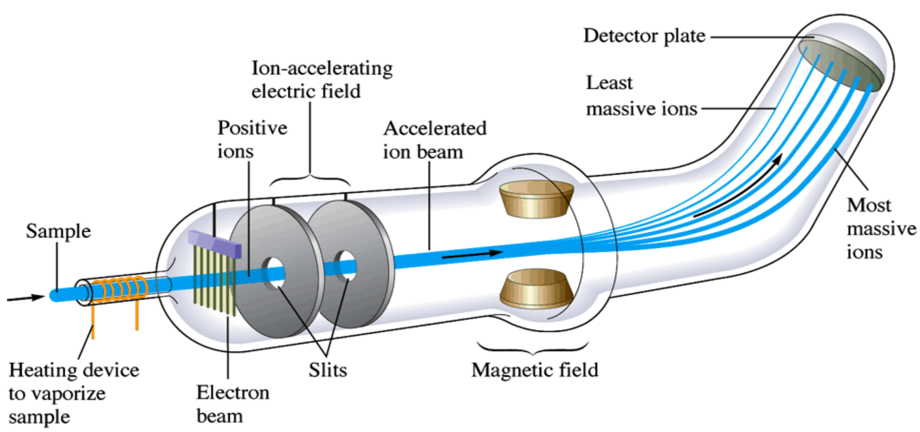
**Name ________________________ AP CHEM ___/___/___**

# Chapter 3 Outline
# Stoichiometry

### Atomic Masses

- The modern system of atomic masses, instituted in 1961, is based on carbon-12.
- Carbon-12 is assigned the mass of exactly 12 atomic mass units (amu)
- The most accurate method currently available for comparing the masses of atoms involves the use of the mass spectrometer. **(LO 1.14)**
- Atoms or molecules are passed through a beam of high speed electrons. This knocks electrons off the substance being analyzed and changes it to a positive ion. The electric field accelerates the positive ions into a magnetic field. Since an accelerated ion creates its own magnetic field, an interaction with the applied magnetic field causes a change in the path of the ion. The amount of deflection is a function of a substance's mass. The most massive ions are deflected the smallest amount, which causes the ions to separate. A comparison of the positions where the ions hit the detector plate gives very accurate values of their relative masses.
- For example: When isotopes carbon-12 and carbon-13 are analyzed in a mass spectrometer, the ratio of their masses is found to be: 1:1.0836129 To calculate the mass of carbon-13: (1.0836129)(12 amu) = 13.003355 amu.




- Since the modern system is based on carbon-12, it may seem surprising that carbon's atomic weight is 12.01 instead of 12.0. This is due to the fact that there are three carbon isotopes($^{12}C$, $^{13}C$ & $^{14}C$) and the atomic weight listed on the periodic table is an average value representing the mass and abundance of each isotope.

### Average Atomic Weight

- The **atomic weight of an element is the weighted average of the masses of the isotopes of that element**. The weighted average is determined using the abundance and mass of each isotope. Most elements have more than one naturally occurring isotope.
- For example, there are two naturally occurring isotopes of copper, ***copper-63*** (62.93 amu) and ***copper-65***. (64.93 amu). The natural abundances of the isotopes are **69.17%** and **30.83%** respectively.
- **To determine the atomic weight:**
  **Step 1:** Multiply the mass number and the relative abundance (as a decimal). The mass of the electron is insignificant in this calculation and is not used.

| isotope name | atomic mass | x | abundance (as a decimal) | = | result |
|---|---|---|---|---|---|
| copper-63 | 62.93 amu | x | 0.6917 | = | 43.53 |
| copper-65 | 64.93 amu | x | 0.3083 | = | 20.02 |

- **Step 2:** Add up your results. **Atomic Weight = 63.55 amu**

- The mass of a proton is approx. 1.008 amu. The mass of a neutron is approximately 1.009 amu. The mass of the electron is insignificant. When elements are formed from the individual subatomic particles a large amount of energy in the form of heat is released. This loss of mass is called the **mass defect** and this conversion of mass into energy is seen in the equation: $E = mc^2$.

**Gram Formula Mass / Molar Mass**

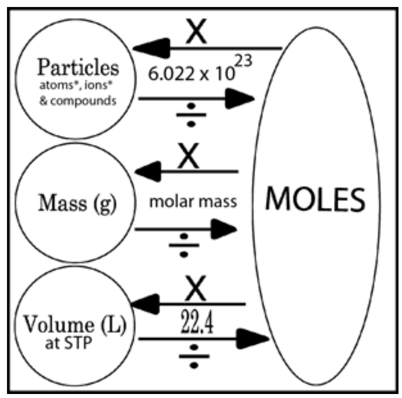
- Gram formula mass (also known as molar mass) is defined as the atomic mass of one mole of an element, molecular compound or ionic compound. The answer must always be written with the unit g/mol (grams per mole).

