CHEM 110: CHAPTER 3: STOICHIOMETRY: CALCULATIONS WITH CHEMICAL FORMULAS AND EQUATIONS

The Chemical Equation

A chemical equation concisely shows the initial (reactants) and final (products) results of a chemical or physical change along with the amounts of each involved. ***Symbols are used as abbreviations.***

<u>Example 1</u>

$$a)\ S_{(s)} + O_{2(g)} \xrightarrow{\Delta} SO_{2\ (g)}$$

$$b)\ SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$$

(s), (g), (l), and (aq) indicate phases each substance is in. The arrow indicates yields or produces and always follows the reactants. Any substance to the left of the arrow is a reactant, to the right a product.

Sometimes other symbols are present above the arrow such as the Greek letter Δ (delta) which indicates heating was done to speed up or initiate the reaction or the formula for a substance that acts as a catalyst to speed up reaction.

WRITING AND BALANCING EQUATIONS

Equations are balanced in order to obey the law of conservation of matter. The same number of atoms for each substance must appear on the reactant (left) and product (right) side of the equation. Balancing involves placing coefficients (whole numbers, never left as fractions) in front of the formulas for substances involved in the reaction as needed. <u>NOTE:</u> It is not always necessary to change the coefficients of an equation in order to balance it.

Balance the following equations:

a. $Ca_{(s)} + H_2O_{(l)} \rightarrow H_{2(g)} + Ca(OH)_{2(s)}$

b. $Cs_{(s)} + Cl_{2(g)} \rightarrow CsCl_{(s)}$

c. $Ba(NO_3)_{2(aq)} + K_2SO_{4(aq)} \rightarrow BaSO_{4\,(s)} + KNO_{3(aq)}$

d. $Hg_2O_{(s)} \rightarrow Hg_{(l)} + O_{2(g)}$

e. $Al^{+3}{}_{(aq)} + OH^{-}{}_{(aq)} \rightarrow Al(OH)_{3(s)}$

NOTE: In e., ions were balanced. This is normal and often occurs for reactions of ions in solutions.

Balancing equations takes practice!

Useful Information Equations Give

When looking at the formulas involved in a balanced equation, several types of information can be determined:

a. atoms or molecules of a substance involved.

Example: $2\ Na_{(s)} + Cl_{2\ (g)} \rightarrow 2\ NaCl_{(s)}$

The coefficients indicate that for every 2 atoms of Na, 2 formula units of NaCl are formed. Also, 1 molecule of Cl_2 is required to react.

b. moles of substances involved

Example: $6\ CO_{2\ (g)} + 6\ H_2O_{(g)} \rightarrow C_6H_{12}O_{6\ (s)} + 6\ O_{2\ (g)}$

For every 6 moles of CO_2 and 6 moles of H_2O used up, 1 mole of $C_6H_{12}O_6$ and 6 moles of O_2 are produced.

c. grams of a substance involved

Example: $P_{4\ (s)} + 5\ O_{2\ (g)} \rightarrow P_4O_{10\ (s)}$

For every 123.88 g of P_4 and 160.00 g of O_2 consumed, 283.88 g of tetraphosphorus decoxide is produced.

NOTE: Check to see if it obeys Law of Conservation of Mass:

123.88 g of P_4 and 160.00 g of O_2 = 283.88 g of P_4O_{10}? YES!

$$P_{4\,(s)} + 5\ O_{2\,(g)} \rightarrow P_4O_{10\,(s)}$$

1 molecule	**5 molecules**	**1 molecule**
1 mole	**5 moles**	**1 mole**
123.88 g	**160.00 g**	**283.88 g**

Classification of Chemical Reactions

There are 4 main types of chemical reactions that can easily be classified.

NOTE: MANY CHEMICAL REACTIONS DO NOT FIT INTO THESE 4 CATEGORIES. We will focus on these types for this course.

1. **SYNTHESIS OR COMBINATION: As the name implies, this involves 2 or more substances combining chemically to produce a new compound.**

Examples: $2\ Na_{(S)} + Cl_{2(g)} \rightarrow 2\ NaCl_{(s)}$

$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$

NOTE: The second example represents combustion.

