Chapter 3: Mass Relationships in Chemical Reactions – Khan Academy videos to watch: *The Mole; Molecular and Empirical Formulas; Formula from Mass Composition; Another Mass Composition Problem; Molecular and Empirical Formula from Percent Composition; Balancing Chemical Reaction; Stoichiometry; Stoichiometry Limiting Reactant.* Suggested homework problems pgs. 107 - 114 SET 1: 5, 6, 12, 14; SET 2: 20, 22, 28; SET 3: 40, 42, 44, 50a, 54; SET 4: 60; SET 5: 76,84, 104; SET 6: 107, 108, 117.

I. The Mole

A The mole is a unit of measure that measures quantity. It is the same concept as a dozen. You can have a dozen doughnuts, oranges, golf balls, etc. and you always have a quantity of 12. The same goes for a mole, we can have a mole of doughnuts, oranges or golf balls and would always have 6.02×10^{23}. This number is so large since we use it to measure things very small like atoms and molecules. 6.02×10^{23} is called Avogadro's number.

B Determining Avogadro's number

1. The number was determined by figuring out how many atoms are present in 12.000g of carbon-12. 12.000g was used because it is the same number as the atomic weight, which is measured in amus and is the mass of a single atom of carbon-12.

2. $$12.000\cancel{g}\ {}^{12}_{6}C \times \frac{1\ \cancel{amu}}{10^{-24}\cancel{g}} \times \frac{1\ \text{atom}\ {}^{12}_{6}C}{12.000\cancel{amu}} = 6.02 \times 10^{23}\ \text{atoms}$$

a. This many atoms is called a mole, from the latin *moles* meaning heaps, and the mass of this many atoms is called the molar mass. All molar masses should be rounded to the nearest 0.1.

b. Examples:

Molar mass of copper is 63.5gCu / 1 mole Cu (see periodic table)

Molar mass of sodium chloride NaCl

1 Na x 23.0g = 23.0g

1 Cl x 35.5g = <u>35.5g</u>

58.5g NaCl / 1 mole NaCl

Practice problems:

What's the mass of 8.93×10^{36} molecules of ammonia, NH_3 ?

How many atoms are present in 15 g of silver nitrate, $AgNO_3$?

C Molar Volume

1. Avogadro's Principle - Equal volume of gases at the same temperature and pressure contain the same number of particles. Therefore, 1 mole of any gas at standard temperature and pressure ($0\,^{O}C$ and 1 atmosphere or 101.3 kPa) has a volume of 22.4L. This give the unitary rate of

$$\frac{22.4 \text{ L any gas}}{1 \text{ mole any gas.}}$$

Practice problems – What is the mass of 4.28 L of Methane gas (CH_4)? How many molecules are present? How many atoms are present?

II. Percent Compostion of compounds

A Divide the individual element molar mass by the entire formula molar mass and multiply by 100% compound.

1. Practice first with % compostion of an ice cream sundae (worksheet)
2. Determine % compostion of calcium carbonate, $CaCO_3$

1 Ca x 40.1 g = 40.1g
1 C x 12.0 g = 12.0g
3 O x 16.0 g = 48.0g
100.1g $CaCO_3$ / 1 mole $CaCO_3$

$$\frac{40.1\text{g Ca}}{100.1\text{g } \cancel{CaCO_3}} \times 100\ \%\ \cancel{CaCO_3} = 40.1\ \%\ \text{Ca}$$

3. All others are done the same way.

III. Empirical Formulas

A Simplest whole number ratio of atoms in a formula. This is always the ionic compound formula but not always the molecular formula or true formula.

B Determine the empirical formula by determining the simplest mole ratio.

