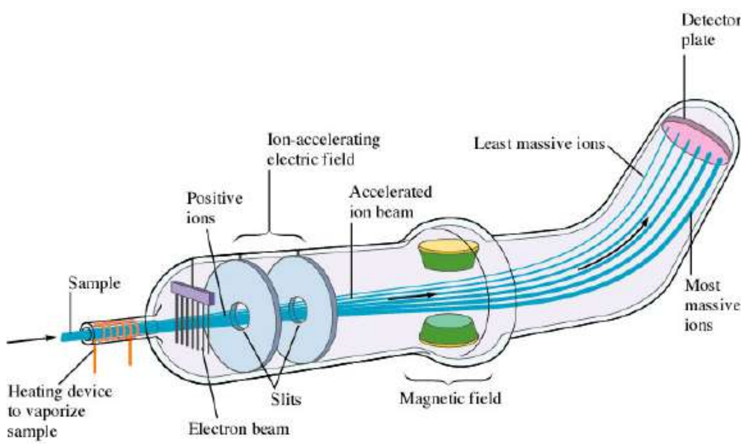
AP* Chemistry
STOICHIOMETRY

3.1 ATOMIC MASSES

- **^{12}C—Carbon 12**—In 1961 it was agreed that this would serve as the standard 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
  - o atoms or molecules are passed into a beam of high-speed electrons
  - o this knocks electrons OFF the atoms or molecules transforming them into cations
  - o apply an electric field
  - o this accelerates the cations since they are repelled from the (+) pole and attracted toward the (−) pole
  - o send the accelerated cations into a magnetic field
  - o an accelerated cation creates it's OWN magnetic field which perturbs the original magnetic field
  - o this perturbation changes the path of the cation
  - o the amount of deflection is proportional to the mass; heavy cations deflect little
  - o ions hit a detector plate where measurements can be obtained.



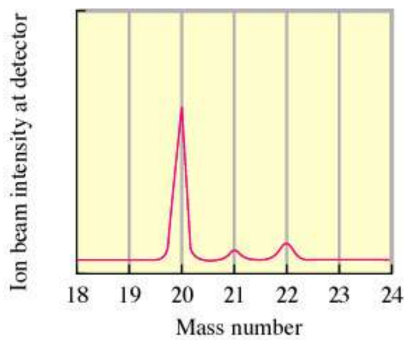
  - o $\frac{Mass^{13}C}{Mass^{12}C} = 1.0836129 \therefore Mass^{13}C = (1.0836129)(12amu) = 13.003355amu$

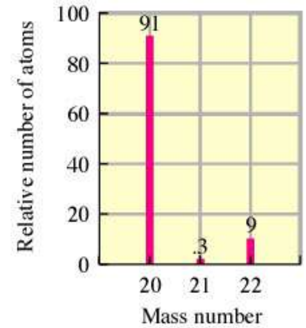
Exact by definition

- **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

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

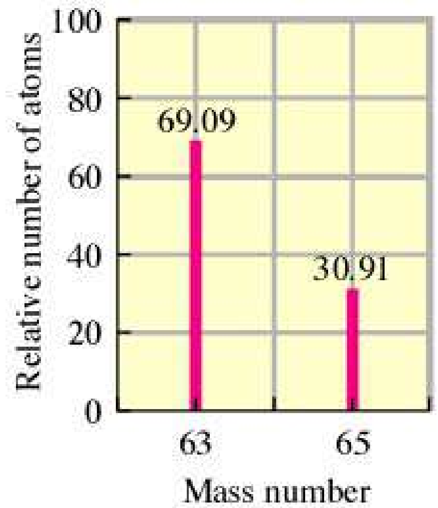
- **counting by mass**—when particles are small this is a matter of convenience. Just as you buy 5 lbs of sugar rather than a number of sugar crystals, or a pound of peanuts rather than counting the individual peanuts….this concept works very well if your know an *average* mass.
- **mass spectrometer to determine isotopic composition**—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 3.1 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 compute 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

