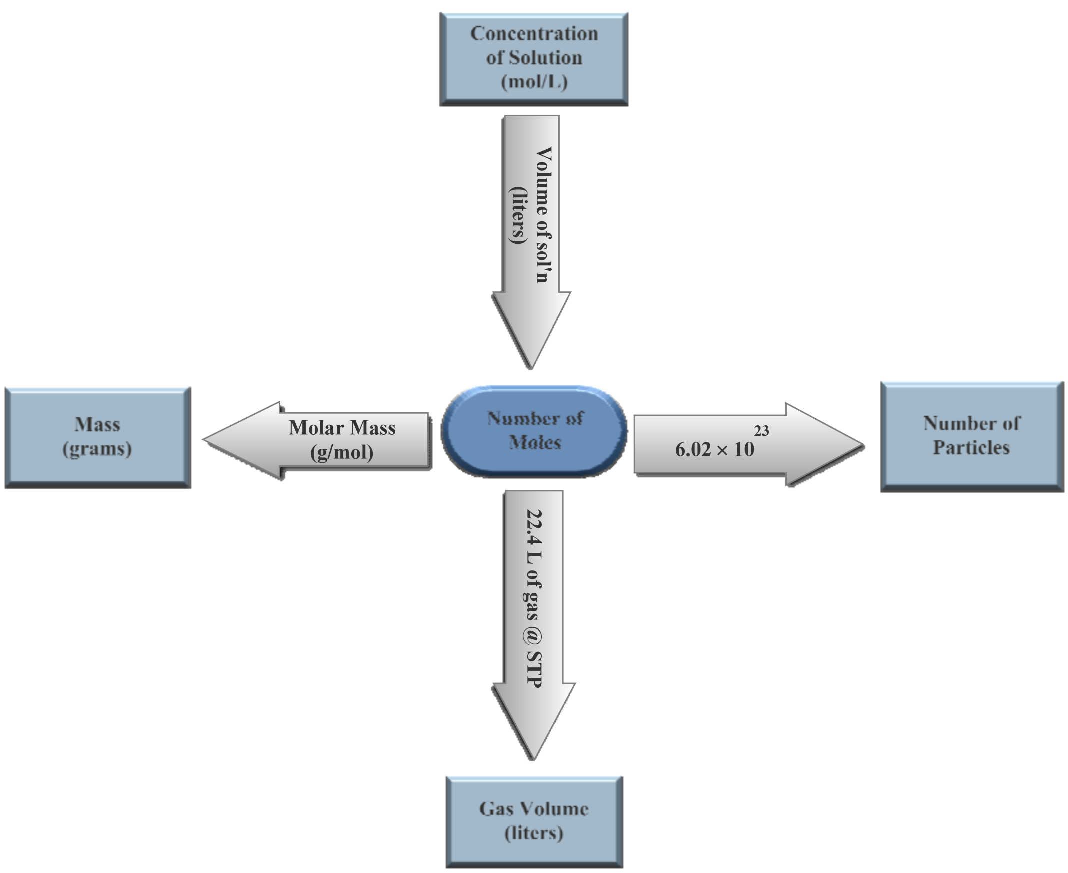
It's time to repeat my dimensional analysis disclaimer.

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!



Here's the "template" for solving the problems…you'll create a chart. Here's a typical example:

Example: Calculate the mass of oxygen will react completely with 96.1 grams of propane?
[notice all words—you supply the chemical formulas!]

Molar Mass:	(44.11)			(32.00)			(44.01)			(18.02)
Balanced Eq'n	C_3H_8	+	5	O_2	$\rightarrow$	3	CO_2	+	4	H_2O
mole:mole	**1**			**5**			**3**			**4**
# moles										
Amount										

1. Write a chemical equation paying special attention to writing correct chemical formulas!

2. Calculate the molar masses and put in parentheses above the formulas—soon you'll figure out you don't have to do this for every reactant and product, just those in which you are interested.

3. Balance the equation! Examine the coefficients on the balanced equation, they ARE the mole:mole ratios! Isolating them helps you internalize the mol:mol until you get the hang of this.

4. Next, re-read the problem and put in an **amount**—in this example it's 96.1 g of propane.

Molar Mass:	(44.11)			(32.00)			(44.01)			(18.02)
Balanced Eq'n	C_3H_8	+	5	O_2	$\rightarrow$	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									

5. Calculate the number of moles of something, anything! Use the mole map. Start at 96.1 grams of C_3H_8, divide the 96.1 g [against the arrow on the mole map] by molar mass to calculate the # moles of propane.

6. USE the mole: mole to find moles of EVERYTHING! If 1 = 2.18 then oxygen is 5(2.18) etc…. [IF the first mol amount you calculate is not a "1", just divide appropriately to make it "1" before moving on to calculate the moles of all the rest!] Leave everything in your calculator—I only rounded to save space!

