Unit 3: QUANTITATIVE RELATIONSHIP IN CHEMICAL CHANGE

Chapter 8: Chemical Reactions

8.1: Describing Chemical Change

Reactants: - chemicals that goes into a reaction.

Products: - chemicals that are produced from a reaction.

$$\text{Reactants} \xrightarrow{\text{"yields"}} \text{Products}$$

Chemical Word Equation: - a chemical reaction written out in words.

Skeletal Equation: - a chemical equation that does not show the relative amount of reactants and products.

States of Chemicals: - (*s*) solid, (*l*) liquid, (*g*) gas, (*aq*) aqueous – dissolved in water

Other Chemical Symbols:

1. **Heat is Added**: - $\xrightarrow{\quad heat \quad}$ or $\xrightarrow{\quad \Delta \quad}$
2. **Catalyst is Added**: - $\xrightarrow{\quad \textit{Name of Catalyst} \quad}$

   (**Catalyst**: - a chemical that is used to speed up a reaction but does not get consumed)

Example 1: Convert the following unbalanced chemical equations into word equations.

a. $N_{2\,(g)} + H_{2\,(g)} \rightarrow NH_{3\,(g)}$

Nitrogen gas reacts with hydrogen gas to produce ammonia gas.

nitrogen gas + hydrogen gas → ammonia gas

b. $Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$

Hydrogen gas and sodium hydroxide solution are produced when solid sodium is added to water.

sodium metal + water → sodium hydroxide solution + hydrogen gas

Example 2: Convert the following word equations into skeletal equations.

a. Heating solid diphosphorous pentaoxide decomposes into phosphorous and oxygen.

$P_2O_{5\,(s)} \rightarrow P_{4\,(s)} + O_{2\,(g)}$	**Recall that phosphorus is a polyatomic element, and oxygen is a diatomic element.**

b. Hypochlorous acid is neutralized by barium hydroxide solution to form water and soluble barium hypochlorite.

$HClO_{(aq)} + Ba(OH)_{2\,(aq)} \rightarrow H_2O_{(l)} + Ba(ClO)_{2\,(aq)}$	**It is very important that the subscripts for these ionic compounds are correct.**

Balancing Chemical Equation: - the process by which we place **coefficient (numbers in front of reactants and products)** in an attempt to equate the number of atoms or polyatomic ions for the elements or compounds in a chemical equation.

Steps involve to Balance Chemical Equation

1. Write the chemical formulas of all reactants and products with their proper subscripts and state to form a skeletal equation.

2. Count the number of atoms / polyatomic ions (always view complex ion as a group) of each chemical formula. Balance the atoms / polyatomic ions by writing the coefficient in front of the reactant / product. Do NOT mess with any of the subscripts.

3. Always balance the atoms that appear in more than two chemicals last.

4. Verify that all atoms / polyatomic ions are balanced.

Example 3: Balance the following chemical equations

a. $N_{2\,(g)} + H_{2\,(g)} \rightarrow NH_{3\,(g)}$

$N_{2\,(g)} + \underline{3}\,H_{2\,(g)} \rightarrow \underline{2}\,NH_{3\,(g)}$

b. $Al_2O_{3\,(s)} \rightarrow Al_{(s)} + O_{2\,(g)}$

$\underline{2}\,Al_2O_{3\,(s)} \rightarrow \underline{4}\,Al_{(s)} + \underline{3}\,O_{2\,(g)}$

c. $Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$

$Na_{(s)} + HOH_{(l)} \rightarrow NaOH_{(aq)} + H_{2\,(g)}$

$\underline{2}\,Na_{(s)} + \underline{2}\,HOH_{(l)} \rightarrow \underline{2}\,NaOH_{(aq)} + H_{2\,(g)}$

Sometimes, it is easier to rewrite H_2O as HOH. This is especially true when a OH^- is part of the products. The first H atom in the HOH becomes $H_{2\,(g)}$ and the remaining part, OH, becomes the polyatomic ion, OH^-.

d. $Fe(NO_3)_{3\,(aq)} + Ba(OH)_{2\,(aq)} \rightarrow Fe(OH)_{3\,(s)} + Ba(NO_3)_{2\,(aq)}$

$$\underline{2}\ Fe(NO_3)_{3\,(aq)} + \underline{3}\ Ba(OH)_{2\,(aq)} \rightarrow \underline{2}\ Fe(OH)_{3\,(s)} + \underline{3}\ Ba(NO_3)_{2\,(aq)}$$

Look at each polyatomic ion as one item. There were $(NO_3)_3$ on the left hand side and $(NO_3)_2$ on the right hand side. Hence, we need to use 2 and 3 as coefficients to balance them. Similarly, there were $(OH)_2$ on the reactant side and $(OH)_3$ on the product side. Therefore, we are required to use 3 and 2 to balance them. Note that once the coefficients are in place, the Fe and Ba atoms are also balanced.

e. $C_2H_{6\,(g)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O_{(l)}$

For burning (adding O_2) with hydrocarbons (compounds containing carbon and hydrogen), we must *balance C, H, O in that order*. This is because oxygen atoms exist in more than two compounds on the product side. After we balance carbon and hydrogen, we have the following number of oxygen atoms.

$C_2H_{6\,(g)}$	+	__ $O_{2\,(g)}$	$\rightarrow$	$\underline{2}\ CO_{2\,(g)}$	+	$3\ H_2O_{(l)}$
		need 7 oxygen		**4 oxygen**		**3 oxygen**
$C_2H_{6\,(g)}$	+	$\underline{\frac{7}{2}}\ O_{2\,(g)}$	$\rightarrow$	$\underline{2}\ CO_{2\,(g)}$	+	$\underline{3}\ H_2O_{(l)}$

Multiply all coefficients by 2:

$$2\ C_2H_{6\,(g)} + \underline{7}\ O_{2\,(g)} \rightarrow \underline{4}\ CO_{2\,(g)} + \underline{6}\ H_2O_{(l)}$$

Since elemental oxygen is diatomic, the 7 oxygen needed on the reactant side needs to have a coefficient of 7/2.

This can easily be converted to a whole number by multiplying all the coefficients by 2.

Example 4: Rewrite the following word equations into skeletal equations and balance them to form chemical equations.

a. Sulfur is burned in the presence of oxygen to form sulfur dioxide gas.