1. Recipe for empirical formula
   a. Change % to grams by assuming 100 grams of the compound
   b. Change grams to moles using molar mass
   c. Divide each by the smallest number of moles to get a simple whole number ratio.
   d. If 0.1 place is 0,1,2,8,9 then round to the closest whole number;
   e. If 0.1 place is 5 then multiply entire ratio by 2 ;
   f. If 0.1 place is 3,7 then multiply entire ratio by 3;
   g. If 0.1 place is 4,6 then something's wrong, unless lab data

Example : A compound is placed into a mass spectrometer and is found to contain 88.8 % copper and 11.2 % oxygen determine the empirical formula

Assume 100 grams of the compound gives 88.8g Cu and 11.2g O.

$$88.8 \text{ g Cu} \times \frac{1\text{mole Cu}}{63.5\text{g Cu}} = 1.40 \text{ moles Cu}$$

$$11.2 \text{ g O} \times \frac{1\text{mole O}}{16.0\text{g O}} = 0.700 \text{ mole O}$$

$$\frac{1.40 \text{ mole Cu}}{0.700\text{mole O}} = 2 \text{ Cu}$$

$$\frac{0.700 \text{ mole O}}{0.700 \text{ mole O}} = 1 \text{ O}$$

Therefore, thus and ergo the empirical formula is Cu_2O and the name is copper (I) oxide.

Example 2: Your lab data reveals the following composition of a compound by mass; 13.5 g calcium, 10.8g oxygen, 0.675g hydrogen. What is the empirical formula of this compound?

IV. Molecular (True) Formulas

A If determining molecular formulas, remember the molecular formula is a multiple of the empirical formula thus the molecular formula mass is a multiple of the empirical formula mass.

1.(empirical formula mass)(multiplier x) = Molecular formula mass

Solve for multiplier then multiply the empirical formula by this number.

a. Example :

Ribose has a molar mass of 150g / mole and the composition of 40.0% C; 6.67% H; and 53.3% O. Determine the molecular formula.

V. Chemical Reactions and Equations

A Definition - A process in which one or more substances are converted into a new substance with different physical and chemical properties.

B Why do things react? Substances react to become more stable and fulfill octect rule; gain a more stable crystal shape; go towards lower energy and higher entropy.

C Evidence that a chemical reaction has occurred

1. dramatic color change
2. gas is evolved
3. water is formed
4. a precipitate is formed (a solid forming from two solutions)
5. heat and / or light is emitted
6. obvious new product with new properties

D General Form

1. Reactants ===> Products
2. word equation is a short description of what occurred.
   a. ex. Gaseous hydrogen reacts with gaseous oxygen to form water vapor
   b. hydrogen gas + oxygen gas ===> water vapor (word equation)
   c. chemical reaction in shorthand form with all symbols
   d. $H_{2(g)} + O_{2(g)}$ ===> $H_2O_{(v)}$ not balanced yet
      i. (g) = gas (l) = liquid
      ii. (s) = solid (aq) = aqueous (water solution)
      iii. (v) = vapor (gaseous state of something normally a solid or liquid at room temp)
   e. Chemical reactions must be balanced to conserve matter; can't create or destroy matter. Must have the same number of atoms on reactants side as on product side.
      i. $2H_{2(g)} + O_{2(g)}$ ===> $2H_2O_{(v)}$
   f. Information from a **B**alanced **C**hemical **E**quation (BCE)
      i. 2 molecules H_2 + 1 molecule O_2 ==> 2 molecules H_2O
      ii. 4 atoms H + 2 atoms O ==> 6 atoms of H & O
      iii. 4 g H_2 + 32 g O_2 ==> 36 g H_2O
      iv. 2 moles H_2 + 2 moles O_2 ==> 2 moles H_2O

E Types of chemical reactions

1. Compostion , Combination or Synthesis
   a. A + B ==> AB
   b. $2H_{2(g)} + O_{2(g)}$ ===> $2H_2O_{(v)}$
   c. only one product
   d. Reactants are element phase and product is usually a solid
2. Decompostion
   a. AB ==> A + B
   b. only one reactant
   c. **Metallic carbonates** decompose, when heated, into metallic oxides and carbon dioxide.

i. ex. - $CaCO_{3(s)}$ $\xrightarrow{\Delta}$ $CaO_{(s)} + CO_{2(g)}$

d. **Metallic chlorates** decompose, when heated, into meatallic chlorides and oxygen.