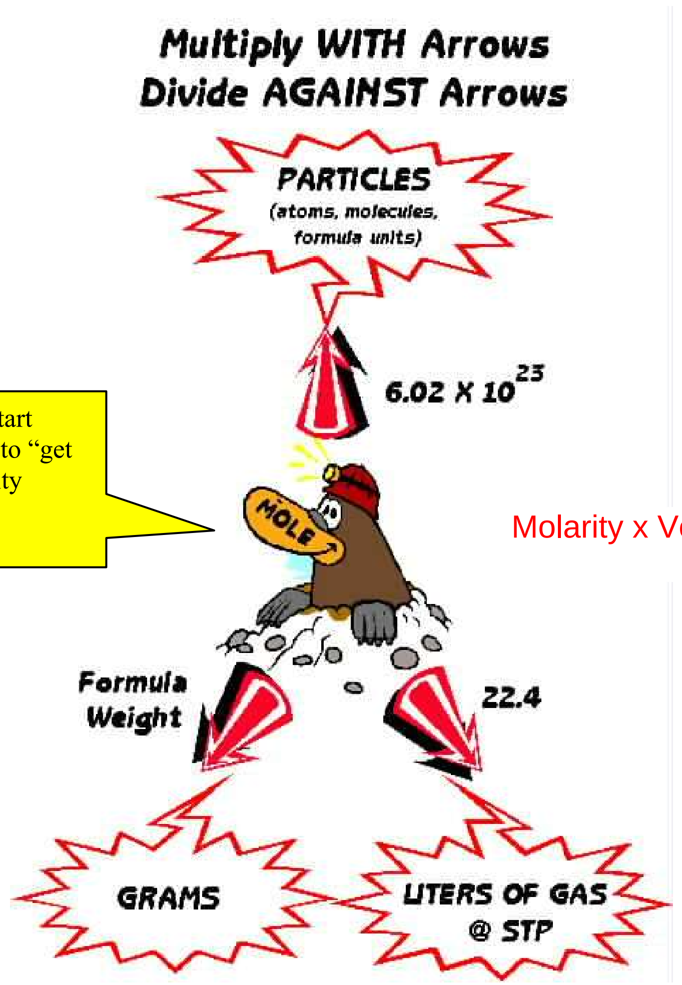
3.2 THE MOLE

- **mole**—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.
- **Avogadro's number**—6.02×10^{23}, the number of particles in a mole of anything

DIMENSIONAL ANALYSIS DISCLAIMER: Beginning on page 84 of the Chapter 3 text files you can find on this CD, you can find all of the remaining exercises worked out with dimensional analysis. This is most likely the way you were taught in Chemistry I. 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*!

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!

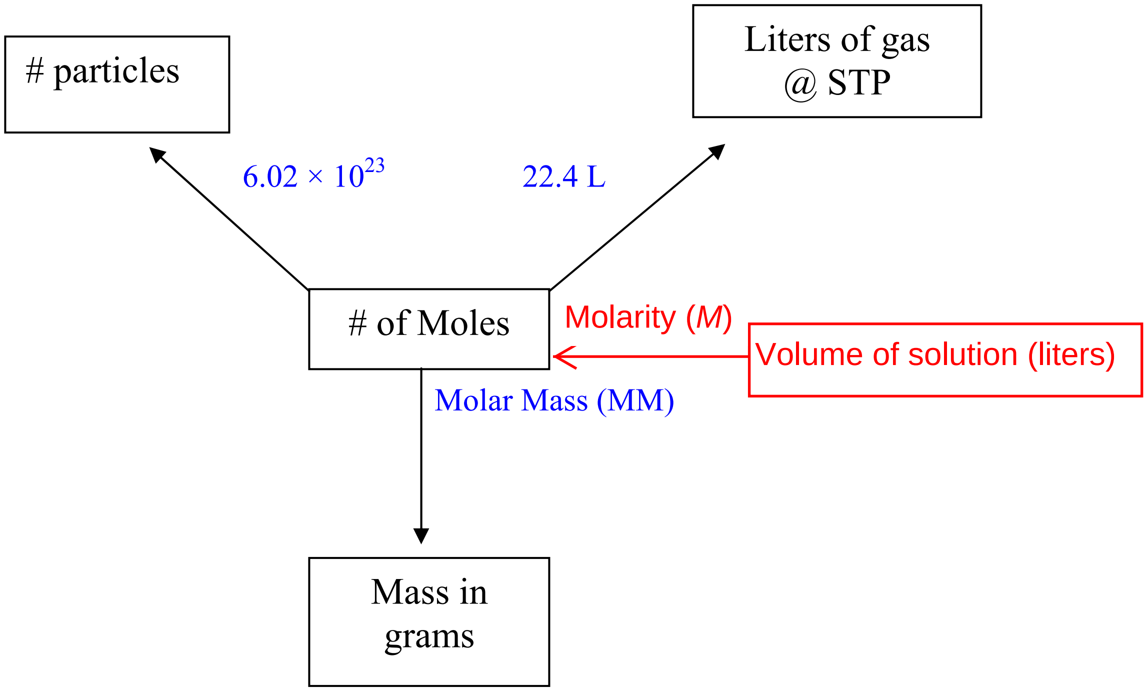
MULTIPLY [by the conversion factor on the arrow] when traveling IN THE DIRECTION OF THE ARROW and obviously, divide when "traveling" against an arrow.