First, we write the skeletal equation.
(Recall that sulfur is a polyatomic element, and oxygen is a diatomic element.)

$$S_{8\,(s)} + O_{2\,(g)} \rightarrow SO_{2\,(g)}$$

Then, we balance the equation.

$$S_{8\,(s)} + \underline{8}\ O_{2\,(g)} \rightarrow \underline{8}\ SO_{2\,(g)}$$

b. Hydrogen gas is added to propyne (C_3H_4) gas with the aid of a platinum catalyst to form propane gas.

First, we write the skeletal equation.
(Propane is C_3H_8 – a formula that you should have memorized.)

$$C_3H_{4\,(g)} + H_{2\,(g)} \xrightarrow{Pt} C_3H_{8\,(g)}$$

Then, we balance the equation.

$C_3H_{4\,(g)}$	+	$\underline{2}\ H_{2\,(g)}$	$\xrightarrow{Pt}$	$C_3H_{8\,(g)}$
4 hydrogen		***need 4 more hydrogen***		**8 hydrogen**

Assignment

8.1 pg. 206 #1, 2; pg. 209 #3, 4; pg. 210 #5 to 7; pg. 211 #8 to 11

8.2: Types of Chemical Reactions

There are 5 basic types of chemical reactions:

1. **Formation or Composition** **(Element + Element → Compound)**

   Example: $2\ Mg_{(s)} + O_{2\,(g)} \rightarrow 2\ MgO_{(s)}$

2. **Deformation or Decomposition** **(Compound → Element + Element)**

   Example: $2\ Al_2O_{3\,(s)} \rightarrow 4\ Al_{(s)} + 3\ O_{2\,(g)}$

3. **Single Replacement** **(Element + Compound → Element + Compound)**

   Example: $2\ AgNO_{3\,(aq)} + Cu_{(s)} \rightarrow 2\ Ag_{(s)} + Cu(NO_3)_{2\,(aq)}$

4. **Double Replacement** **(Compound + Compound → Compound + Compound)**

   Example: $AgNO_{3\,(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3\,(aq)}$

5. **Hydrocarbon Combustion** **(Hydrocarbon + Oxygen → Carbon Dioxide + Water)**

   Example: $CH_{4\,(g)} + 2\ O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\ H_2O_{(g)}$

Precipitation Reaction: - a reaction where a precipitate (new solid) is formed as a product.

Neutralization Reaction: - a reaction between an acid and a base where water is formed as a product.

To Predict Products and Balance Chemical Equations:

1. Write the correct chemical formulas for all products and reactants with proper subscripts. The presence of metals or ionic compounds indicates that we will need to use ions and charges to form any products.
2. For hydrocarbon combustion, balance in the order of C, H, and then O.
3. For other type of reactions, balance the equation for each type of cations and anions. Do NOT break up complex ions. Water may be written as HOH (H^+ and OH^-) in single and double replacement reactions.
4. Check with the Solubility Table (see Section 18.1) and the Table of Elements for the states of chemicals.

Example 1: Predict the product(s) along with the states, indicate the type of reaction, and balance the following chemical reactions.

a. Sulfur trioxide gas is produced from its elements.

Formation: $S_{8\,(s)} + \underline{12}\ O_{2\,(g)} \rightarrow \underline{8}\ SO_{3\,(g)}$

b. A solid piece of zinc is immersed in an iron (III) chloride solution.

Single Replacement: $\underline{3}\ Zn_{(s)} + \underline{2}\ FeCl_{3\,(aq)} \rightarrow \underline{3}\ ZnCl_{2\,(aq)} + \underline{2}\ Fe_{(s)}$

Zn^{2+} Fe^{3+} Cl^-

c. Propane ($C_3H_{8\,(g)}$) is burned in a gas barbecue.

Hydrocarbon Combustion: $C_3H_{8\,(g)} + 5\ O_{2\,(g)} \rightarrow 3\ CO_{2\,(g)} + 4\ H_2O_{(g)}$

d. Chlorine gas is bubbled through a copper (II) iodide solution.

Single Replacement: $Cl_{2\,(g)} + CuI_{2\,(aq)} \rightarrow CuCl_{2\,(aq)} + I_{2\,(s)}$

Cl^- Cu^{2+} I^-

e. Ammonia gas is decomposed into its elements.

Decomposition: $2\ NH_{3\,(g)} \rightarrow N_{2\,(g)} + 3\ H_{2\,(g)}$

f. Sulfuric acid is neutralized by sodium hydroxide solution.

Double Replacement: $H_2SO_{4\,(aq)} + 2\ NaOH_{(aq)} \rightarrow 2\ HOH_{(l)} + Na_2SO_{4\,(aq)}$

H^+ SO_4^{2-} Na^+ OH^-

g. Propanol ($C_3H_7OH_{(l)}$) is accidentally ignited.

Hydrocarbon Combustion: $C_3H_7OH_{(l)} + \frac{9}{2} O_{2\,(g)} \rightarrow 3\ CO_{2\,(g)} + 4\ H_2O_{(g)}$

(Multiply Coefficients by 2) $2\ C_3H_7OH_{(l)} + 9\ O_{2\,(g)} \rightarrow 6\ CO_{2\,(g)} + 8\ H_2O_{(g)}$

h. Lead (II) nitrate solution is reacted with chromium (III) sulfate solution.

Double Replacement: $3\ Pb(NO_3)_{2\,(aq)} + Cr_2(SO_4)_{3\,(aq)} \rightarrow 2\ Cr(NO_3)_{3\,(aq)} + 3\ PbSO_{4\,(s)}$

Pb^{2+} NO_3^- Cr^{3+} SO_4^{2-}

i. Octane ($C_8H_{18\,(l)}$) is combusted in an automobile.

Hydrocarbon Combustion: $C_8H_{18\,(l)} + \frac{25}{2} O_{2\,(g)} \rightarrow 8\ CO_{2\,(g)} + 9\ H_2O_{(g)}$

(Multiply Coefficients by 2) $2\ C_8H_{18\,(l)} + 25\ O_{2\,(g)} \rightarrow 16\ CO_{2\,(g)} + 18\ H_2O_{(g)}$

Assignment

8.2 pg. 214 #13, 14; pg. 216 #15, 16; pg. 218 #17; pg. 220–221 #18 to 21; pg.224 #22, 23 (omit 23c)

8.3: Reactions in Aqueous Solution

Molecular Equation: - a chemical equation where compounds are written in their chemical formulas.

