**Unit III: Quantitative Chemistry – Chemical Reactions and Stoichiometry**

**Vocabulary**: stoichiometry, limiting reagent/reactant, excess reactant, theoretical yield, hydrocarbon combustion, percentage yield, titration, precipitate

**1. Chemical Equations and Stoichiometry**

In Grade 10 you enjoyed writing balanced chemical equations because the law of conservation of mass tells us that atoms cannot be created or destroyed in a chemical reaction.

Consider the combustion of hydrogen to form water:

2H2(g) + O2(g) 🡪 2 H20(l)

The coefficients (the numbers you write in front of the formulas) indicate the number of atoms, but they also represent a ratio: there are two H2 molecules to one O2 molecule. Or you could say there is one O2 molecule to two H20 molecules.

If there were 200 H2 molecules, the ratio would be the same: 100 O2 molecules reacting with them, producing 200 H20 molecules.

Now the useful part: if you have 2 moles of H2 (remember that’s 2 x (6.02 x 1023) molecules), it would react with 1 mole of O2, yielding 2 moles of H20. This is called the mole ratio.

e.g. How many moles of moles of CO2 gas are produced when 2 moles of ethane (C2H6) combusts?

e.g. Limestone (CaCO3) undergoes thermal decomposition to turn into lime (CaO) and carbon dioxide gas. If we had 2 kg of limestone, what would be the mass of the lime produced?

a) write a balanced chemical equation

b) determine moles of the known

c) use the mole ratio to determine moles of the unknown

d) convert from moles to mass

Recall Avogadro’s hypothesis that at the same T and P, equal volumes of different gases contain the same number of particles. Therefore, as long as T and P don’t change, the mole ratio is the same as the volume ratio. This only applies to reactions between gases!

e.g. If 40 cm3 of carbon monoxide is reacted with 40 cm3 of oxygen, what volume of CO2 is produced?

Using the mole ratio to answer questions about chemical reactions is called Stoichiometry

**2. Limiting and excess reactants**

In the chemical reactions we have looked at, we have made some assumptions. We have assumed that a) there is enough of each reactant, and b) that they react completely. Here we will look at the first assumption.

In reality, the amount of product is determined by which reactant you have the least of (remember the bike shop analogy). This reactant is called the limiting reactant. The reactant that you have more than enough of is called the excess reactant.

e.g. 4.04g of hydrogen gas is mixed with 16.00g of oxygen gas and the mixture is exploded. What is the mass of water produced?

a) write a balanced chemical reaction

b) determine the limiting reactant i.e. find which reactant would yield less product, assuming the other reactant is in excess.

Theoretical yield: the maximum amount of product that can be obtained, according to the balanced equation and using the given quantities of reactants

Experimental yield: the amount of product that is actually produced

Why is the experimental yield usually less than the theoretical yield?

* Incomplete reaction
* Side reactions in which unwanted substances are produced
* Complete separation of product from reaction mixture is impossible

Percentage yield: (experimental yield ÷ theoretical yield) x 100%

e.g. Johnny wrote a balanced chemical equation for a reaction and used the limiting reactant to calculate that 2.5 g of Unobtainium should have been produced. However, in the actual experiment he only measured 2.0 g of Unobtainium. The % yield would be 2.0/2.5 x 100% = 80%.

e.g. Aspirin, C9H8O4, is made by reacting ethanoic anhydride C4H6O3, with 2-hydroxybenzoic acid, C7H6O3, according to the equation:

2C7H6O3 + C4H6O3 🡪 2C9H8O4 + H20

13.80 g of 2-hydroxybenzoic acid is reacted with 10.26 g of ethanoic anhydride

a) determine the limiting reagent

b) the mass of aspirin actually obtained was 10.90 g. Calculate the percentage yield.

*\*ACT – Precipitate lab*

**3. Reactions in Solutions**

Until now, we have looked at mixing two solutions together and determining the concentration of ions. There were no reactions between the ions in solution. Now we will look at some reactions involving solutions: precipitation reactions acid-base neutralization reactions.

*ACT: (Demo : 0.1 M Pb(NO3)2 + 0.1 M KI).*

*Draw representation of what’s happening*

1. **Net Ionic Equations and Precipitation reactions**

The reactions that we are going to talk about here are all double replacement reactions so if a reaction does take place, we can determine the products. All of the ions in the reactants are aqueous. However, it is clear that one of the products is a solid. This is called a precipitate. Later you will learn how to determine which of the products the precipitate is, but for this reaction it is PbI2

*Pb(NO3)2 + 2 KI ---> PbI2 + 2 KNO3*

We could write this in a way that better represents what is happening, called the overall ionic equation:

Pb2+(aq) + 2NO3-(aq) + 2K+(aq) + 2I-(aq) ---> PbI2 (s) + 2K+(aq) + 2NO3-(aq)

Not used that often because it takes up too much space! Also, you’ll notice that there are 2NO3-(aq) ions and 2K+(aq) ions on either side of the equation. If we cancel them out, we get:

This is called a NET IONIC EQUATION(NIE). It shows exactly what is happening in the solution. The NO3-(aq) ions and the K+(aq) ions are actually doing nothing in the reaction. They can be thought of as just watching the lead ions and the iodide ions get together. For this reason, they are referred to as SPECTATOR IONS. Spectator ions are never included in net ionic equations.