**Calculating a substance's gram formula mass (molar mass)**:
Calculate the gram formula mass(molar mass) of **ammonium phosphate**.

| Description of Action | Action |
|---|---|
| **1.** Write the formula for the compound. | **1.** $(NH_4)_3PO_4$ |
| **2.** Determine how many of each atom are in the compound. If there is a number outside the parenthesis, multiply each subscript by this number. If there is no subscript, assume it is one. | **2.** N: 3<br>H:12<br>P: 1<br>O:4 |
| **3.** Multiply the number of atoms by its atomic weight**.** **Round the atomic weight to the tenth's place.** | **3.** N: 3 x 14.0 = 42.0<br>H:12 x 1.0 = 12.0<br>P: 1 x 31.0 = 31.0 (31.0 is rounded from 30.97)<br>O:4 x 16.0 = 64.0 (16.0 is rounded from 15.99) |
| **4.** Add your results and use the unit **g/mol** (grams per mole) on the end. | **4.** 42.0 + 12.0 + 31.0 + 64.0 = **149.0 g/mol.** |

**Mole Conversions (LO 1.4)**

- To convert between moles, atoms, molecules and liters you should use the following chart:
- This topic and chart are covered extensively in first year chemistry. For more practice/information check out the link to my Honors/Chemistry I worksheet: http://www.sartep.com/chem/worksheets/pdfs/7_04.pdf
- One example is shown below:




Calculate the number of hydroxide ions in 23.5 grams of aluminum hydroxide.
23.5 g $Al(OH)_3$ ÷ 78.0 g $Al(OH)_3$ x 6.022 x $10^{23}$ x 3 = **5.44 x $10^{23}$ $(OH)^-$**

**Percent Composition (LO 1.3)**

- When we calculate percent composition, we are determining the relative mass that each element contributes to the total mass of the compound. For example, if we calculate the gram formula mass of $H_2O$ we will find it to be 18.0 g/mol. We arrived at this number by adding the mass of oxygen, 16.0, and the mass of two hydrogens, 2.0. Oxygen makes up 16.0 of the total 18.0 grams. Hydrogen is 2.0 of 18.0 grams. If we divide each elements total mass by the compounds total mass and multiplying this result by 100, we get a percentage. This percentage is the element's percent composition.

**Example (from above):**

| **STEP 1** (find g.f.m.) | **STEP 2** (divide each element mass by g.f.m.) | **STEP 3** (multiply by 100) |
|---|---|---|
| $H_2O$: | | |
| H: 2 x 1.0 = 2.0 → | H: 2.0 ÷ 18.0 = 0.111 → | H: 0.111 x 100. = **11.1%** |
| O: 1 x 16.0 = 16.0 | O: 16.0 ÷ 18.0 = 0.889 (rounded) | O: 0.889 x 100. = **88.9%** |
| **Total = 18.0 g/mol** | | |

**Example:** Calculate the percent composition of ammonium nitrate.

| Description of Action | Action |
|---|---|
| **1.** Write the formula for the given compound. | **1.** $NH_4NO_3$ |
| **2.** Record the amount of each element in the compound. (Note: We have 2 total nitrogen, so we record them together.) Multiply the amount of each element by its atomic weight (measured to the tenths place). Add the results to find the gram formula mass of the compound. | **2.** N: 2 x 14.0 = **28.0**<br>H: 4 x 1.0 = **4.0**<br>O: 3 x 16.0 = **<u>48.0</u>**<br>**80.0 g/mol** |
| **3.** Divide the total mass of each element by the gram formula mass. For these calculations your answer should have 3 places after the decimal (round if necessary). | **3.** N: 28.0 ÷ 80.0 = **0.350**<br>H: 4.0 ÷ 80.0 = **0.050**<br>O: 48.0 ÷ 80.0 = **0.600** |
| **4.** Multiply each result by 100. Add the % symbol to your new result. (If you were to add up your percentages the must equal 100.) | **4.** N: 0.350 x 100 = **35.0%**<br>H: 0.050 x 100 = **5.0%**<br>O: 0.600 x 100 = **<u>60.0%</u>**<br>**100.0%** |