Complete Ionic Equation: - a chemical equation where all compounds that are soluble are written in the ionic components (**slightly soluble compounds are not separated into ions**).

Net Ionic Equation: - an ionic equation that only shows the ions responsible in forming the precipitate. **Spectator Ions** (ions that do not form the precipitate) are omitted.

Example 1: Predict all products form when an ammonium phosphate solution reacts with a calcium chloride solution. Explain the reaction in a form of a balanced

a. Molecular Equation

$$2\ (NH_4)_3PO_4\,_{(aq)} + 3\ CaCl_2\,_{(aq)} \rightarrow 6\ NH_4Cl\,_{(aq)} + Ca_3(PO_4)_2\,_{(s)} \downarrow$$

NH_4^+ PO_4^{3-} Ca^{2+} Cl^- **(precipitate)**

b. Complete Ionic Equation

$$6\ NH_4^+\,_{(aq)} + 2\ PO_4^{3-}\,_{(aq)} + 3\ Ca^{2+}\,_{(aq)} + 6\ Cl^-\,_{(aq)} \rightarrow 6\ NH_4^+\,_{(aq)} + 6\ Cl^-\,_{(aq)} + Ca_3(PO_4)_2\,_{(s)} \downarrow$$

(Precipitate does NOT separate into ions)

c. Net Ionic Equation

$$2\ PO_4^{3-}\,_{(aq)} + 3\ Ca^{2+}\,_{(aq)} \rightarrow Ca_3(PO_4)_2\,_{(s)}$$

(Only write the ions that contribute to the precipitated chemical species)

Example 2: Predict all products form when sulfuric acid reacts with a lithium hydroxide solution. Explain the reaction in a form of a balanced

a. Molecular Equation

$$H_2SO_4\,_{(aq)} + 2\ LiOH\,_{(aq)} \rightarrow Li_2SO_4\,_{(aq)} + 2\ HOH\,_{(l)}$$

H^+ SO_4^{2-} Li^+ OH^- **(liquid water)**

b. Complete Ionic Equation

$$2\ H^+\,_{(aq)} + SO_4^{2-}\,_{(aq)} + 2\ Li^+\,_{(aq)} + 2\ OH^-\,_{(aq)} \rightarrow 2\ Li^+\,_{(aq)} + SO_4^{2-}\,_{(aq)} + 2\ HOH\,_{(l)}$$

(Pure Liquid does NOT separate into ions)

c. Net Ionic Equation

$$2\ H^+\,_{(aq)} + 2\ OH^-\,_{(aq)} \rightarrow 2\ HOH\,_{(l)}$$

$$H^+\,_{(aq)} + OH^-\,_{(aq)} \rightarrow H_2O\,_{(l)}$$

This is the main result of acid-base neutralization (the formation of water).

(Only write the ions that contribute to the pure liquid species)

Example 3: Predict all products form when solid aluminum reacts with a copper (II) nitrate solution. Explain the reaction in a form of a balanced

a. Molecular Equation

$$2\ Al\,_{(s)} + 3\ Cu(NO_3)_2\,_{(aq)} \rightarrow 2\ Al(NO_3)_3\,_{(aq)} + 3\ Cu\,_{(s)} \downarrow$$

Al^{3+} Cu^{2+} NO_3^- **(precipitate)**

b. Complete Ionic Equation

$$2\ Al\,_{(s)} + 3\ Cu^{2+}\,_{(aq)} + 6\ NO_3^-\,_{(aq)} \rightarrow 2\ Al^{3+}\,_{(aq)} + 6\ NO_3^-\,_{(aq)} + 3\ Cu\,_{(s)} \downarrow$$

c. Net Ionic Equation

$$2\ Al\,_{(s)} + 3\ Cu^{2+}\,_{(aq)} \rightarrow 2\ Al^{3+}\,_{(aq)} + 3\ Cu\,_{(s)}$$

(Need to write all the ions on both sides that correspond to any solid used or formed)

Assignment

8.3 pg. 226 #25, pg. 228 #26 to 31

Chapter 8 Review: pg. 232 #32 to 36, 38 to 53 (omit 43e, 46 & 52c)

Chapter 9: Stoichiometry

9.1: The Arithmetic of Equations

Stoichiometry: - the calculation of quantities in a chemical reaction.
- the coefficients of various reactants and /or products form mole ratios.
- these mole ratio hold for moles, molecules / atoms, and volumes of gas at constant temperatures and pressures (such as STP and SATP).

Example 1: Interpret the chemical equation $4\ NH_{3\,(g)} + 7\ O_{2\,(g)} \rightarrow 4\ NO_{2\,(g)} + 6\ H_2O_{(g)}$ in terms of

a. moles.
b. molecules.
c. masses.
d. volumes STP.
e. volumes at SATP.

	$4\ NH_{3\,(g)}$	$7\ O_{2\,(g)}$	$4\ NO_{2\,(g)}$	$6\ H_2O_{(g)}$
a.	**4 moles of NH_3**	**7 moles of O_2**	**4 moles of NO_2**	**6 moles of H_2O**
b.	**4 molecules of NH_3**	**7 molecules of O_2**	**4 molecules of NO_2**	**6 molecules of H_2O**
c.	$m = nM$ $m = (4\ \cancel{\text{mol}})(17.04\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 68.16\ \text{g}}$	$m = nM$ $m = (7\ \cancel{\text{mol}})(32.00\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 224.0\ \text{g}}$	$m = nM$ $m = (4\ \cancel{\text{mol}})(46.01\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 184.0\ \text{g}}$	$m = nM$ $m = (6\ \cancel{\text{mol}})(18.02\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 108.1\ \text{g}}$
d.	$V = (4\ \cancel{\text{mol}})(22.4\ \text{L}/\cancel{\text{mol}})$ $\boldsymbol{V = 89.60\ \text{L}}$	$V = (7\ \cancel{\text{mol}})(22.4\ \text{L}/\cancel{\text{mol}})$ $\boldsymbol{V = 156.8\ \text{L}}$	$V = (4\ \cancel{\text{mol}})(22.4\ \text{L}/\cancel{\text{mol}})$ $\boldsymbol{V = 89.60\ \text{L}}$	$V = (6\ \cancel{\text{mol}})(22.4\ \text{L}/\cancel{\text{mol}})$ $\boldsymbol{V = 134.4\ \text{L}}$
e.	$V = (4\ \cancel{\text{mol}})(24.8\ \text{L}/\cancel{\text{mol}})$ $\boldsymbol{V = 99.20\ \text{L}}$	$V = (7\ \cancel{\text{mol}})(24.8\ \text{L}/\cancel{\text{mol}})$ $\boldsymbol{V = 173.6\ \text{L}}$	$V = (4\ \cancel{\text{mol}})(24.8\ \text{L}/\cancel{\text{mol}})$ $\boldsymbol{V = 99.20\ \text{L}}$	$V = (6\ \cancel{\text{mol}})(24.8\ \text{L}/\cancel{\text{mol}})$ $\boldsymbol{V = 148.8\ \text{L}}$

