

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}

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

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

Example from AISS Notes

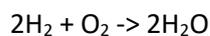
Calculate the concentration of hydrochloric acid solution in mol/dm^3 if 400 cm^3 of the solution contains 0.25 mole of HCl.

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.

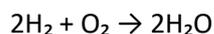


This means that for every 2 mol of Hydrogen Gas and 1 mol of Oxygen gas that reacts, 2 mol of water will be produced.

- 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

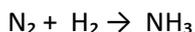
Calculate the mass of water produced when 0.5 g of hydrogen gas is burnt in oxygen.



Magnesium reacts with hydrogen chloride according to the equation: $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$

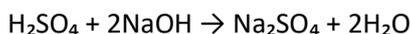
Calculate the mass of hydrogen chloride, HCl required to make 20 g of magnesium chloride, MgCl_2 .

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



Balance the equation above and calculate the volume of ammonia gas produced when 12 dm³ of nitrogen gas is used.

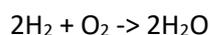
Sulfuric acid and sodium hydroxide can react as follows:



(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



If we use 2 mol of H₂ to react with 2 mol of O₂, 1 mol of O₂ will be in excess, so H₂ is the limiting reagent in this reaction.

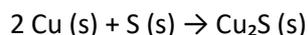
2 mol of H₂O will be produced.

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

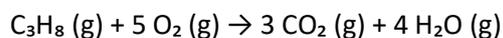
Hint: Excess Reagent remains with product! DO NOT FORGET!

Example from AISS Notes

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

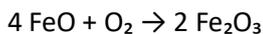


A sample of 50 cm³ of propane was burnt in 50 cm³ of oxygen gas.



(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 dm³ of oxygen gas, according to this equation:

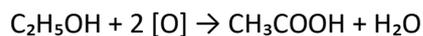


(a) find the mass of iron(III) oxide formed. (b) find the mass of excess reactant left.

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 (CH₃COOH) were obtained from the oxidation of 69 g of ethanol (C₂H₅OH). Calculate the percentage yield in the reaction.



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

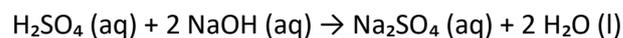
- methyl orange : red $\xrightarrow{4}$ orange $\xrightarrow{4}$ yellow
(pH 3-5)
- phenolphthalein : colourless $\xrightarrow{9}$ pale pink $\xrightarrow{9}$ red
(pH 8-10)
- thymolphthalein : colourless $\xrightarrow{9-10}$ pale blue $\xrightarrow{9-10}$ blue.
(pH 9-10.5)
- thymol blue : red $\xrightarrow{2}$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

Dilution Formula

$$C_1 V_1 = C_2 V_2$$

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:



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/dm^3 . What volume of concentrated acid would be needed to make 1 dm^3 of dilute acid with a concentration of 1 mol/dm^3 .

Summary of Common Formulas

