

CHAPTER ONE

THE MOLE CONCEPT

The Mole (Mol)

A mole of a substance is the amount of substance which contains as many particles as there are carbon particles in 12 grams of carbon -12.

Avogadro Number or Avogadro constant, N_A

The number of particles in one mole of any substance is 6.02×10^{23} . This number is called Avogadro number or Avogadro constant, N_A .

The particles being referred to here can be atoms, molecules, ions, electrons, protons etc.

E.g

1 mole of hydrogen atoms contains 6.02×10^{23} atoms

1 mole of hydrogen molecules contains 6.02×10^{23} molecules

1 mole of hydrogen ions contains 6.02×10^{23} ions.

The following hints are useful when working with moles and Avogadro number.

$$\text{Number of particles in a substance} = \text{Number of moles} \times \text{Avogadro's constant}$$

$$\text{Or } \text{Number of moles in a substance} = \frac{\text{Number of particles}}{\text{Avogadro number}}$$

Worked examples;

1. Calculate
 - (i) the number of particles in 0.5 moles of sodium hydroxide.
 - (ii) the number of moles of particles in 1.8×10^{24} .

Solution

$$\begin{aligned} \text{(i) Number of particles in a substance} &= \text{Number of moles} \times \text{Avogadro's constant} \\ &= 0.5 \times 6.02 \times 10^{23} \\ &= 3.01 \times 10^{23} \text{ particles.} \end{aligned}$$

$$\begin{aligned} \text{(ii) 1 mole of a substance contains } &6.02 \times 10^{23}. \\ \text{Moles in } 1.8 \times 10^{24} \text{ particles} &= \frac{1.8 \times 10^{24}}{6.02 \times 10^{23}} \\ &= 2.99 \times 10^0 \\ &= \mathbf{3 \text{ moles}} \end{aligned}$$

The terms used:

The following are the common terms used in the study of the mole concept.

- Relative Atomic Mass (RAM).
- Relative Molecular Mass (RAM).
- Molar Mass (MM).

(a) Relative Atomic Mass (RAM)

The relative atomic mass of an element is the mass of one atom of the element divide by the mass of one-twelfth of an atom of Carbon – 12.

$$\text{I.e. Relative Atomic Mass} = \frac{\text{Mass of one atom of an element}}{\frac{1}{12} \times \text{Mass of one atom of Carbon-12}}$$

(b) Relative Molecular Mass (RMM)

The Relative molecular mass of an element or compound is the mass of one molecule divided by the mass of one-twelfth of an atom of Carbon-12.

$$\text{Relative Molecular Mass} = \frac{\text{Mass of one atom of an element}}{\frac{1}{12} \times \text{Mass of one atom of Carbon-12}}$$

Example

1. Given, H=1.008, calculate
 - (i) RAM
 - (ii) RMM of hydrogen

Solution

$$\begin{aligned} \text{(i) Relative Atomic Mass} &= \frac{\text{Mass of one atom of an element}}{\frac{1}{12} \times \text{Mass of one atom of Carbon-12}} \\ &= \frac{\text{Mass of one atom of hydrogen}}{\frac{1}{12} \times \text{Mass of one atom of Carbon-12}} \\ &= \frac{1.008}{\frac{1}{12} \times 12} \\ &= 1.008 \\ &= \mathbf{1} \end{aligned}$$

$$\begin{aligned}
 \text{(ii) Relative Molecular Mass} &= \frac{\text{Mass of one molecule of a substance}}{\frac{1}{12} \times \text{Mass of one atom of Carbon-12}} \\
 &= \frac{\text{Mass of one molecule of hydrogen}}{\frac{1}{12} \times \text{Mass of one atom of Carbon-12}} \\
 &= \frac{2 \times 1.008}{\frac{1}{12} \times 12} \\
 &= 2 \times 1.008 \\
 &= 2.016 \\
 &= 2
 \end{aligned}$$

(c) Molar Mass (MM)

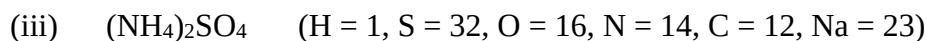
The molar mass of a compound is the mass of one mole of a compound.
It is got by getting the sum of the relative atomic masses of the respective elements in the compound.