Assignment
9.1 pg. 240-241 #3, 4, 6 to 8

9.2A: Chemical Calculations (Gravimetric, STP & SATP Stoichiometry)

Mole Ratio: - a ratio form between the coefficient of the required chemical amount to the given chemical amount. $\left(\dfrac{\textbf{require coefficient}}{\textbf{given coefficient}}\right)$

Example 1: 4.35 mol of $PCl_{5\,(g)}$ is decomposed into its elements. Write a balance equation and determined the amount of chlorine produced.

$$\underline{4}\ PCl_{5\,(g)} \rightarrow P_{4\,(s)} + \underline{10}\ Cl_{2\,(g)}$$
4.35 mol (PCl_5) ; ? mol (Cl_2)

$$n_{Cl_2} = 4.35\ \cancel{\text{mol } PCl_5} \times \frac{\mathbf{10\ mol\ Cl_2}}{\mathbf{4\ \cancel{mol\ PCl_5}}} = 10.875\ \text{mol } Cl_2$$

$$\boldsymbol{n_{Cl_2} = 10.9\ \text{mol}}$$

Gravimetric Stoichiometry: - stoichiometry that involves quantities of masses.

Gravimetric Stoichiometry Procedure:

1. Predict the products and balance the chemical equation.
2. Put all the information given under the appropriate chemicals and determine the molar masses of the chemical involved.
3. Find the moles of the given chemical. $\left(\boldsymbol{n = \dfrac{m}{M}}\right)$
4. Find the mole of the required chemical using mole ratio. $\left(\textbf{mol of require} = \textbf{mol of given} \times \dfrac{\textbf{require coefficient}}{\textbf{given coefficient}}\right)$
5. Convert mole of the required chemical to its mass equivalence. $(\boldsymbol{m = nM})$

Example 2: Determine the mass of carbon dioxide formed when 50.0 kg of butane ($C_4H_{10\,(l)}$) is burned.

$$2\ C_4H_{10\,(g)} + 13\ O_{2\,(g)} \rightarrow 8\ CO_{2\,(g)} + 10\ H_2O_{(g)}$$
C_4H_{10}: 50.0 kg, 58.14 g/mol ; CO_2: ? g, M = 44.01 g/mol

① $$n_{C_4H_{10}} = \frac{50.0\ \text{k}\cancel{\text{g}}}{58.14\ \cancel{\text{g}}/\text{mol}} = 0.8599931201\ \text{kmol } C_4H_{10}$$

② $$n_{CO_2} = 0.8599931201\ \text{k}\cancel{\text{mol } C_4H_{10}} \times \frac{8\ \text{mol } CO_2}{2\ \cancel{\text{mol } C_4H_{10}}} = 3.43997248\ \text{kmol } CO_2$$

③ $$m_{CO_2} = nM = (3.43997248\ \text{k}\cancel{\text{mol}}\ CO_2)(44.01\ \text{g}/\cancel{\text{mol}})$$

$$\boldsymbol{m_{CO_2} = 151\ \text{kg}}$$

Example 3: Barium chloride solution was mixed with an excess sodium phosphate solution. What was the mass of barium chloride solid needed in the original solution to form 3.21 g of precipitate?

$3\ BaCl_{2\,(aq)}$	+	$2\ Na_3PO_{4\,(aq)}$	→	$Ba_3(PO_4)_{2\,(s)}$	+	$6\ NaCl_{(aq)}$
? g				3.21 g		
208.23 g/mol				M = 601.93 g/mol		

① $n_{Ba_3(PO_4)_2} = \dfrac{3.21\ \cancel{g}}{601.93\ \cancel{g}/mol}$ = 0.005332846 mol $Ba_3(PO_4)_2$

② n_{BaCl_2} = 0.005332846 $\cancel{mol\ Ba_3(PO_4)_2}$ × $\dfrac{3\ mol\ BaCl_2}{1\ \cancel{mol\ Ba_3(PO_4)_2}}$ = 0.015998538 mol $BaCl_2$

③ $m_{BaCl_2} = nM$ = (0.015998538 $\cancel{mol}$ $BaCl_2$)(208.23 g/$\cancel{mol}$) m_{BaCl_2} = **3.33 g**

<u>Gaseous Stoichiometry</u> (STP and SATP): - stoichiometry that involves quantities of volumes.

<u>Gaseous Stoichiometry Procedure</u> (STP and SATP)

1. Predict the products and balance the chemical equation.
2. Put all the information given under the appropriate chemicals.
3. Find the moles of the given chemical.

   (At STP: n = Volume × $\dfrac{1\ mol}{22.4\ L}$; At SATP: n = Volume × $\dfrac{1\ mol}{24.8\ L}$)

4. Find the mole of the required chemical using mole ratio.

   $$\left(\textbf{mol of require} = \textbf{mol of given} \times \frac{\textbf{require coefficient}}{\textbf{given coefficient}}\right)$$

5. Convert mole of the required chemical to its volume equivalence.
   (At STP: $V = n$ × 22.4 L/mol; At SATP = n × 24.8 L/mol)

Example 4: If 15.25 L of hydrogen at STP is reacted with excess amount of nitrogen, determine the volume of ammonia formed at SATP.

