Concentration of Solutions: Amount of Solute (in g or mol) dissolved in a unit volume (dm^3) of a solvent

- Can be expressed as g dm^{-3} or mol dm^{-3}

$$\text{Concentration in g } dm^{-3} = \frac{\text{mass of solute (g)}}{\text{volume of solvent } (dm^3)}$$

$$\text{Concentration in mol } dm^{-3} = \frac{\text{No. of moles of solute (mol)}}{\text{Volume of solvent } (dm^3)}$$

Example from AISS Notes

Calculate the concentration of hydrochloric acid solution in mol/dm^3 if 400 cm^3 of the solution contains 0.25 mole of HCl.

Calculate the number of moles of sodium hydroxide contained in 150 cm^3 of 2.0 mol/dm^3 solution of sodium hydroxide.

A solution of sodium hydroxide contains 9.6 g of sodium hydroxide in 100 cm^3 of solution. Find the concentration of the solution in (a) g/dm^3; (b) mol/dm^3

Calculation from Equation

A chemical Equation tells us the mole ratio of reactant and products in a reaction.

$2H_2 + O_2 \text{ -> } 2H_2O$

This means that for every 2 mol of Hydrogen Gas and 1 mol of Oxygen gas that reacts, 2 mol of water will be produced.

- If 0.4 mol of H_2 react with 0.2 mol of O_2, 0.4 mol of H_2O is produced

Example from AISS Notes

Calculate the mass of water produced when 0.5 g of hydrogen gas is burnt in oxygen.

$2H_2 + O_2 \rightarrow 2H_2O$

Magnesium reacts with hydrogen chloride according to the equation: $Mg + 2HCl \rightarrow MgCl_2 + H_2$

Calculate the mass of hydrogen chloride, HCl required to make 20 g of magnesium chloride, $MgCl_2$.

Nitrogen gas and hydrogen gas reacts together to form ammonia gas.

$N_2 + H_2 \rightarrow NH_3$

Balance the equation above and calculate the volume of ammonia gas produced when 12 dm^3 of nitrogen gas is used.

Sulfuric acid and sodium hydroxide can react as follows:

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

(a) Calculate the relative formula mass of sodium sulfate and sodium hydroxide. (b) What is the mass of sodium sulfate that can be formed from 40 g of sodium hydroxide?

Limiting and Excess Reactants

Let's say we have a chemical equation

$2H_2 + O_2 \rightarrow 2H_2O$

If we use 2 mol of H_2 to react with 2 mol of O_2, 1 mol of O_2 will be in excess, so H_2 is the limiting reagent in this reaction.

2 mol of H_2O will be produced.

Hint: If the question provides too much information of the starting reactants, most likely there is an excess reagent. DON'T GET TRICKED!!

Hint: Excess Reagent remains with product! DO NOT FORGET!

Example from AISS Notes

If 8 g of copper was mixed with 4 g of sulfur, find the mass of copper(I) sulfide formed.

$2\ Cu\ (s) + S\ (s) \rightarrow Cu_2S\ (s)$

A sample of 50 cm^3 of propane was burnt in 50 cm^3 of oxygen gas.

$C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$

(a) Find the volume of carbon dioxide gas produced.
(b) Find the volume of excess reactant left.

If 7.2 g of iron(II) oxide was mixed with 0.4 dm^3 of oxygen gas, according to this equation:

$4\ FeO + O_2 \rightarrow 2\ Fe_2O_3$

(a) find the mass of iron(III) oxide formed. (b) find the mass of excess reactant left.

Example from AISS Notes

When 1.92 g of magnesium was heated in excess oxygen, 3.0 g of magnesium oxide was obtained. Calculate the percentage yield of magnesium oxide.

In an experiment, 30 g of ethanoic acid (CH_3COOH) were obtained from the oxidation of 69 g of ethanol (C_2H_5OH). Calculate the percentage yield in the reaction.

$C_2H_5OH + 2\ [O] \rightarrow CH_3COOH + H_2O$

Percentage Yield and Purity

- Amount of product formed in a reaction is the yield
- Theoretical yield is the calculated amount of products that would be obtained if the reaction is completed
- Actual Yield is the amount of pure product actually produced in the experiment
- Actual yield is usually less than the theoretical yield

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Percentage Purity: Amount of substance present in a mixture (impure substance)

$$\% \text{ purity} = \frac{\text{Mass of pure substance in sample}}{\text{Mass of sample}} \times 100\%$$

Example from AISS Notes

3.2 g of copper was heated in air. 3.8 g of copper(II) oxide was formed. What is the percentage purity of copper?

3.21 g sample of impure copper(II) carbonate was reacted with excess hydrochloric acid. It was found that 480 cm³ of carbon dioxide gas measured at r.t.p. was given off. What is the percentage purity of the copper(II) carbonate in the given sample?

Titration Calculations

Equivalence Point: Reactants have reacted just nicely with each other via mole ratios (from balanced eqn)

End Point: Indicator in Titration just changed its color

Common Indicators

- methyl orange : red $\xrightarrow{}$ orange (4) $\rightarrow$ yellow
  (pH 3-5)
- phenolphthalein : colourless $\rightarrow$ pale pink (9) $\rightarrow$ red
  (pH 8-10)
- thymolphthalein : colourless $\rightarrow$ pale blue (9-10) $\rightarrow$ blue.
  (pH 9-10.5)
- thymol blue: red $\rightarrow$ orange (2) $\rightarrow$ yellow $\rightarrow$ green (9) $\rightarrow$ blue
  (pH 1-3)
  (pH 8-10)

<u>Dilution Formula</u>

$$C_1V_1 = C_2V_2$$

Example from AISS Notes

Aqueous sodium sulfate can be prepared by titrating dilute sulfuric acid with aqueous sodium hydroxide. The equation for the reaction is:

H_2SO_4 (aq) + 2 NaOH (aq) → Na_2SO_4 (aq) + 2 H_2O (l)

In the titration, 25.0 cm³ of 2.0 mol/dm³ sulfuric acid was used.

(a) Calculate the volume of 0.5 mol/dm³ sodium hydroxide used in the reaction. (b) Calculate the mass in grams, of sodium sulfate produced in the reaction.

A concentrated acid has a concentration of 12 mol/ dm³. What volume of concentrated acid would be needed to make 1 dm³ of dilute acid with a concentration of 1 mol/ dm³.

Summary of Common Formulas