Practice:

CuSO4 + 2 NaOH ---> Cu(OH)2 + Na2SO4 \*assume that the precipitate is Cu(OH)2

1. Write the overall ionic equation:

ii) Write the net ionic equation:

You will notice that for these examples, NIE is simply how we would go about making the solid.

e.g. If we knew we got a precipitate of Ca3(PO4)2 in a reaction, the net ionic equation would be

3 Ca2+(aq) + 2 PO43-(aq) ---> Ca3(PO4)2 (s)

This would be the net ionic equation for the formation of the precipitate Ca3(PO4)2 no matter what solutions were used. The other ions would just be spectator ions.

1. **Acid-Base neutralization reactions**

When an acid and a base are mixed together, a precipitate is NOT formed but a reaction does take place.

Acids we know - HCl, H2SO4 , HNO3 etc. If we look at them carefully, all acids contain an H atom. More specifically, when it ionizes, an H+ ion is released.

Bases can be identified by the presence of an OH- ion.

When an acid is added to a base, water and a salt are formed (review from types of reactions). Let's see what happens in the reaction between NaOH and HCl.

NaOH + HCl ---> H2O + NaCl

Na+(aq) + OH-(aq) + H+(aq) + Cl-(aq) ---> H2O (l) + Na+(aq) + Cl-(aq)

NIE:

This is the net ionic equation for the reaction between any acid and any base. Notice the change in state from aqueous to a liquid.

*\*ACT: Lab 16-D*

1. **Formation of precipitates**

We saw in the lab that some compounds are soluble in water and some are not. We will use the definition here that soluble means that we can make at least a 0.1 M solution of the compound at 25°C. How can we predict if a precipitate will form? The answer is not easy. In fact we don't have too many rules that always apply. Chemists have done thousands of experiments similar to the one you did yesterday and have collated the results into pages and pages of data tables. We will use a portion of those tables and rules.

To use this table, we determine what the product will be in a reaction (remember, the reactions here are double replacement reactions.) From here we must check out both products to see if there is a precipitate formed.

Examples : 1) Identify the precipitate (if any) when the following solutions are mixed:

a) lead (II) nitrate and sodium chloride

b) ammonium hydroxide and copper (I) sulfate

c) potassium phosphate and cesium sulfide

2) Write the net ionic equation for the formation of any precipitate when solutions of the following are mixed. If no precipitate is formed, write NO REACTION.

a) barium sulfide and ammonium carbonate

b) silver nitrate and sodium acetate

c) iron (III) chloride and copper (II) chloride

d) lead (II) acetate and lithium phosphate

1. **Precipitation Formation and Ionic Concentrations**

As we can see from the above and from the lab, sometimes we get a precipitate formed in a reaction. What does the formation of a precipitate do to the concentrations of the ions in solution? It lowers it for those involved. We will, for the sake of argument here, assume that one of the ions is completely used up in the reaction while some of the other one may still remain. If this sounds familiar, it should. This is really a stoichiometric INXS calculation.

Example : 25.0 mL of 1.00 M NaCl is mixed with 35.0 mL of 2.20 M AgNO3. A precipitate is formed and falls out of solution. Calculate the concentrations of all ions remaining in solution after the precipitation stops.

1. Use the data table to write the NIE: *Ag+(aq) + Cl-(aq) ---> AgCl (s)*

Set up a table similar to the one below :

**Ion Moles Before Moles After Volume Concentration**

Na+ 0.0250 mol 0.0250 mol 0.0600 L 0.417 M

Cl- 0.0250 mol 0 mol 0.0600 L 0.000 M

Ag+ 0.0770 mol 0.0520 mol 0.0600 L 0.867 M

NO3- 0.0770 mol 0.0770 mol 0.0600 L 1.28 M

Example : 20.0 mL of 0.500 M AgNO3 is mixed with 15.2 mL of 0.750 M Na2S. Calculate the concentrations of all ions remaining in solution.

NIE : 2 Ag+(aq) + S2-(aq) ---> Ag2S (s)

**Ion Moles Before Moles After Volume Concentration**

1. **Acids and Bases**

We should be able to identify an acid and a base from its properties and its formula.

**Acids :**

1) taste sour

2) conduct an electrical current

3) cause certain dyes to change colour (e.g. litmus goes red)

4) liberates hydrogen when it reacts with certain metals

5) loses the above properties when mixed with a base though the resulting solution conducts electricity.

**Bases :**

1) taste bitter

2) conduct an electrical current

3) cause certain dyes to change colour (e.g. litmus goes blue, phenolphthalein goes pink)

4) feels slippery

5) loses the above properties when mixed with a base though the resulting solution conducts electricity.

In their respective chemical formula, and more importantly when they dissolve in water, acids have an H+ ion while bases have an OH- ion. (This is an incomplete definition of an acid and a base but will suffice for chemistry 11.)

When an acid and a base react, the result is a salt and water. The type of reaction is called a **NEUTRALIZATION** reaction because the properties of the acid and the base are neutralized by each other. The net ionic reaction for this reaction is as follows : H+(aq) + OH-(aq) ---> H2O (l)

We must remember that in any neutralization reaction, the moles of H+ MUST equal the moles of OH-.

Example : What volume of 6.00 M HCl must be added to 125 mL of 1.59 M NaOH in order to neutralize it ?

*Answer :*

Example : What volume of 0.10 M Ca(OH)2 is needed to neutralize 115.2 mL of 0.55 M H3PO4?

*Answer :*

1. **Titrations**

A titration reaction is simply a neutralization reaction using special techniques. A given volume of an acid is placed in a flask (as well as some indicator - phenolphthalein is commonly used). The base is placed in a burette and added drop wise until the colour of the solution in the flask just changes. From here we can get the volumes needed for the neutralization calculation.