$3\ H_{2\,(g)}$	+	$N_{2\,(g)}$	→	$2\ NH_{3\,(g)}$
V = 15.25 L				? L
STP = 22.4 L/mol				SATP = 24.8 L/mol

① n_{H_2} = 15.25 $\cancel{L}$ × $\dfrac{1\ mol}{22.4\ \cancel{L}}$ = 0.6808035714 mol

② n_{NH_3} = 0.6808035714 $\cancel{mol\ H_2}$ × $\dfrac{2\ mol\ NH_3}{3\ \cancel{mol\ H_2}}$ = 0.4538690476 mol NH_3

③ V_{NH_3} = (0.4538690476 $\cancel{mol}$ NH_3)(24.8 L/$\cancel{mol}$) = 11.25595238 L V_{NH_3} = **11.3 L**

Example 5: A piece aluminum metal is placed in an excess amount of sulfuric acid until all the metal is reacted. Calculate the mass of the aluminum used if 5.76 L of hydrogen gas is evolved at STP.

$2\,Al_{(s)}$	+	$3\,H_2SO_{4\,(aq)}$	→	$3\,H_{2\,(g)}$	+	$Al_2(SO_4)_{3\,(aq)}$
M = 26.98 g/mol				5.76 L		
? g				STP = 22.4 L/mol		

① $n_{H_2} = 5.76\ \cancel{L} \times \dfrac{1\text{ mol}}{22.4\ \cancel{L}} = 0.2571428571\text{ mol}$

② $n_{Al} = 0.2571428571\ \cancel{\text{mol } H_2} \times \dfrac{2\text{ mol Al}}{3\ \cancel{\text{mol } H_2}} = 0.1714285714\text{ mol Al}$

③ $m_{Al} = nM = (0.1714285714\ \cancel{\text{mol}}\text{ Al})(26.98\text{ g/}\cancel{\text{mol}})$

m_{Al} = **4.63 g**

Example 6: 35.24 kg of octane ($C_8H_{18\,(l)}$) is burned under excess oxygen. Determine the volume of carbon dioxide gas produced at SATP.

$2\,C_8H_{18\,(l)}$	+	$25\,O_{2\,(g)}$	→	$16\,CO_{2\,(g)}$	+	$18\,H_2O_{(g)}$
35.24 kg				SATP = 24.8 L/mol		
M = 114.26 g/mol				? L		

① $n_{C_8H_{18}} = \dfrac{35.24\text{ k}\cancel{g}}{114.26\ \cancel{g}\text{/mol}} = 0.3084193944\text{ kmol}$

② $n_{CO_2} = 0.3084193944\text{ k}\cancel{\text{mol } C_8H_{18}} \times \dfrac{16\text{ mol } CO_2}{2\ \cancel{\text{mol } C_8H_{18}}} = 2.467355155\text{ kmol } CO_2$

③ $V_{CO_2} = (2.467355155\text{ k}\cancel{\text{mol}}\ CO_2)(24.8\text{ L/}\cancel{\text{mol}})$

V_{CO_2} = **61.2 kL**

Assignment

9.2A pg. 244 #9b, 10; pg. 245 #11, 12; pg. 248 #13, 14; pg. 249 #15, 16; pg. 250 #17 to 22

9.2B: Chemical Calculations (Gaseous and Solution Stoichiometry)

Gaseous Stoichiometry Procedure (Ideal Gas)

1. Predict the products and balance the chemical equation.
2. Put all the information given under the appropriate chemicals.
3. Find the moles of the given chemical $\left(n = \frac{PV}{RT}\right)$.
4. Find the mole of the required chemical using mole ratio.
   $\left(\textbf{mol of require} = \textbf{mol of given} \times \frac{\textbf{require coefficient}}{\textbf{given coefficient}}\right)$
5. Convert mole of the required chemical to its volume equivalence **($PV = nRT$)**

Example 1: Ammonia is reacted with oxygen to form nitrogen monoxide and water vapour.

$$4\ NH_{3\,(g)} + 5\ O_{2\,(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$$

If 50.0 L of ammonia at 90.0 kPa at 25.0°C were allowed to react with excess oxygen, what would be the pressure of nitrogen monoxide in a collector vessel measuring 30.0 L at a temperature of 10.0°C?

$4\ NH_{3\,(g)}$	+	$5\ O_{2\,(g)}$	→	$4\ NO_{(g)}$	+	$6\ H_2O_{(g)}$
50.0 L				30.0 L		
25.0°C = 298.15 K				10.0°C = 283.15 K		
90.0 kPa				? kPa		

① $n_{NH_3} = \frac{PV}{RT} = \frac{(90.0\ \cancel{kPa})(50.0\ \cancel{L})}{\left(8.314\ \frac{\cancel{kPa \bullet L}}{mol \bullet \cancel{K}}\right)(298.15\ \cancel{K})} = 1.815380558\ \text{mol}$

② $n_{NO} = 1.815380558\ \cancel{\text{mol NH}_3} \times \frac{4\ \text{mol NO}}{4\ \cancel{\text{mol NH}_3}} = 1.815380558\ \text{mol NO}$