The SI unit is - Kg/mol or Kgmol^{-1}

The common unit used is - g/mol or gmol^{-1}

Examples

1. Calculate the molar masses of the following compounds.



Solution

(i) The molar mass of sulphuric acid, H_2SO_4 = $(2 \times 1) + 32 + (4 \times 16)$
 = $2 + 32 + 64$
 = **98 g/mol**

(ii) The molar mass of sodium carbonate, Na_2CO_3 ,
 = $(2 \times 23) + 12 + (3 \times 16)$
 = $46 + 12 + 48$
 = **106 g/mol**

(iii) The molar mass of Ammonium sulphate = $(\text{NH}_4)_2\text{SO}_4$
 $= 2[14 + (1 \times 4)] + 32 + (4 \times 16)$
 $= (2 \times 18) + 32 + 64$
 $= 36 + 32 + 64$
 $= 132 \text{ g/mol}$

Relative Atomic Masses of the common Elements

Name	Symbol	RAM
Aluminum	Al	27
Calcium	Ca	40
Carbon	C	12
Chlorine	Cl	35.5
Copper	Cu	63.5
Hydrogen	H	1
Iodine	I	127
Iron	Fe	56
Lead	Pb	207
Magnesium	Mg	24
Nitrogen	N	14
Oxygen	O	16
Silver	Ag	108
Sodium	Na	23
Sulphur	S	32
Zinc	Zn	65

Calculations from Formulae and Equations

(a) To Calculate the Formula Mass

- Write the chemical formula of the compound.
- Get the sum of the relative atomic masses of the element present in the compound.

Examples

1. Calculate the formula masses of the following compounds

(a) Sulphuric acid.

(b) Calcium Nitrate.

(c) Ammonium Nitrate

(H = 1, S = 32, O = 16, N = 14, Ca = 40)

Solution

(a) Formula of Sulphuric acid $= \text{H}_2\text{SO}_4$
 $= (2 \times 1) + 32 + (4 \times 16)$
 $= 2 + 32 + 64$
 $= \mathbf{98 \text{ g}}$

(b) Formula of Calcium Nitrate $= \text{Ca}(\text{NO}_3)_2$
 $= 40 + 2(14 + 3 \times 16)$
 $= 40 + 2 \times 62$
 $= 40 + 124$
 $= \mathbf{164 \text{ g}}$

(c) Formula of Ammonium Sulphate $= (\text{NH}_4)_2\text{SO}_4$
 $= 2(14 + 4 \times 1) + 32 + (4 \times 16)$
 $= 2(14 + 4) + 32 + 64$
 $= (2 \times 18) + 96$
 $= 36 + 96$
 $= \mathbf{132 \text{ g}}$

(b) To Calculate the Percentage by mass of an Element in a compound

The percentage by mass of an element in a compound can be calculated by using the formula:

$$\text{Percentage by Mass of an element} = \frac{\text{Mass of Element in a Compound}}{\text{Formula Mass}} \times 100$$

Examples

1. Calculate the percentage by mass of oxygen in Calcium oxide.

(Ca = 40, O = 16)

Solution

Formula of Calcium oxide = CaO

Mass of oxygen in CaO = 16

Formula mass of CaO
= 40 + 16
= 56

$$\begin{aligned}\text{Percentage by mass of an element} &= \frac{\text{Mass of Element in a Compound}}{\text{Formula Mass}} \times 100 \\ &= \frac{\text{Mass of oxygen}}{\text{Formula mass}} \times 100 \\ &= \frac{16}{56} \times 100 \\ &= 28.57 \%\end{aligned}$$

(c) To Calculate the Percentage Composition by mass of a compound

Examples

1. Calculate the percentage composition by mass of sodium hydroxide.

(Na = 23, O = 16, H = 1)

The formula mass of NaOH = 23 + 16 + 1
= 40 g

$$\begin{aligned}\text{Percentage by mass of sodium} &= \frac{\text{Mass of sodium}}{\text{Formula mass}} \times 100 \\ &= \frac{23}{40} \times 100 \\ &= 57.5\%\end{aligned}$$

$$\begin{aligned}
 \text{Percentage by mass of oxygen} &= \frac{\text{Mass of oxygen}}{\text{Formula mass}} \times 100 \\
 &= \frac{16}{40} \times 100 \\
 &= \mathbf{40\%}
 \end{aligned}$$

$$\begin{aligned}
 \text{Percentage by mass of hydrogen} &= \frac{\text{Mass of hydrogen}}{\text{Formula mass}} \times 100 \\
 &= \frac{1}{40} \times 100 \\
 &= \mathbf{2.5\%}
 \end{aligned}$$

Or

$$\begin{aligned}
 \text{Percentage by mass of hydrogen} &= 100\% - (57.5 + 40)\% \\
 &= (100 - 97.5)\% \\
 &= \mathbf{2.5\%}
 \end{aligned}$$

2. Calculate the percentage composition by mass of the following compounds.

- (a) Calcium oxide.
- (b) Sodium sulphate.
- (c) Ammonium sulphate.