The form of each of these equations is A + B → AB.

2. DECOMPOSITION: The reverse of a synthesis reaction. 1 substance breaks down into 2 or more substances. Usually this is accomplished using heat, electricity, or shock.

Examples: $2\ NaCl_{(s)} \rightarrow 2\ Na_{(S)} + Cl_{2(g)}$ **(electricity)**

$$2\,H_2O_{(l)} \rightarrow 2\,H_{2(g)} + O_{2(g)} \quad \text{(electricity)}$$

$$Pb(N_3)_{(s)} \rightarrow Pb_{(S)} + 3\,N_{2(g)} \quad \text{(heat or shock)}$$

The form of each of these equations is $AB \rightarrow A + B$.

3. SINGLE-REPLACEMENT OR DISPLACEMENT: One element replaces (displaces) another element from a compound. The element replaced is then set free as the native element.

Examples: $Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$

$$Br_{2(g)} + 2\,KI_{(s)} \rightarrow I_{2(g)} + 2\,KI_{(s)}$$

The form of each equation is $A + BC \rightarrow B + AC$.

NOTE: The reverse reactions of the 2 examples above would NOT occur spontaneously. Reactions occur spontaneously in one direction only. Outside energy would need to be added to cause the reverse reactions to occur (termed an endothermic process).

How can one determine whether or not a reaction that is single replacement will be spontaneous or not?

THE ACTIVITY SERIES IS USED FOR THIS. See Table 4.1 on pg. 121.

***If no reaction occurs, write "No reaction" or NR after the arrow.**

4. DOUBLE DISPLACEMENT OR DOUBLE DECOMPOSITION: Two *<u>ionic</u>* compounds dissolved in water displace ions from each other twice and form a solid precipitate or weak electrolyte.

<u>Example</u>: $Ca(NO_3)_{2(aq)}$ + $Na_2CO_{3(aq)}$ → $CaCO_{3(s)}$ + 2 $NaNO_{3(aq)}$

To determine which of the products the solid would be, see a Solubility Table.

<u>Example</u>: $NaOH_{(aq)}$ + $HCl_{(aq)}$ → $NaCl_{(aq)}$ + $H_2O_{(l)}$

This reaction is a special type of DD reaction termed a "neutralization reaction."

In this type, a base (NaOH) reacts with an acid (HCl) to yield a salt (NaCl) and water. *Sometimes it makes balancing easier to write H_2O as HOH.*

It is double displacement because H^+ goes with OH- and Na^+ goes with Cl-. (The positive and negative switched.)

The Mole

A mole is a number used to measure a given quantity of a substance (element or compound). This number represents a certain amount of atoms, molecules, or ions and is expressed in one of the following ways:

1. 1 Mole = <u>Avogadro's Number</u> of atoms, molecules, formula units, or ions = 6.022×10^{23} atoms, molecules, formula units, or ions. A humungous number!

2. **1 mole = the atomic weight (mass) of an element (from the Periodic Table, see below symbol)**

The atomic weight can be used to convert directly from grams to moles of a substance and vice versa.

<u>Example 1</u>: 55.2 g of As = ? moles As

Using the atomic weight (mass) of As (round to nearest hundredths place),

$$\mathbf{55.2}\ \cancel{\mathbf{g\ As}} \times \frac{1\ mole\ As}{\cancel{74.92\ g\ As}} = 0.74 \text{ moles As}$$

OR to go from <u>moles to grams</u> use the reciprocal

<u>Example 2</u>: 0.25 moles I_2 = ? g I_2

$$\mathbf{0.25}\ \cancel{\mathbf{moles\ I_2}} \times \frac{253.8\ g\ I2}{\cancel{1\ mole\ I2}} = 63.45 \text{ moles } I_2$$

3. Going from moles to atoms, molecules, formula units, or ions: USE AVOGADRO'S NUMBER. *Be familiar with Figure 3.8 pg.91.*

<u>EXAMPLE 3</u>: 5.276 moles Li = ? atoms Li

$$\mathbf{5.276}\ \cancel{\mathbf{moles\ Li}} \times \frac{6.022 \times 10\ 23\ atoms\ Li}{\cancel{1\ mole\ Li}} = 3.18 \times 10^{24} \text{ atoms Li}$$

OR to go from atoms to moles, use the reciprocal:

Example 4: 5.59 x 10^{29} molecules N_2 = ? moles of N_2

5.59 x 10^{29} molecules N_2 × $\frac{1\ mole\ N2}{6.022 \times 10\ 23\ molecules\ N2}$

= 9.28 × 10^{5} molecules N_2

THE SAME PROCEDURE IS USED FOR ATOMS, FORMULA UNITS (Ionic Compounds), OR IONS.