**Determining Empirical and Molecular Formula (LO 1.2)**

**Empirical & Molecular Formula**

**Empirical Formula**

**Given: Percentage or mass of each element or compound.**

1. Divide each percentage or mass given by the molar mass of the element or compound.
2. Divide each result by the smallest result.
3. Round and multiply (if necessary) each result by the SAME whole number to get a whole number result. (This step is not necessary for hydrates). Use the following chart:

| Round to: | Multiply by: |
|---|---|
| x.00 | (use the other factor) |
| x.20 | 5 |
| x.25 | 4 |
| x.33 | 3 |
| x.50 | 2 |
| x.66 | 3 |
| x.75 | 4 |
| x.80 | 5 |

4. Write the formula using your results.

**Molecular Formula** (do all of the above)

5. Find the gram formula mass of the empirical formula.
6. Divide the molecular formula mass (given) by the empirical formula mass (calculated).
7. Multiply each subscript by the result. (The result MUST be a whole number.)

**Example:** Ascorbic acid, also known as vitamin C, has a percentage composition of 40.9% carbon, 4.58% hydrogen, and 54.5% oxygen. Its molecular mass is 176.1 g/mol. What is its molecular formula?

**C: 40.9 ÷ 12.0 = 3.41 ÷ 3.41 = 1.00 x 3 = 3**
**H: 4.58 ÷ 1.0 = 4.58 ÷ 3.41 = 1.32 x 3 = 4**
**O: 54.5 ÷ 16.0 = 3.41 ÷ 3.41 = 1.00 x 3 = 3**
**$C_3H_4O_3$ (empirical formula)** →

**C: 3 x 12.0 = 36.0**
**H: 4 x 1.0 = 4.0**
**O: 3 x 16.0 = <u>48.0</u>**
**88.0 g/mol** →

**176.1 ÷ 88.0 = 2**

$\boxed{C_6H_8O_6}$

A favorite of mine and one that sometimes comes up on the AP exam is combustion analysis. To solve a combustion analysis problem one must know that all of the carbon in a compound is converted to carbon dioxide and all of the hydrogen in a compound is converted to water. Usually the percent analysis is given for any additional element, except oxygen. To find the mass of oxygen you must subtract the masses of the other elements from the total. Once you know the mass of each element you can find the empirical and molecular formulas. Try the one below; the answers are given but the work is not.

Answer the following questions that relate to the analysis of chemical compounds.
A compound containing the elements C, H, N, and O is analyzed. When a 1.2359 g sample is burned in excess oxygen, 2.241 g of $CO_2(g)$ is formed. The combustion analysis also showed that the sample contained 0.0648 g of H.

(i) Determine the mass, in grams, of C in the 1.2359 g sample of the compound. **0.6116 g C**

(ii) When the compound is analyzed for N content only, the mass percent of N is found to be 28.84 percent. Determine the mass, in grams, of N in the original 1.2359 g sample of the compound. **0.3564 g N**

(iii) Determine the mass, in grams, of O in the original 1.2359 g sample of the compound. **0.2031 g O**

(iv) Determine the empirical formula of the compound. **$C_4H_5N_2O$**

(v) The molecular mass of the compound is 194.2 g/mol. Determine the molecular formula of the compound. **$C_8H_{10}N_4O_2$**

**Chemical Equations**

- Chemical changes are represented by chemical reactions.
- In a chemical reaction, the matter is not created or destroyed (Law of Conservation of Mass). Instead the atoms in the chemical reaction are rearranged. A balanced equation has equal numbers of each type of atom on each side of the equation.

coefficient coefficient

$\mathbf{2}\,H_2 + O_2 \rightarrow \mathbf{2}\,H_2O$

Reactants yield Products

**Coefficients** - the numbers written in front of the chemical formulas used in balancing chemical equations.
**Reactants** - the elements or compounds on the left hand side of the equation before the → are known as the reactants.
**Products** - the elements or compounds on the right side of the equation after the → are known as the products.
**Yield(→) -** this arrow is called a yield sign. It separates the reactants from the products.

| ***Diatomic Elements*** |
|---|
| $H_2$ $N_2$ $O_2$ |
| $F_2$ $Cl_2$ $Br_2$ $I_2$ |

The word **diatomic** is used above when we read the equation. A **diatomic molecule** is made up of **two atoms of the same element**. These elements never exist as one naturally. When we are writing chemical equations you must always remember to write a subscript of two when these elements are by themselves. (When they are bonded to other elements they **do not** have to have a subscript of 2.) There are seven diatomic elements; they are listed in the chart to the right. You must MEMORIZE these seven diatomic elements. You can see that they kind of form a pattern on the periodic table. Many pneumonic exist. Ask me if you need one.