(Ca = 40, O = 16, Na = 23, S = 32, N = 14, H = 1)

Solution

(a)

$$\begin{aligned}
 \text{Formula of Calcium oxide} &= \text{CaO} \\
 \text{Mass of oxygen in CaO} &= 16 \\
 \text{Formula mass of CaO} &= 40 + 16 \\
 &= \mathbf{56}
 \end{aligned}$$

$$\begin{aligned}
 \text{Percentage by mass of oxygen} &= \frac{\text{Mass of oxygen}}{\text{Formula mass}} \times 100 \\
 &= \frac{16}{56} \times 100 \\
 &= \mathbf{28.57\%}
 \end{aligned}$$

$$\begin{aligned}
 \text{Percentage by mass of calcium} &= \frac{\text{Mass of calcium}}{\text{Formula mass}} \times 100 \\
 &= \frac{40}{56} \times 100 \\
 &= \mathbf{71.43\%}
 \end{aligned}$$

Or Percentage by mass of calcium = (100 – 28.57) %
= **71.43 %**

(b) Formula of Sodium sulphate = Na_2SO_4
Formula mass of Sodium sulphate = (2 x 23) + 32 + (4 x 16)
= 46 + 32 + 64
= **142**

Mass of sodium in Na_2SO_4 = 2 x 23
= **46**

$$\begin{aligned}
 \text{Percentage by mass of sodium} &= \frac{\text{Mass of sodium}}{\text{Formula mass}} \times 100 \\
 &= \frac{46}{142} \times 100 \\
 &= \mathbf{32.39\%}
 \end{aligned}$$

$$\begin{aligned}
 \text{Percentage by mass of sulphur} &= \frac{\text{Mass of sulphur}}{\text{Formula mass}} \times 100 \\
 &= \frac{32}{142} \times 100 \\
 &= \mathbf{22.54\%}
 \end{aligned}$$

$$\begin{aligned}
 \text{Percentage by mass of oxygen} &= \frac{\text{Mass of oxygen}}{\text{Formula mass}} \times 100 \\
 &= \frac{64}{142} \times 100 \\
 &= \mathbf{45.07\%}
 \end{aligned}$$

Or Percentage by mass of oxygen = 100% - (32.39% + 22.54%)
= 100% - 54.93%
= **45.07 %**

(d) To calculate the Number of moles for a given mass of an element or a compound

$$(i) \quad \text{Number of moles of an element} = \frac{\text{Mass of element}}{\text{RAM}}$$

$$(ii) \quad \text{Number of moles of a compound} = \frac{\text{Mass of compound}}{\text{RFM}}$$

Examples

1. Calculate the number of moles in 2.3 g of sodium (Na = 23).

Solution

$$\begin{aligned} \text{Number of moles of an element} &= \frac{\text{Mass of element}}{\text{RAM}} \\ &= \frac{2.3}{23} \\ &= \mathbf{0.1 \text{ moles}} \end{aligned}$$

2. Calculate the number of moles in 50.3g of sodium carbonate.
(Na = 23, C = 12, O = 16)

Solution

$$\begin{aligned} \text{Mass of sodium carbonate} &= 50.3 \text{ g,} \\ \text{RFM of sodium carbonate} &= (2 \times 23) + 12 + (3 \times 16) \\ &= 46 + 12 + 48 \\ &= \mathbf{106 \text{ g}} \end{aligned}$$

$$\begin{aligned} \text{Number of moles of a compound} &= \frac{\text{Mass of compound}}{\text{RFM}} \\ &= \frac{50.3}{106} \\ &= \mathbf{0.5 \text{ moles}} \end{aligned}$$

(e) To Calculate Empirical formula and Molecular formula

(i) Empirical formula

Empirical formula of a compound is the formula which shows the simplest ratio of the different atoms in the compound.

Or Empirical formula of a compound is the formula which expresses its composition by mass and the ratio of the numbers of the different atoms present in the compound.

(ii) Molecular formula

Molecular formula of a compound is the formula which expresses the actual number of each kind of atoms present in one molecule of the compound.