IF you start with moles, start here, notice you multiply to "get away" to any other quantity FROM moles.

Molarity x Volume (L) = moles also!

When you draw this it will look more like this:



Exercise 3.2 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*. Compute the mass in grams of a sample of americium containing six atoms.
[There's a copy of the AP Chem periodic table at apchemistrynmsi.wikispaces.com PRINT—print it!]

2.42×10^{-21} g

Exercise 3.4 Determining Moles of Atoms

Aluminum (A1) is a metal with a high strength-to-mass ratio and a high resistance to corrosion; thus it is often used for structural purposes. Compute both the number of moles of atoms and the number of atoms in a 10.0-g sample of aluminum.

2.23×10^{23} atoms

Exercise 3.5 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

3.3 MOLAR MASS, MOLECULAR WEIGHT, AND FORMULA WEIGHT

- **molar mass, *MM***--the mass in grams of Avogadro's number of molecules; i.e. the mass of a mole!
- **molecular weight, MW**--sum of all the atomic weights of all the atoms in the formula (it is *essential you have a correct formula as you'll painfully discover!)*
- **empirical formula**--that ratio in the network for an ionic substance.
- **formula weight**--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}**. The AP Exam uses MM for all formula masses.**
- **A WORD ABOUT SIG. FIG.'s—**It is correct to "pull" from the periodic table as many sig. figs for your MM's as are in your problem—just stick with 2 decimal places for all—much simpler!

Exercise 3.6 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. How many moles of juglone does this sample represent?

a. 174.16 g/mol

b. 8.96×10^{-5} mol juglone

Exercise 3.7 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. What is the mass in grams of this sample? What is the mass of the CO_3^{2-} ions present?

a. 100.09 g/mol
b. 486 g; 292g CO_3^{2-}

Exercise 3.8 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. How many molecules of isopentyl acetate are released in a typical bee sting?

How many atoms of carbon are present?

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

ELEMENTS THAT EXIST AS MOLECULES

Pure hydrogen, nitrogen, oxygen and the halogens [I call them the "gens" collectively—easier to remember!] exist as **DIATOMIC** molecules *under normal conditions.* MEMORIZE!!! Be sure you compute their molar masses as diatomics. 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

3.4 PERCENT COMPOSITION OF COMPOUNDS

Two common ways of describing the composition of a compound: in terms of the number of its constituent atoms and in terms of the percentages (by mass) of its elements.

Percent (by mass) Composition: 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.*

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

Consider ethanol, C_2H_5OH

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

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

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

Mass of 1 mol of C_2H_5OH = 46.08 g/mol

NEXT THE MASS PERCENT CAN BE CALCULATED:

$$\text{Mass percent of C} = \frac{24.02\ g\ C}{46.08\ g/mol} \times 100\% = 52.13\%$$

Repeat for the H and O present.

Exercise 3.9 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. Compute the mass percent of each element in carvone.

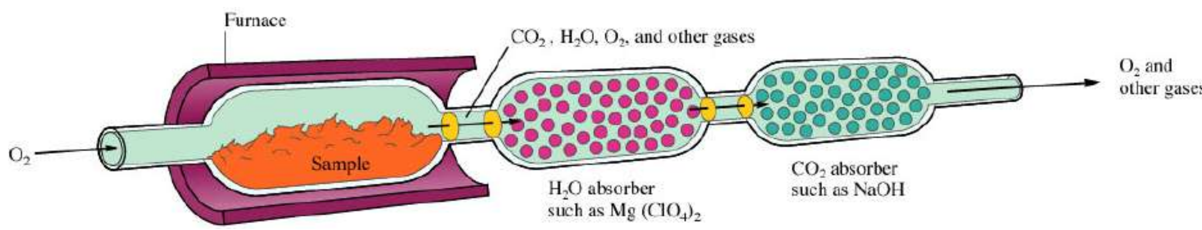
C = 79.94%
H = 9.41%
O = 10.65%

Exercise 3.10 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 $C_{14}H_{20}N_2SO_4$. Compute the mass percent of each element.