**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 + 2 molecules | ⟶ | 1 molecule + 2 molecules |
| 1 mole + 2 moles | ⟶ | 1 mole + 2 moles |
| $6.022 \times 10^{23}$ molecules + 2 ($6.022 \times 10^{23}$ molecules) | ⟶ | $6.022 \times 10^{23}$ molecules + 2 ($6.022 \times 10^{23}$ molecules) |
| 16 g + 2 (32 g) | ⟶ | 44 g + 2 (18 g) |
| 80 g reactants | ⟶ | 80 g products |

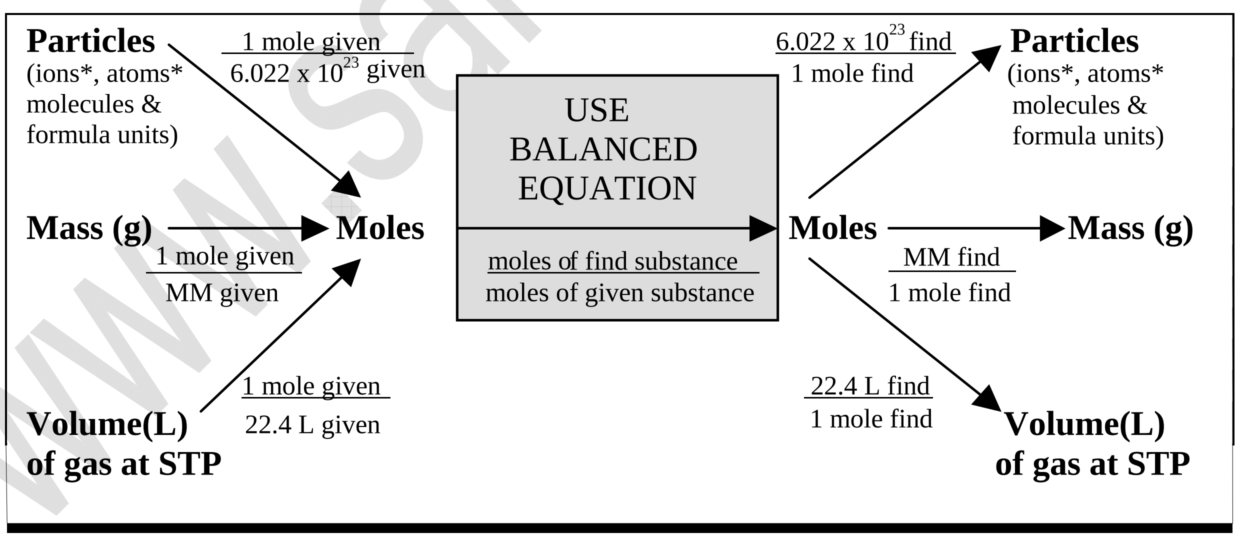
**Balancing Chemical Equations**

- There are three rules to follow when balancing chemical equations. They do not work every time, but they do work most times and most people find them helpful.
- If there is an element that is not bonded to any other atom or is diatomic, balance it last. For example in the equation: $CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$, oxygen is diatomic. When balancing this equation, balance oxygen last.
- If you are balancing an equation with both hydrogen and oxygen and neither is a diatomic molecule, balance hydrogen last and balance oxygen second to last. For example in the equation: $CO_2 + NH_3$ --> $OC(NH_2)_2 + H_2O$, both hydrogen and oxygen are used and neither exist as a diatomic molecule. When balancing this equation, balance hydrogen last and oxygen second to last.
- If you come to the point where you have an odd number of a certain element on one side of the equation and an even number of the same element on the opposite side on the equation, double the coefficient of the formula with the odd number of the element. If the coefficient is 1, change it to 2.
- Ultimately it is a trial and error process. Keep at it until it works. Here are 10 practice problems. The answers are on the last page. Don't cheat.

1. $C_{10}H_{16} + Cl_2 \rightarrow C + HCl$
2. $CO_2 + NH_3 \rightarrow OC(NH_2)_2 + H_2O$
3. $Si_2H_3 + O_2 \rightarrow SiO_2 + H_2O$
4. $Al(OH)_3 + H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2O$
5. $Fe + O_2 \rightarrow Fe_2O_3$
6. $Fe_2(SO_4)_3 + KOH \rightarrow K_2SO_4 + Fe(OH)_3$
7. $C_7H_6O_2 + O_2 \rightarrow CO_2 + H_2O$
8. $H_2SO_4 + HI \rightarrow H_2S + I_2 + H_2O$

**Stoichiometry (LO 3.3, 3.4)**

- I use the chart to the right to teach stoichiometry.