Note: *In some compounds, the empirical formula is the same as the molecular formula. While in others they are different as shown in the table below.*

Name of compound	Molecular formula	Empirical formula
Methane	CH ₄	CH ₄
Ethane	C ₂ H ₆	CH ₃
Ethene	C ₂ H ₄	CH ₂
Propene	C ₃ H ₈	C ₃ H ₈
Glucose	C ₆ H ₁₂ O ₆	CH ₂ O
Benzene	C ₆ H ₆	CH
Butane	C ₄ H ₁₀	C ₂ H ₅
Water	H ₂ O	H ₂ O

Examples

❖ **To calculate empirical formula**

1. A compound contains 40% carbon, 6.7% hydrogen the rest being oxygen. Calculate the empirical formula of the compound. (C = 12, H = 1, O = 16)

Percentage of oxygen = [100 – (40 + 6.7)] % = 53.3%

Elements	C	H	O
Percentages by mass	40	6.7	53.33
Moles	$\frac{40}{12}$	$\frac{6.7}{1}$	$\frac{53.3}{16}$
	0.33	6.7	3.33
Mole ratios	$\frac{3.33}{3.33}$	$\frac{6.7}{3.33}$	$\frac{3.33}{3.33}$
	1	2	1

Hence the empirical formula of the compound is CH₂O

2. (a) A hydrated salt has the following percentage composition: iron 20.14%; sulphur 11.53%; oxygen 23.01% and water 45.32%. Find the simplest formula.
(Fe = 56, S = 32, O = 16, RFM of water = 18)

Elements	Fe	S	O	Water
Percentages	20.14	11.53	23.01	45.32
Moles	$\frac{20.14}{56}$	$\frac{11.53}{32}$	$\frac{23.01}{16}$	$\frac{45.32}{18}$
	0.3596	0.3603	1.4381	2.5178
Mole ratios	$\frac{0.3596}{0.3596}$	$\frac{0.3609}{0.3596}$	$\frac{1.4381}{0.3596}$	$\frac{2.5178}{0.3596}$
	1	1.0019	3.9991	7.0222
	1	1	4	7

Therefore the simplest formula = $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$

3. A compound W contains 3.6 g of P and 1.4 g of Q. Calculate the empirical formula of W. (P = 24, Q = 14)

Solution

Elements	P	Q
Mass	3.6	1.4
Moles	$\frac{3.6}{24}$	$\frac{1.4}{14}$
	0.15	0.1
Mole ratios	$\frac{0.15}{0.1}$	$\frac{0.1}{0.1}$
	1.5	1
	2 x 1.5	2 x 1
	3	2

The empirical formula of W is P_3Q_2

❖ **To calculate molecular formula of a compound**

The molecular formula of a compound is calculated from:

$$(i) \quad (\text{Empirical formula})_n = \text{Relative Molecular Mass}$$

$$(ii) \quad (\text{Empirical formula})_n = 2 \times \text{Vapour density}$$

Where: $\Rightarrow n = \text{a whole number } 1, 2, 3, 4, \dots$

$\Rightarrow$ Vapour density is the ratio of the mass of certain volume of the gas or vapour to the mass of the same volume of hydrogen at the same temperature and pressure.

Example

1. A gaseous hydrocarbon contains 80% carbon by mass. If the molecular mass of the gas is 30 g, find the molecular formula.

Solution:

$$\text{Percentage of H} = (100 - 80) \% = 20\%$$

Elements	C	H
Percentages	80	20
Moles	$\frac{80}{12}$	$\frac{20}{1}$
	6.7	20
Mole ratios	$\frac{6.7}{6.7}$	$\frac{20}{6.7}$
	1	3

The empirical formula is CH_3

$(\text{Empirical formula})_n = \text{Relative Molecular Mass}$

$$(\text{CH}_3)_n = 30$$

$$(12 + 1 \times 3)_n = 30$$

$$\frac{15n}{15} = 30$$

$$n = \frac{30}{15}$$

$$= 2$$

$$(\text{CH}_3)_2 = \text{C}_2\text{H}_6$$

2. A compound X contains 85.7% carbon and 14.3% hydrogen. If the vapour density of the gas is 14, calculate:
- The simplest formula of X.
 - Determine the molecular formula of X.
 - Write the name of X.
 - Write the structural formula of X.

Solution

(a)	Elements	C	H
	Percentages	85.7	14.3
	Moles	$\frac{85.7}{12}$	$\frac{14.3}{1}$
		7.14	14.3
	Mole ratios	$\frac{7.14}{7.14}$	$\frac{14.3}{7.14}$
		1	2

The empirical formula is CH₂

(b) (i) *(Empirical formula) $n = 2 \times \text{Vapour density}$*

$$\begin{aligned}
 (\text{CH}_2)_n &= 2 \times 14 \\
 [12 + (1 \times 2)] n &= 28 \\
 \frac{14n}{14} &= 28 \\
 n &= \frac{28}{14} \\
 &= 2 \\
 (\text{CH}_2)_2 &= \text{C}_2\text{H}_4
 \end{aligned}$$

The molecular formula is C₂H₄

- X is ethene.
- CH₂ = CH₂

(d) Calculations from Chemical Equations

- ❖ Write a balanced equation for the reaction or use the reaction equation written on the question paper.
- ❖ Take the numbers in front of each formula or symbol to represent moles.
- ❖ Convert the moles to masses if necessary.
- ❖ Compare the masses of the compounds concerned.