C = 53.82%
H = 6.466%
N = 8.969%
S = 10.27%
O = 20.49%

3.5 DETERMINING THE FORMULA OF A COMPOUND



When faced with a compound of "unknown" formula, one of the most common techniques is to combust it with oxygen to produce CO_2, H_2O, and N_2 which are then collected and weighed.

- **empirical and molecular formulas:** assume a 100 gram sample if given %
  - o empirical gives smallest ratio
  - need to know molar mass to establish molecular formula which is (empirical formula)$_n$, where n is an integer

- **determining empirical and molecular formulas**
  - **<u>hydrates</u>**—"dot waters" used to cement crystal structures.
  - **<u>anhydrous</u>**--*without* water

Example: A compound is composed of carbon, nitrogen and hydrogen. When 0.1156 g of this compound is reacted with oxygen [burned, combusted], 0.1638 g or carbon dioxide and 0.1676 g of water are collected. What is the empirical formula of the compound?

Compound + O_2 → CO_2 + H_2O + N_2 but NOT balanced!!

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

0.1638 g ÷ 44.01 g/mol = 0.003781 moles of CO_2 = 0.003781 moles of C

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

0.1676 ÷ 18.02 g/mol = 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—2 lungs per body, 2 atoms H in water and so on…] so DOUBLE THE NUMBER OF MOLES TO GET THE NUMBER OF MOLES OF HYDROGEN!! moles H = 0.01860 moles of H

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

g C = 0.003781 moles C × 12.01 = 0.04540 grams C

g H = 0.01860 moles H × 1.01 = 0.01879 grams H
0.06419 grams total thus far

SUBTRACT!

0.1156 g sample – 0.06419 g thus far = grams N left = 0.05141 g N so….

0.05141 g N ÷ 14.01 = 0.003670 moles N

Chemical formulas represent mole to mole ratios, so…divide the number of moles of each by the smallest # of moles of 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!!

Element	# moles	ALL Divided by 0.003670
C	0.003781	1
H	0.01860	5
N	0.003670	1

Therefore the correct EMPIRICAL formula is CH_5N.

Next, if we are told that the MM is 31.06 g/mol, then simply use this relationship:

(Empirical mass)	×	n	=	*MM*
(12.01 + 5.05 + 14.01)	×	n	=	31.06

Solve for n

n = 0.999678… or essentially one, so the empirical formula and the molecular formula are one in the same.

One last trick of the trade: When you don't know the mass of your sample, assume 100 grams so that any percents become grams….proceed by finding the number of moles!

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 3.11 Determining Empirical and Molecular Formulas I

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

71.65% Cl 24.27% C 4.07% H

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

Empirical formula = CH_2Cl

Molecular formula = $C_2H_4Cl_2$

Exercise 3.12 Determining Empirical and Molecular Formulas II

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 3.13 Determining a Molecular Formula

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 = $C_8H_{10}N_4O_2$

3.6 & 3.7 BALANCING CHEMICAL EQUATIONS

CHEMICAL REACTIONS

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].

CHEMICAL EQUATIONS

chemical reaction--transforms elements and compounds into new substances

balanced chemical equation--shows the relative amounts of reactants [left] and products [right] by molecule or by mole.

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

- *s, l, g, aq*--solid, liquid, gas, aqueous solution
- NO ENERGY or TIME is alluded to
- Antoine Lavoisier (1743-1794)--law of conservation of matter: *matter can neither be created nor destroyed*
  - ☞ this means "balancing equations" is all 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		*Products*
$CH_4(g) + 2O_2(g)$	⟶	$CO_2(g) + 2H_2O(g)$
1 molecule CH_4 + 2 molecules O_2	⟶	1 molecule CO_2 + 2 molecules H_2O
1 mol CH_4 molecules + 2 mol O_2 molecules	⟶	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	⟶	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	⟶	44 g CO_2 + 2(18 g) H_2O
80 g reactants	⟶	80 g products

Exercise 3.14 Balancing a Chemical Equation I

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, as shown in the two photographs on page 105. 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.

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

(4 × 2) H (4 × 2) H

Exercise 3.15 Balancing a Chemical Equation II

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.

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

3.8 STOICHIOMETRIC CALCULATIONS: AMOUNTS OF REACTANTS AND PRODUCTS

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.