**START on THIS SIDE** **FINISH on THIS SIDE**

In a stoichiometry problem you always begin with a value of one of the units on the left and you always finish with a value of one of the units on the right. In every stoichiometry problem you must make a mole relationship comparison between substances being compared using their coefficients from the balanced chemical equation. This concept is used extensively. If you need help, e-mail me or see me after school for help. Additional problems can be found here: http://www.sartep.com/chem/worksheets/pdfs/8_02.pdf. E-mail me for the answers.

**Limiting Reactant (LO 3.4)**

- In a limiting reactant problem you will be given amounts of two of your reactants. The goal is to determine which you will run out of first. The only way to do this is to solve the stoichiometry problem using each reactant and its value. The reactant that produces **less** of the same product is the one you would run out of first, and thus your **limiting reactant**. The other reactant, the one you will have extra of when finished, is your **excess reactant**.

**Percent Yield**

- The percent yield equation is used to determine the efficiency of a chemical procedure. When you perform a percent yield calculation you will be comparing how much product you **actually** produce in a lab situation with how much should theoretically be produced if the situation was ideal. The theoretical yield is what you determine in a stoichiometry problem. The percent yield equation is listed below.

$$\textbf{Percent Yield} = \frac{\textbf{Actual Yield}}{\textbf{Theoretical Yield}} \times \textbf{100}$$

**An example showing all three topics and the worked out solutions are included below.**

.
$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

Zinc reacts with hydrochloric acid to form zinc chloride and hydrogen gas. 98.2 grams of zinc and 98.2 grams of hydrogen chloride react?

(a) Identify the limiting reactant. Support your answer with calculations. (3 points)
(b) How much of the excess reagent remains? (3 points)
(c) Calculate the volume of hydrogen gas produced. (3 points)
(d) If 12.5 liters of hydrogen gas are actually produced, what is the percent yield? (3 points)

(a) $\frac{98.2 \text{ g Zn}}{1} \times \frac{1 \text{ mole Zn}}{65.4 \text{ g Zn}} \times \frac{1 \text{ mole } H_2}{1 \text{ mole Zn}} = \mathbf{1.50 \text{ mol } H_2}$

$\frac{98.2 \text{ g HCl}}{1} \times \frac{1 \text{ mole HCl}}{36.5 \text{ g HCl}} \times \frac{1 \text{ mole } H_2}{2 \text{ mole HCl}} = \mathbf{1.34 \text{mol } H_2}$

Limiting Reagent: **HCl**

(b) $\frac{98.2 \text{ g HCl}}{1} \times \frac{1 \text{ mole HCl}}{36.5 \text{ g HCl}} \times \frac{1 \text{ mole Zn}}{2 \text{ mole HCl}} \times \frac{65.4 \text{ g Zn}}{1 \text{ mole Zn}} = \mathbf{88.0 \text{ g Zn needed}}$

98.2 – 88.0 = **10.2 g of excess Zn**

(c) $\frac{98.2 \text{ g HCl}}{1} \times \frac{1 \text{ mole HCl}}{36.5 \text{ g HCl}} \times \frac{1 \text{ mole } H_2}{2 \text{ moles HCl}} \times \frac{22.4 \text{ L } H_2}{1 \text{ mole } NH_2} = \mathbf{30.1 \text{ L } H_2}$

(d) Percent Yield = $\frac{12.5}{30.1} \times 100 = \mathbf{41.5\%}$

**Balancing Answers:**

1. $C_{10}H_{16}$ + **8** $Cl_2$ → **10** C + **16** HCl
2. $CO_2$ + **2** $NH_3$ → $OC(NH_2)_2$ + $H_2O$
3. **4** $Si_2H_3$ + **11** $O_2$ → **8** $SiO_2$ + **6** $H_2O$
4. **2** $Al(OH)_3$ + **3** $H_2SO_4$ → $Al_2(SO_4)_3$ + **6** $H_2O$
5. **4** Fe + **3** $O_2$ → **2** $Fe_2O_3$
6. $Fe_2(SO_4)_3$ + **6** KOH → **3** $K_2SO_4$ + **2** $Fe(OH)_3$
7. **2** $C_7H_6O_2$ + **15** $O_2$ → **14** $CO_2$ + **6** $H_2O$
8. $H_2SO_4$ + **8** HI → $H_2S$ + **4** $I_2$ + **4** $H_2O$