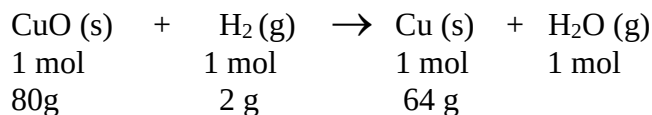
(i) To calculate mass of a reactant or product.

1. Excess hydrogen gas was passed over 20 g of heated copper (II) oxide.

Calculate:

- (a) the mass of copper formed.
- (b) the mass of hydrogen used up in the reaction.

Solution:



- (a) 80 g of CuO forms 64 g of copper.

1 g of CuO forms $\frac{64}{80}$ g of copper.

20 g of CuO forms $\frac{64}{80} \times 20 = \mathbf{16\text{ g}}$ of copper.

- (b) 80 g of CuO requires 2 g of hydrogen.

1 g of CuO requires $\frac{2}{80}$ g of hydrogen.

20 g of CuO requires $\frac{2}{80} \times 20 = \mathbf{0.5\text{ g}}$ of hydrogen

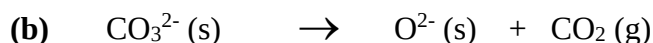
2. 5.00 g of zinc carbonate was heated to a constant mass.

- (a) State what was observed.
- (b) Write an ionic equation for the reaction that took place.
- (c) Calculate the apparent loss in mass.

Solution:

(a) The carbonate decomposed to give:

- a colourless gas that turns lime water milky and
- a solid residue which is yellow when hot and turned to white on cooling.



(c) RFM of ZnCO_3 = $65 + 12 + (3 \times 16)$
= $65 + 12 + 48$
= **125 g**

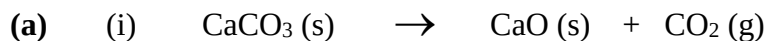
125 g of zinc carbonate gives a loss in mass of 44 g of carbon dioxide.

1 g of zinc carbonate gives a loss in mass of $\frac{44}{125}$ g of carbon.

5 g of zinc carbonate gives a loss in mass of $\frac{44}{125} \times 5 = \mathbf{1.76 \text{ g}}$

3. 5.00 g of calcium carbonate was heated until there was no change in mass.

- (a) (i) Write equation for the reaction that took place.
(ii) Calculate the mass of the solid left.
- (b) The residue was shaken with water and the product was tested with blue litmus paper. State what was observed.
(Ca = 40, C = 12, O = 16)

Solution

(ii) RFM of CaCO_3 = $40 + 12 + (3 \times 16)$
= $40 + 12 + 48$
= **100 g**

RFM of CaO = $40 + 16$
= **56 g**

100 g of calcium carbonate produces 56 g of calcium oxide.

1 g of calcium carbonate produces $\frac{56}{100}$ g of calcium oxide.

5.00 g of calcium carbonate produces $\frac{56}{100} \times 5 = \mathbf{2.8 \text{ g}}$

- (c) No observable change (The blue litmus paper remained blue because the resultant solution is basic).

(e) **Calculations involving Gas volumes**

To calculate volume of a reactant or a product (Calculations involving Gas Volumes)

Summary of data on gas volumes

1. *The volume of all gases varies with temperature and pressure.*
2. *Therefore, always the standards of temperature and pressure must be stated.
The standard chosen are:*

stp - Standard temperature (0 °C or 273 K) and pressure
760 mmHg or 1 atmosphere.

rt - room temperature.

3. *Molar gas Volume (MGV).
The volume occupied by one mole of any gas at:*

stp = 22.4 dm³ or 22,400 cm³.

rt = 24 dm³ or 24,00 cm³.

3. *Equal volume of all gases at the same temperature and pressure contain the same number of molecules (Avogadro's law).*

The following statement is also true:

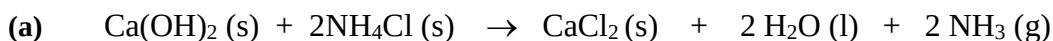
Equal numbers of molecules of different substances in the gaseous state occupy the same volume, provided all volume measurement take place under the same temperature and pressure.