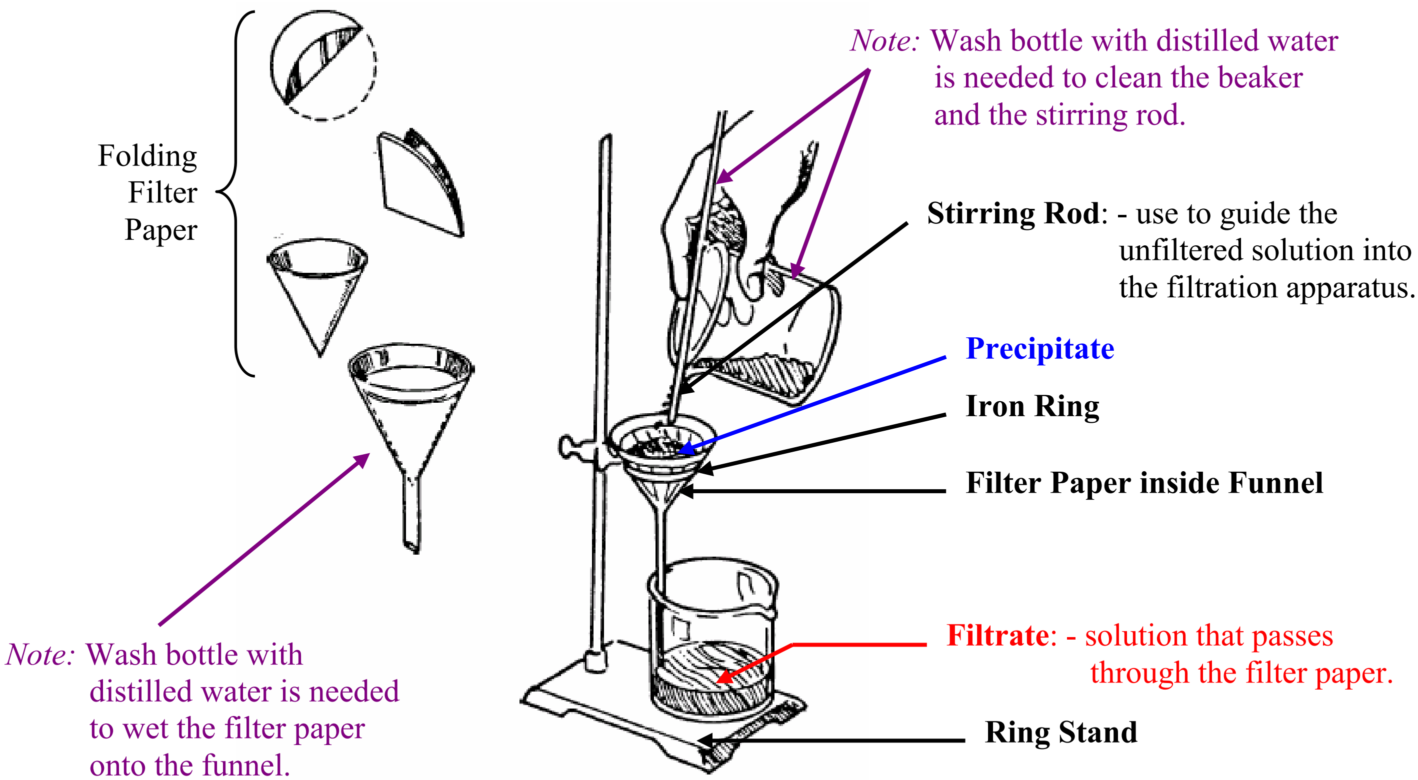
③ $P_{NO} = \frac{nRT}{V} = \frac{(1.815380558\ \cancel{mol})\left(8.314\ \frac{kPa \bullet \cancel{L}}{\cancel{mol \bullet K}}\right)(283.15\ \cancel{K})}{30.0\ \cancel{L}} = 142.453463\ \text{kPa}$ $\boldsymbol{P_{NO} = 142\ \text{kPa}}$

Steps to Solve a Precipitation Reaction:

1. Write a balanced molecular equation. Identify the precipitate.
2. Put the given information underneath the proper chemicals. Identify the limiting reagent if any.
3. Using $n = CV$, convert all given information to moles.
4. Identify and use the information of the limiting reagent if necessary.
5. Determine the moles of precipitate form by using the mole ratio $\left(\frac{\text{Require Coefficient}}{\text{Given Coefficient}}\right)$.
6. Covert moles of precipitate to mass ($m = nM$).

Filtration: - a separation process to isolate the precipitate formed.

Filtration Set-up



Example 2: 200 mL of 0.0500 M of calcium chloride is reacted with an excess amount of ammonium phosphate.

a. Determine the mass of the precipitate formed in this reaction.
b. If the experimental mass of the precipitate is 1.28 g, calculate the % error. How can you interpret this result?

a.

$3\ CaCl_{2\ (aq)}$	+	$2\ (NH_4)_3PO_{4\ (aq)}$	→	$Ca_3(PO_4)_{2\ (s)}$	$+\ 6\ NH_4Cl_{\ (aq)}$
200 mL = 0.200 L		excess		? g	
0.0500 mol/L				M = 310.18 g/mol	

① $n_{CaCl_2} = CV = (0.0500\ \text{mol}/\cancel{\text{L}})(0.200\ \cancel{\text{L}}) = 0.01\ \text{mol}$

② $n_{Ca_3(PO_4)_2} = 0.01\ \cancel{\text{mol CaCl}_2} \times \dfrac{1\ \text{mol Ca}_3(\text{PO}_4)_2}{3\ \cancel{\text{mol CaCl}_2}} = 0.003333\ldots\ \text{mol Ca}_3(\text{PO}_4)_2$

③ $m_{Ca_3(PO_4)_2} = nM = (0.003333\ldots\ \cancel{\text{mol}}\ \text{Ca}_3(\text{PO}_4)_2)(310.18\ \text{g}/\cancel{\text{mol}})$

$\boldsymbol{m_{Ca_3(PO_4)_2} = 1.03\ \text{g}}$

b.

$$\%\ \text{error} = \frac{|\text{Experimental} - \text{Theoretiocal}|}{\text{Theoretical}} \times 100\% = \frac{|1.28\ \text{g} - 1.03\ \text{g}|}{1.03\ \text{g}} \times 100\%$$

% error = 24.3%

This is a significant error. Since the experimental is much higher than the theoretical, we can say that there were a lot of impurities in the precipitate (from the excess ammonium phosphate).

Neutralization: - the reaction between acid and base to produce water and salt.

Example 3:

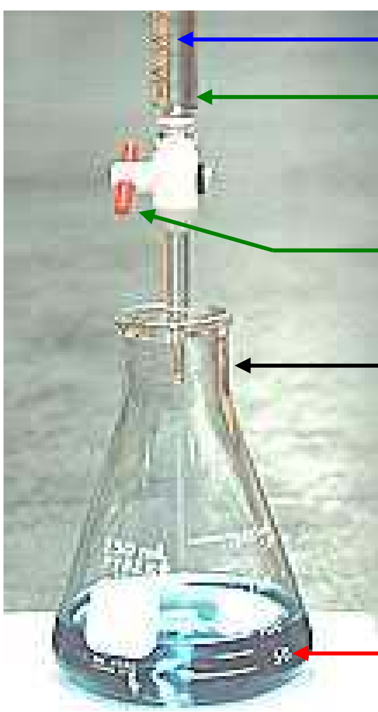
$$HCl_{(aq)} + KOH_{(aq)} \rightarrow HOH_{(l)} + KCl_{(aq)} \quad \text{(Molecular Equation)}$$
$$H^{+}_{(aq)} + Cl^{-}_{(aq)} + K^{+}_{(aq)} + OH^{-}_{(aq)} \rightarrow HOH_{(l)} + K^{+}_{(aq)} + Cl^{-}_{(aq)} \quad \text{(Complete Ionic Equation)}$$
$$H^{+}_{(aq)} + OH^{-}_{(aq)} \rightarrow HOH_{(l)} \quad \text{(Net Ionic Equation)}$$