7. Re-read the problem to determine which amount was asked for…here's the payoff….AP problems ask for several amounts! First, we'll find the mass of oxygen required since that's what the problem asked.
   10.9 moles × 32.00 g/mol = 349 g of oxygen

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	$\rightarrow$	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 × $6.02 \times 10^{23} \dfrac{\text{molecules}}{\text{mol}} = 5.24 \times 10^{24}$ 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

Exercise 16
Baking soda ($NaHCO_3$) is often used as an antacid. It neutralizes excess hydrochloric acid secreted by the stomach:

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(aq)$$

Milk of magnesia, which is an aqueous suspension of magnesium hydroxide, is also used as an antacid:

$$Mg(OH)_2(s) + 2HCl(aq) \rightarrow 2H_2O(l) + MgCl_2(aq)$$

Which is the more effective antacid per gram, $NaHCO_3$ or $Mg(OH)_2$? Justify your answer.

$Mg(OH)_2$

CALCULATIONS INVOLVING A LIMITING REACTANT

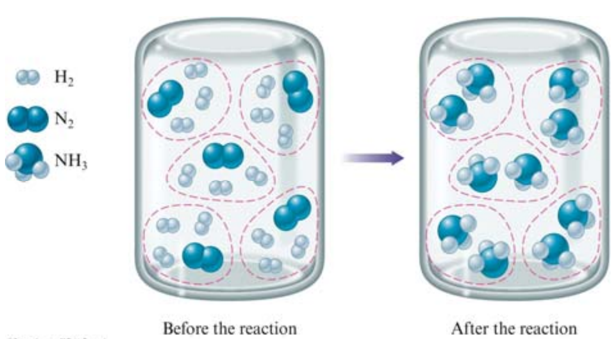
Ever notice how hot dogs are sold in packages of 10 while the buns come in packages of 8? What's up with that?! The bun is the limiting reactant and limits the hot dog production to 8 as well! The limiting reactant [or reagent] is the one consumed most entirely in the chemical reaction.

Let's use a famous process [meaning one the AP exam likes to ask questions about!], the Haber process. This reaction is essentially making ammonia for fertilizer production from the nitrogen in the air reacted with hydrogen gas. The hydrogen gas is obtained from the reaction of methane with water vapor. This process has saved millions from starvation!! The reaction is shown below.

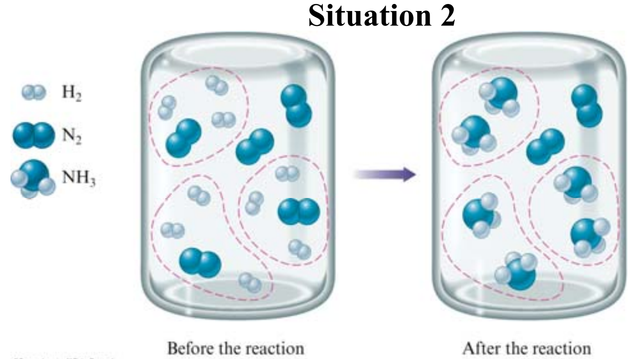
Exercise17
Examine the particle views and explain the differences between the two situations pictured below with regard to what is or is not reacting and total yield of ammonia.

$$N_2(g) \; + \; 3H_2(g) \; \rightarrow \; 2NH_3(g)$$

Situation 1



Situation 2



<u>Plan of attack</u>: 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.

Either way, I've established that hydrogen is the limiting reactant so I modify the table:

In English, that means I'll use up all the hydrogen but not all the nitrogen!

Molar Mass:	(28.02)	(2.02)	(17.04)
Balanced Eq'n	N_2 +	3 H_2 →	2 NH_3
mole:mole	**1**	**3**	**2**
# moles	∴825 mol used ~~892~~ moles	2,475 moles	∴1650 mol produced
amount	825 mol (28.02) = 23,116 g used ~~25,000~~ g 1,884 g excess!!	5,000 g	1650 mol (17.04) = 28,116 g produced

Here's the question again, let's clean up any sig.fig issues:
Suppose 25.0 kg of nitrogen reacts with 5.00 kg of hydrogen to form ammonia. (3 sig. fig. limit)

What mass of ammonia can be produced? 28,100 g produced = 28.1 kg
(It is always polite to respond in the unit given).

Which reactant is the limiting reactant? Hydrogen—once that's established, discard the nitrogen amounts and **let hydrogen be your guide**!

What is the mass of the reactant that is in excess? 1,884 g = 1.88 kg excess nitrogen!!

Exercise 18
Nitrogen gas can be prepared by passing gaseous ammonia over solid copper(II) oxide at high temperatures. The other products of the reaction are solid copper and water vapor. If a sample containing 18.1 g of NH_3 is reacted with 90.4 g of CuO, which is the limiting reactant? How many grams of N_2 will be formed?

CuO is limiting; 10.6 g N_2

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%