4. *Number of moles of a gas* = $\frac{\text{Volume of gas}}{\text{Molar Gas Volume}}$

EXAMPLES

1. When 0.107 g of ammonium chloride was heated with calcium hydroxide, a gas was evolved.
 - (a) Write equation for the reaction.
 - (b) Calculate the volume of the gas that was evolved at room temperature. (1 mole of gas occupies 24 dm³ at rt)

Solution



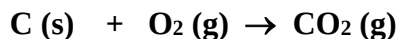
(b) Mass of NH₄Cl that reacted = 2[14 + (1 x 4) + 35.5]
= 2(14 + 4 + 35.5)
= 2 x 53.5
= **107 g**

107 g of ammonium chloride liberates (2 x 24) dm³ of ammonia.

1 g of ammonium chloride liberates $\frac{2 \times 24}{107}$ dm³ of ammonia.

0.107 g of ammonium chloride will liberate $\frac{2 \times 24}{107} \times 0.107 = \mathbf{0.048 \text{ dm}^3}$.

2. 6.0 g of carbon are burnt completely in air at s.t.p. Calculate the volume of carbon dioxide evolved. [C = 12, Molar volume = 22400 cm³ at s.t.p.]

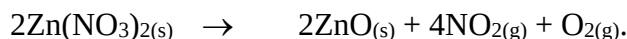


12 g of carbon liberates 22400 cm³ of carbon dioxide.

1 g of carbon liberates $\frac{22400}{12}$ cm³ of carbon dioxide.

6.0 g of carbon will liberate $\frac{22400}{12} \times 6.0 = \mathbf{11200 \text{ cm}^3}$

3. Zinc nitrate decomposes on heating according to the equation



Calculate the maximum volume of oxygen evolved in this reaction at s.t.p. when 7.56 g of zinc nitrate is heated.

[Zn = 65, N = 14, O = 16, molar gas volume at s.t.p. = 22.4 dm³]

$$\begin{aligned} \text{Mass of Zn(NO}_3)_2 \text{ that reacted} &= 2[65 + 2(14 + 3 \times 16)] \\ &= 2 [65 + 2(14 + 48)] \\ &= 2[65 + (2 \times 62)] \\ &= 2(65 + 124) \\ &= 2 \times 189 \\ &= \mathbf{378 \text{ g}} \end{aligned}$$

378 g of zinc nitrate liberates 22.4 dm³ of oxygen.

1 g of zinc nitrate liberates $\frac{22.4}{378}$ dm³ of oxygen.

7.56 g of zinc nitrate will liberate $\frac{22.4}{378} \times 7.56 = \mathbf{0.448 \text{ dm}^3}$

4. Ammonia is oxidised by copper (II) oxide according to the equation



Calculate the volume of ammonia that will be oxidised by 6.0 g of copper (II) oxide at s.t.p.? (1 mole of a gas at s.t.p. occupies 22400 cm³)

$$\begin{aligned} \text{Mass of CuO that oxidised ammonia} &= 3(64 + 16) \\ &= 3 \times 80 \\ &= \mathbf{240 \text{ g}} \end{aligned}$$

240 g of CuO oxidised 2 x 22.4 dm³ of ammonia.

1 g of CuO oxidises $\frac{2 \times 22.4}{240}$ dm³ of ammonia.

6.0 g of CuO will oxidise $\frac{2 \times 22.4}{240} \times 6 = \mathbf{1.12 \text{ dm}^3}$

GAS LAWS

Points to know:

Gas laws deal with the influence of physical factors such as:

- Temperature and
- Pressure - on the behaviour of gases.

Recall that:

The gas molecules are spread far apart compared to the molecules of liquids and solids.
The gas molecules are always in a continuous random motion (Brownian motion)

There are three gas laws, namely:

Boyle's law,
Charles law and
Pressure law.

(a) Boyle's law.

Boyle's law states that:

The volume of a fixed mass of gas is inversely proportional to the pressure at constant temperature.

Boyle's law implies that at constant temperature as the pressure of a given mass of a gas increases, the volume decreases.

Mathematically, Boyle's law is expressed as;

$$p \propto \frac{1}{V} \quad \text{or} \quad PV = \text{Constant}$$

That is when pressure of a gas changes from initial value P_1 to final value P_2 , its volume changes from initial value V_1 to final value V_2 .

In general, this is written as; $P_1V_1 = P_2V_2$

Example:

The volume of a fixed mass of gas at constant temperature is 20 cm^3 when the pressure is 75 cmHg . Find the new pressure when the volume increases to 50 cm^3 .