Steps to Solve a Neutralization Reaction:

1. Write a balanced molecular equation.
2. Put the given information underneath the proper chemicals.
3. Using $n = CV$, convert the given information to moles.
4. Determine the moles of the required chemical by using the mole ratio $\left(\frac{\text{Require Coefficient}}{\text{Given Coefficient}}\right)$.
5. Covert moles of the required chemical to concentration or volume $\left(C = \frac{n}{V} \text{ or } V = \frac{n}{C}\right)$.

Titration: - a volumetric analysis that involves measuring the volume of known concentration solution to measure a given volume of an unknown concentration solution.

Titration Set-up



Titrant: - the solution of known concentration.

Buret: - a precise apparatus to deliver the titrant.
- the volume of the titrant added is read by subtracting the final volume and the initial volume.

Buret Valve: - can be adjusted to let one drop out at a time.

Erlenmeyer Flask: - a container commonly uses to hold the analyte. (Narrow mouth prevents splash and spillage.)

Analyte: - the solution of an unknown concentration.
- the exact volume is usually delivered by a pipet.

Acid-Base Titration: - volumetric analysis that assist in determining the unknown concentration in an acid and base neutralization.

Equivalent Point (Stoichiometric Point): - a point where the number of moles of H^+ is equivalent to the number of moles of OH^-. ($n_{H^+} = n_{OH^-}$)

Endpoint: - a point where the indicator actually changes colour to indicate neutralization is completed.

Indicator: - a chemical that changes colour due to the pH of the solution.

Common Acid-Base Indicators:

a. Bromothymol Blue – Green at pH = 7
b. Phenol Red – Light Orange at pH = 7
c. Phenolphthalein – Light Pink at pH = 9

Example 4: Use the following observation table to determine the concentration of sulfuric acid.

10.0 mL of $H_2SO_{4\,(aq)}$ titrated by 0.0350 mol/L of $KOH_{(aq)}$				
	Trial 1	**Trial 2**	**Trial 3**	**Trial 4**
Initial Volume	0.32 mL	24.19 mL	3.48 mL	24.97 mL
Final Volume	24.19 mL	45.71 mL	24.97 mL	46.47 mL
Volume of KOH added	**23.87 mL**	**21.52 mL**	**21.49 mL**	**21.50 mL**
Bromothymol Blue Colour	Blue	Green	Green	Green

First, we have to complete the table by subtracting the final and the initial volumes. Since the titration is completed when the indicator turns green, we only average the result of the last 3 trials.

$$\text{Average Volume of KOH added} = \frac{21.52\text{ mL} + 21.49\text{ mL} + 21.50\text{ mL}}{3} = 21.50\text{ mL}$$

$2\,KOH_{(aq)}$	+	$H_2SO_{4\,(aq)}$	→	$2\,HOH_{(l)}$	+	$K_2SO_{4\,(aq)}$
21.50 mL		10.0 mL				
0.0350 mol/L		? mol/L				

① $n_{KOH} = CV = (0.0350\text{ mol}/\cancel{\text{L}})(21.50\text{ m}\cancel{\text{L}}) = 0.7525\text{ mmol}$

② $n_{H_2SO_4} = 0.7525\text{ m}\cancel{\text{mol KOH}} \times \dfrac{1\text{ mol } H_2SO_4}{2\ \cancel{\text{mol}}\ \cancel{\text{KOH}}} = 0.37625\text{ mmol } H_2SO_4$

③ $C_{H_2SO_4} = \dfrac{n}{V} = \dfrac{0.37625\ \cancel{\text{m}}\text{mol}}{10.0\ \cancel{\text{m}}\text{L}} = 0.037625\text{ mol/L}$

$[H_2SO_4] = 0.0376$ mol/L

Assignment
9.2B Gas and Solution Stoichiometry Worksheet

Gas and Solution Stoichiometry Worksheet

1. A 25.0 L propane tank at 500.0 kPa and 30.0°C is completely used up during a barbecue cookout. Calculate the volume of carbon dioxide when it reached the upper atmosphere where the pressure is at 50.0 kPa and the temperature is at −20.0°C.

2. 60.0 kL of hydrogen at SATP is reacted with excess amount of nitrogen. The resulting ammonia gas needs to be stored in a gas tank with a volume of 6.25 kL and under 2000 kPa of pressure. What should the temperature regulator of the gas tank be set at?

3. Nitrogen and oxygen in the air regularly react within an automobile engine to form nitrogen monoxide. If 40.0 L of oxygen at 300 kPa is reacted inside a 200°C engine, calculate the volume of smog produced at STP in the exhaust.

4. When nitrogen monoxide reacts with oxygen in the atmosphere under sunlight, it forms photochemical smog (nitrogen dioxide – a brown colour gas). If on an average day, 400,000 ML of nitrogen monoxide at SATP is produced in a large city during rush hour; determine the volume of smog produced when it reached a higher altitude with the pressure at 80.0 kPa and the temperature at −15.0°C.

5. A 15.0 g piece of zinc metal is completely reacted with excess amount of hydrochloric acid at 10.0°C and under 90.0 kPa. Determine the volume of gas produced in this reaction.

6. 56.0 mL of 0.250 M of silver nitrate solution is reacted with a large piece of copper wire. Calculate the mass of the precipitate formed in this reaction.

7. Excess aluminum nitrate is reacted with a 22.6 mL of 0.850 M strontium hydroxide solution.
   a. What is the mass of the precipitate formed in this reaction?
   b. If the experimental mass of the precipitate is 0.92 g, calculate the % error in this experiment. What could contribute to this error?

8. 25.0 mL of hydrochloric acid is titrated by 0.300 mol/L of sodium hydroxide solution. Determine the concentration of the acid if 32.8 mL of base is needed to change the bromothymol blue added to green at the endpoint.