i. ex. - $Mg(ClO_3)_{2(s)}$ $\xrightarrow{\Delta}$ $MgCl_{2(s)} + 3O_{2(g)}$

e. **Metallic hydoxides** decompose, when heated, into metallic oxides and water.

i. ex. - $2Fe(OH)_{3(s)}$ $\xrightarrow{\Delta}$ $Fe_2O_{3(s)} + 3H_2O_{(v)}$

f. Some metallic oxides decompose when heated.

i. ex. - $2SnO_{2(s)}$ $\xrightarrow{\Delta}$ $2SnO_{(s)} + O_{2(g)}$

g. Binary ionic compounds decompose, with electricity, into corresponding elements.

i. ex. - $2NaCl_{(l)}$ $\xrightarrow{\Delta}$ $2Na_{(s)} + Cl_{2(g)}$

h. Reactants are solid, products are solid + gas or water.

3.Single Replacement

a. A + BX ===> AX + B metals or (+) replace metals or (+)

b. Y + BX ===> BY + X nonmetals (-) replace nonmetals (-)

c. In order for A,Y to replace B,X , respectively, A,Y must be more active than B,X. Check the Activity Series. Halogens see Periodic Table.

i. ex. - $Cu_{(s)} + 2AgNO_{3(aq)}$ ===> $2Ag_{(s)} + Cu(NO_3)_{2(aq)}$

d. use rule of thumbs to determine Cu^{+1} or Cu^{+2}

i. $Cu_{(s)} + H_2O_{(l)}$ ===> No RXN

e. $2NaCl_{(aq)} + F_{2(g)}$ ===> $2NaF_{(aq)} + Cl_{2(g)}$

f. compounds are always aqueous unless water; elements are elemental phase.

4.Double Replacement

a. AX + BY ===> BX + AY

b. In order for this to occur, one of three must be produced :

i. H_2O in form of hydrogen hydroxide;

ii. gas (HCl; CO_2; H_2S; SO_2) ;

iii. precipitate must form, i.e. a solid from two solutions. To determine precipitate check solubility table.

iv. ex. - $Pb(NO_3)_{2(aq)} + K_2CrO_{4(aq)}$ ==> $2KNO_{3(aq)} + PbCrO_{4(s)}\downarrow$

c. Both reactants are aqueous, one product is aqueous the other will be a solid, gas or liquid water.

5.Complete Combustion of a Hydrocarbon (anything that has H or C)

a. $C_xH_y + O_{2(g)}$ ===> $CO_{2(g)} + H_2O_{(v)}$

b. Always add oxygen and produce $CO_{2(g)} + H_2O_{(v)}$

c. ex. - propane, $C_3H_{8(g)} + O_{2(g)}$ ==> $3CO_{2(g)} + 4H_2O_{(v)}$

d. The hard part is balancing. Try the complete combustion of C_4H_{10} .

VI. Stoichiometry – mass and mole relationships in chemical reactions

A Four step recipe

1. Write the balanced chemical equation (BCE)
2. Change grams given to moles given using molar mass from periodic table
3. Change moles given to moles needed using the mole ratio from BCE
4. Change moles needed to grams needed using molar mass from periodic table

B Try practice problem 3.13 pg. 100 and problem 72 pg. 110.

C Limiting reactant

1. the reactant that controls the amount of reaction occurring by being completely consumed in the reaction is the limiting reactant
2. the reactant not completely consumed is said to be the excess reactant
3. Try practice problems 3.15 pg. 103 and problem 86 pg. 111.

D Percent yield (reaction yield)

1. equals the actual yield / theoretical yield x 100
2. actual yield is what is measured in lab
3. theoretical yield is the maximum amount that can be produced, using stoichiometry
4. Try practice problem 3.16 pg. 106 and problem 90 pg. 112.

E Impure reactant

1. % purity problems
   a. % pure = pure / impure
   b. assume 100 grams of impure thus no moving of decimal
   c. Try ore problems then incorporate limiting reactant then incorporate percent yield

F Molarity problems

1. Molarity, M = moles solute / 1 liter solution
2. Same as impure reactant.

G Incorporate limiting and % yield into problem.