Solution: $P_1 = 75 \text{ cmHg}$, $P_2 = ?$, $V_1 = 20 \text{ cm}^3$, $V_2 = 50 \text{ cm}^3$

$$P_1V_1 = P_2V_2$$

$$75 \times 20 = P_2 \times 50$$

$$P_2 = \frac{75 \times 20}{50}$$

$$\therefore P_2 = 30 \text{ cmHg}$$

(b) Charles's law

Charles's law states that:

The volume of a fixed mass of gas at constant pressure is directly proportional to the absolute temperature.

Where the absolute temperature is the temperature at which a gas occupies zero volume.

Charles law implies that the volume of a fixed mass of a gas increases or decreases in the same proportion as the absolute temperature provided the pressure is kept constant.

Mathematically, Charles law is expressed as:

$$V \propto T \quad \text{or} \quad \frac{V}{T} = \text{Constant}$$

That is at a constant pressure, when the volume of a gas changes from V_1 to V_2 , the temperature changes from T_1 to T_2 .

In general, this is written as;
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Example

The volume of a fixed mass of gas at constant pressure is 50 cm^3 at a temperature of 27°C . What will the volume be at 42°C .

Solution: $T_1 = 27^\circ\text{C} = 273 + 27 = 300 \text{ K}$, $T_2 = 42^\circ\text{C} = 273 + 42 = 315 \text{ K}$

$V_1 = 50 \text{ cm}^3$, $V_2 = ?$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow \frac{50}{300} = \frac{V_2}{315} \Rightarrow V_2 = \frac{50 \times 315}{300}$$
$$\therefore V_2 = 52.5 \text{ cm}^3$$

(c) Pressure Law:

Pressure Law states that:

The pressure of a fixed mass of a gas is directly proportional to its absolute temperature if the volume is kept constant.

Mathematically, pressure law is expressed as:

$$P \propto T \quad \text{or} \quad \frac{P}{T} = \text{Constant}$$

That is at a constant volume, when the pressure of a gas changes from initial value P_1 to final value P_2 , the temperature changes from initial value T_1 to final value T_2 .

In general, this is written as; $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Example:

The pressure of a given mass of gas at constant volume at a temperature of 100 °C is 300 mmHg. What will the pressure be if the temperature falls to 27 °C.

Solution: $T_1 = 273 + 100 = 373\text{K}, \quad T_2 = 273 + 27 = 300\text{K}$
 $P_1 = 300 \text{ mm Hg}, \quad P_2 = ?$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \Rightarrow \frac{300}{373} = \frac{P_2}{300}$$

$$P_2 = \frac{300 \times 300}{373}$$

$$P_2 = 241.3 \text{ mmHg}$$

Equation of state or Ideal gas equation

The three equations can be combined into one equation.

If a fixed mass of a gas of volume V_1 exerts a pressure P_1 at absolute temperature T_1 , the

expression may be written as: $\frac{P_1 V_1}{T_1} = \text{Constant.} \dots\dots\dots 1$

Now suppose the same mass of a gas has a volume V_2 and exerts a pressure P_2 at absolute temperature T_2 , then new expression becomes:

$$\frac{P_2 V_2}{T_2} = \text{Constant.} \dots\dots\dots 2$$

Equating equation (1) to equation (2), we have the equation of state expressed as:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

❖ *This equation holds for ideal gases under all conditions.*

Examples

1. Air in a 2.5 litre vessel at 127 °C exerts a pressure of 3 atm. Calculate the pressure that the same mass of air would exert if contained in a 4 litre vessel at – 43 °C.

Solution: $T_1 = 273 + 127 = 400\text{K}, \quad T_2 = 273 - 43 = 230\text{K}$
 $P_1 = 3 \text{ atm}, \quad P_2 = ? \text{ atm}$
 $V_1 = 2.5 \text{ dm}^3 \quad V_2 = 4 \text{ dm}^3$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \Rightarrow \frac{3 \times 2.5}{400} = \frac{P_2 \times 4}{230}$$

$$P_2 = \frac{3 \times 2.5 \times 230}{400 \times 4}$$

$$P_2 = 1.078 \text{ atm}$$

2. The volume of a gas is 90 cm³ at 25 °C and 780 mmHg pressure. Calculate the volume of the gas at s.t.p.

Solution: $T_1 = 25^\circ\text{C} = 25 + 273 = 298\text{K}$, $T_2 = 0^\circ\text{C} = 273\text{K}$ (at stp)
 $P_1 = 780 \text{ mmHg}$, $P_2 = 760 \text{ mmHg}$ (at stp)

$V_1 = 90 \text{ cm}^3$, $V_2 = ?$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \Rightarrow \frac{780 \times 90}{298} = \frac{760 \times V_2}{273}$$

$$V_2 = \frac{780 \times 90 \times 273}{298 \times 760}$$

$$\therefore V_2 = 84.6 \text{ cm}^3$$

Graham's law of diffusion of gases

Graham's law of diffusion of gases states that:

Under the same conditions of temperature and pressure, the rate of diffusion of a gas is inversely proportional to the square root of its density.