9. What is the volume of 0.750 M of potassium hydroxide needed to completely neutralized 230 mL of 0.325 M phosphoric acid?

10. In an analysis, 15.0 mL of barium hydroxide solution is titrated by 0.650 mol/L of hypochlorous acid. Calculate the concentration of the base if 23.4 mL of acid is required to reach the equivalence point.

Answers:
1. 626 L **2.** 932 K **3.** 137 L **4.** 4.33×10^5 ML **5.** 6.00 L **6.** 1.51 g
7a. 1.00 g **7b.** 8.00%; The experimental value is less than the theoretical. This can mean that some precipitate is lost during filtering or transferring to the funnel. **8.** 0.394 mol/L **9.** 299 mL
10. 0.507 mol/L

9.3: Calculations Involving Limiting Reagents

Excess: - the reactant with more than enough amount for the reaction.

Limiting Reagent: - the reactant with the smaller amount (after taken account of the mole ratio) for the reaction.

Note: A limiting reagent question will always have enough information to find the moles of both reactants.

Steps to deal with Limiting Reagent Problems:

1. Assume one of the reactants is the limiting reagent and determine its mole amount.
2. Determine the mole amount of the other reactant.
3. Use the mole amount of the assumed limiting reagent and the mole ratio, calculate the mole amount of the other reactant actually needed.
4. If the mole amount of the other reactant is smaller than what is needed, then our assumption was wrong. The other reactant is the limiting reagent.
5. If the mole amount of the other reactant is bigger than what is needed, then our assumption was correct. It means that the other reactant is the excess.

Example 1: 5.00 g of phosphorus is reacted with 15.00 g of chlorine gas to produce phosphorus trichloride. Determine the mass of the product produced.

$P_{4\,(s)}$	+	$6\ Cl_{2\,(g)}$	→	$4\ PCl_{3\,(s)}$
5.00 g		15.00 g		? g
M = 123.88 g/mol		M = 70.90 g/mol		M = 137.32 g/mol

Since there is enough information to determine the moles of two reactants, we need to determine which one is the limiting reagent.

① $n_{P_4} = \frac{m}{M} = \frac{5.00\ \cancel{g}}{123.88\ \cancel{g}/mol}$ = 0.04036… mol P_4 ② $n_{Cl_2} = \frac{m}{M} = \frac{15.00\ \cancel{g}}{70.90\ \cancel{g}/mol}$ = 0.2115… mol Cl_2

Let's assume P_4 is the limiting reagent. Calculate the mol Cl_2 actually needed.

③ $n_{Cl_2} = 0.04036\ \cancel{mol\ P_4} \times \frac{6\ mol\ Cl_2}{1\ \cancel{mol}\ P_4}$ = 0.24216…mol Cl_2 needed

But we don't have 0.2416 mol of Cl_2, we only have 0.2115… mol of Cl_2. Therefore, Cl_2 is the limiting reagent. (*Note: the limiting reagent is NOT always the chemical with the smaller number of moles. You have to always compare like we did above.)*

Now, we calculate the moles of PCl_3 formed by using moles of limiting reagent Cl_2.

④ $n_{PCl_3} = 0.2115655853\ \cancel{mol\ Cl_2} \times \frac{4\ mol\ PCl_3}{6\ \cancel{mol}\ \cancel{Cl_2}}$ = 0.1410437236 mol PCl_3

Finally, we determine the mass of PCl_3 produced.

⑤ $m_{PCl_3} = nM$ = (0.1410437236 $\cancel{mol}$ PCl_3)(137.32 g/$\cancel{mol}$) = 19.36812412 g **m_{PCl_3} = 19.4 g**

Example 2: A 30.0 mL of 0.200 M of lead (II) nitrate solution is reacted with 20.0 mL of 0.150 mol/L of sodium sulfate solution. What is the mass of the precipitate formed at the end of this reaction?

$Pb(NO_3)_{2\,(aq)}$	+	$Na_2SO_{4\,(aq)}$	→	$PbSO_{4\,(s)}$	+	$2\ NaNO_{3\,(aq)}$
30.0 mL		20.0 mL		? g		
0.200 mol/L		0.150 mol/L		M = 303.28 g/mol		

Since there is enough information to determine the moles of two reactants, we need to determine which one is the limiting reagent.

① $n_{Pb(NO_3)_2} = CV = (0.200\ \text{mol}/\cancel{\text{L}})(30.0\ \text{m}\cancel{\text{L}}) = 6\ \text{mmol}$

② $n_{Na_2SO_4} = CV = (0.150\ \text{mol}/\cancel{\text{L}})(20.0\ \text{m}\cancel{\text{L}}) = 3\ \text{mmol}$

<u>Since the two reactants are of 1:1 ratio, the chemical wth the smaller moles becomes the limiting reagent.</u> ***(Note: Again, the limiting reagent is <u>NOT</u> always the chemical with the smaller number of moles. We don't have to do the "required over given" to compare only because they are of 1:1 ratio.)***

③ Limiting Reagent is Na_2SO_4

Now, we calculate the moles of $PbSO_4$ formed by using moles of limiting reagent Na_2SO_4.

④ $n_{PbSO_4} = 3\ \text{m}\cancel{\text{mol}\ Na_2SO_4} \times \dfrac{1\ \text{mol}\ PbSO_4}{1\ \cancel{\text{mol}\ Na_2SO_4}} = 3\ \text{mmol}\ PbSO_4$

Finally, we determine the mass of $PbSO_4$ produced.

⑤ $m_{PbSO_4} = nM = (3\ \text{m}\cancel{\text{mol}}\ PbSO_4)(303.28\ \text{g}/\cancel{\text{mol}}) = 909.84\ \text{mg} = 0.90984\ \text{g}$

m_{PbSO_4} **= 0.910 g**

<u>Assignment</u>

9.3 pg. 254 #23, 24; pg. 255 #25; pg. 256 #26; pg. 258 #27, 28; pg. 259 #29 to 32
Chapter 9 Review: pg. 262 #34 to 41, 43 to 47

<u>Assignment</u>

Unit 3 Review: pg. 233–234 #54 to 58, 60 to 62, 64; pg. 235 #1 to 13
pg. 263–264 #51 to 56, 58; pg. 265 #1 to 12