DETERMINING MOLAR MASS (Formula Weight)

1 mole = the mass of each element (from atomic mass on Periodic Table) in a given substance added together.

NOTE: The subscript indicates how many of each atom is present and how many times to multiply each atomic mass by.

Example 5: Determine the molar mass of fructose (fruit sugar, $C_6H_{12}O_6$)

Number of C atoms X 12.01 (atomic mass) = 6 x 12.01 g = 72.06 g C
Number of H atoms X 1.01 = 12 x 1.01 g = 12.12 g H
Number of O atoms x 16.00 = 6 x 16.00 g = 96.00 g O
TOTAL = 72.06 g + 12.12 g + 96.00 g = 180.18 g/mole
This is the molar mass (gram formula mass) of fructose.

SEE CLASS NOTES FOR DETERMINING % COMPOSITION, EMPIRICAL AND MOLECULAR FORMULAS.

Mole Relationships

1. Recall that 1 mole equals the following:

$6.022x10^{23}$ atoms, molecules, ions, or formula units of a substance

2. Mass of a compound (or element) in grams (molar mass): determined by adding up all the atomic masses of the elements present

3. 22.4 L of a gas at Standard Temperature and Pressure (STP)

Stoichiometry: "mole measure" (Greek) Here "mole" represents a quantity

****A mathematical method of determining quantitative relationships among the reactants and products of a chemical reaction***

One of the most important steps of stoichiometry is determining the mole ratios between reactants and products in a *balanced* equation.

Example: $2\ C_4H_{10(g)} + 9\ O_{2(g)} \rightarrow 8\ CO_{(g)} + 10\ H_2O_{(g)}$

The mole ratios for C_4H_{10} would be:

a. **2 moles C_4H_{10} : 9 moles O_2**
b. **1 mole C_4H_{10} : 4 moles CO → Reduces to 1:4.**
c. **1 mole C_4H_{10} : 5 moles H_2O → Reduces to 1:5.**

2 $C_4H_{10(g)}$ + 9 $O_{2(g)}$ $\rightarrow$ 8 $CO_{(g)}$ + 10 $H_2O_{(g)}$

For O_2

a. 9 moles O_2 : 2 moles C_4H_{10}

b. 9 moles O_2 : 8 moles CO

c. 9 moles O_2 : 10 moles H_2O

For CO

a. 4 moles CO : 1 mole C_4H_{10}

b. 8 moles CO : 9 moles O_2

c. 4 moles CO : 5 moles H_2O

For H_2O

a. 5 moles H_2O : 1 mole C_4H_{10}

b. 10 moles H_2O : 9 moles O_2

c. 5 moles H_2O : 4 mole CO

***Knowing how to determine mole ratios is an important step towards solving these problems.**

TYPES OF STOICHIOMETRY PROBLEMS

A. mole – mole: given moles of one substance, find moles of another

B. mole – mass: given moles of one substance, find mass of another

C. mass – mole: given mass of one substance, find moles of another

d. mass – mass: given mass of one substance, find mass of another

NOTE: Variants of these can occur if given and/or desired units are atoms, ions, formula units, or liters rather than masses.

e. Limiting Reactant: given quantity of 2 reactants, determine how much product can be produced and/or how much of excess reactant is left over after reaction.

f. % Yield: Determining the theoretical amount of product by doing a stochiometry calculation that can be produced and given actual amount of product produced, *divide actual amount by theoretical amount and multiply by 100 to get % Yield*.

SEE course website (www.psduboischem110.homestead.com) for description of steps towards solving different types of stoichiometry problems.