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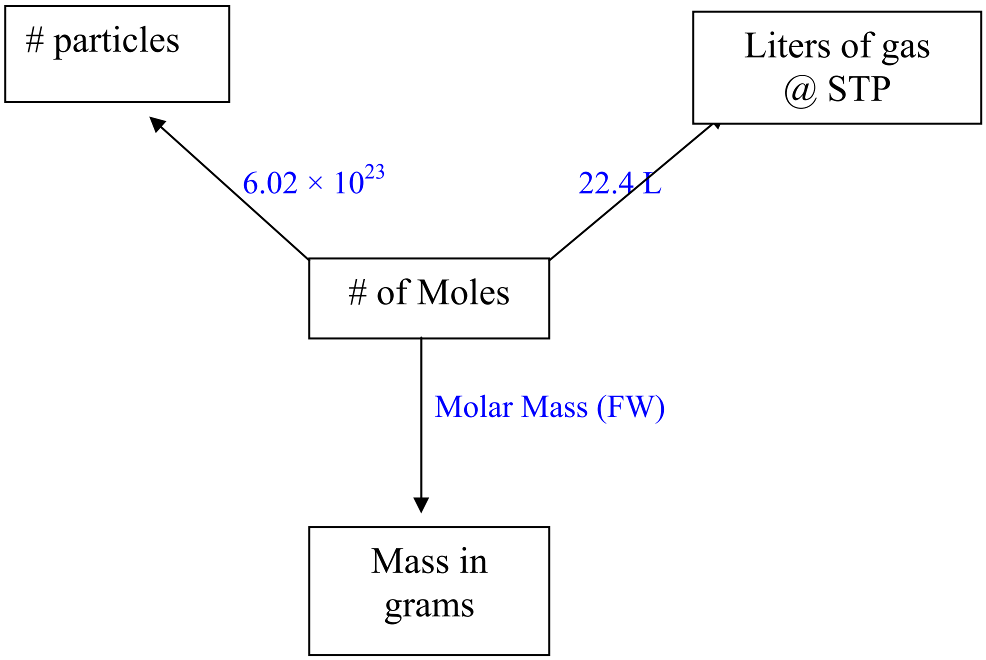
First you have to be proficient at the following no matter which method you choose!:

- 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:

<u>What mass of oxygen will react with 96.1 grams of propane?</u>

[notice all words—you supply 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 you're interested in.

3. Look at the coefficients on the balanced equation, they ARE the mole:mole ratios!

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 →	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. Find the number of moles of something, anything! Use the mole map. Start at 96.1 grams, divide [against the arrow] by molar mass to get 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 you find is not a "1", just divide to make it "1" and then it's easy greasy!] 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

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	

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

AND…how many water molecules are produced? Use the mole map , start in the middle with 8.71 mol (6.02×10^{23}) = 5.24×10^{24} molecules of water.

Try these two exercises with whichever method you like best!

Exercise 3.16 Chemical Stoichiometry I

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?

919 g

Exercise 3.17 Chemical Stoichiometry II

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

$$NaHCO_3(s) + HC1(aq) \rightarrow NaC1(aq) + H_2O(l) + CO_2(aq)$$

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

$$Mg(OH)_2(s) + 2HC1(aq) \rightarrow 2H_2O(l) + MgC1_2(aq)$$

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

$Mg(OH)_2$

3.9 CALCULATIONS INVOLVING A LIMITING REACTANT

Ever notice how hot dogs are sold in packages of 10 while the buns come in packages of 8?? 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.

Plan of attack: First, you'll know you ***need*** a plan if you are given TWO amounts of matter that react.

Next, when in doubt…find the number of moles. Set up your table like before, only now you'll have TWO amounts and thus TWO # 's of moles to get you started. I cover one up and "What if?" More to follow! It doesn't matter where you start the "What if?" game….you get there either way.

Let's use a famous process [meaning one the AP exam likes to ask questions about!], the Haber process. This is basically 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 below:

Molar Mass:	(28.04)		(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.

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

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? 23,100 g produced = 23.1 kg (always polite to respond in the unit given).

Which reactant is the limiting reactant? hydrogen—once that's established, chunk 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 3.18 Stoichiometry: Limiting Reactant

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

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

<u>Actual yield</u>: That which is realistic!

<u>Percent yield:</u> The ratio of actual to theoretical yield.

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

Exercise 3.19 Calculating Percent Yield

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 6.86×10^4 g
Percent yield is 52.3%