The following formulae are used when using Graham's law.

$$1. \quad \frac{\text{Rate of diffusion of gas A}}{\text{Rate of diffusion of gas B}} = \sqrt{\frac{\text{Molecular mass of gas B}}{\text{Molecular mass of gas A}}}$$

$$2. \quad \frac{\text{Time taken by gas A to diffuse}}{\text{Time taken by gas B to diffuse}} = \sqrt{\frac{\text{Molecular mass of gas B}}{\text{Molecular mass of gas A}}}$$

Where: A and B are gases of equal volumes.

Examples

1. Equal volumes of carbon monoxide and carbon dioxide are allowed to diffuse through the same medium. Calculate the relative rate of diffusion of carbon monoxide. (C = 12, O = 16).

Solution: Molecular mass of CO = 12 + 16
 $= 28 \text{ g}$
 Molecular mass of CO₂ = 12 + (2 x 16)
 $= 44 \text{ g}$

$$\begin{aligned}\frac{\text{Rate of diffusion of CO}}{\text{Rate of diffusion of CO}_2} &= \sqrt{\frac{\text{Molecular mass of CO}_2}{\text{Molecular mass of CO}}} \\ &= \sqrt{\frac{44}{28}} \\ &= 1.254\end{aligned}$$

$\therefore$ Carbon monoxide diffuses 1.254 times faster than carbon dioxide.

2. If it takes 20 s for 200 cm³ of oxygen gas to diffuse across a porous material. How long will it take an equal volume of sulphur dioxide to diffuse across the same porous material? (S = 32, O = 16)

Solution:

Molecular mass of O₂ = 2 x 16 = 32 g, Time taken by O₂ = 20 s
Molecular mass of SO₂ = 32 + (2 x 16) = 64 g, Time taken by SO₂ = ?

$$\begin{aligned}\frac{\text{Time taken by O}_2 \text{ to diffuse}}{\text{Time taken by SO}_2 \text{ to diffuse}} &= \sqrt{\frac{\text{Molecular mass of O}_2}{\text{Molecular mass of SO}_2}} \\ \frac{20}{\text{Time taken by SO}_2 \text{ to diffuse}} &= \sqrt{\frac{32}{64}} \\ &= \sqrt{\frac{1}{2}} \\ &= \frac{\sqrt{1}}{\sqrt{2}} \\ &= \frac{1}{\sqrt{2}} \\ \text{Time taken by SO}_2 \text{ to diffuse} &= 20\sqrt{2} \\ \therefore \text{Time taken by SO}_2 \text{ to diffuse} &= 28.3 \text{ s}\end{aligned}$$

3. Determine the molecular mass of gas Y when oxygen diffuses 1½ times faster than gas Y. (O = 16)

Solution: Rate of diffusion of O₂ = 1½, Molecular mass of O₂ = 32 g
Rate of diffusion of Y = 1 Molecular mass of Y = ?

$$\begin{aligned}\frac{\text{Rate of diffusion of gas Y}}{\text{Rate of diffusion of gas O}_2} &= \sqrt{\frac{\text{Molecular mass of O}_2}{\text{Molecular mass of Y}}} \\ \frac{1}{1\frac{1}{2}} &= \sqrt{\frac{32}{\text{Molecular mass of Y}}} \\ \frac{1^2}{1.5^2} &= \frac{32}{Y} \\ Y &= 1.5^2 \times 32 \\ &= 2.25 \times 32\end{aligned}$$

$\therefore$ The molecular mass of Y = 72 g

4. If it takes 30 seconds for 100 cm³ of carbon dioxide to diffuse across a porous plate. How long will it take 150 cm³ of nitrogen dioxide to diffuse across the same plate under similar conditions? (C = 12, N = 14, O = 16)

Solution:

Volume of CO ₂	= 100 cm ³ ,	Volume of NO ₂	= 150 cm ³
Molecular mass of CO ₂	= 44 g,	Molecular mass of NO ₂	= 46 g
Time taken by CO ₂	= 30 s,	Time taken by NO ₂	= ?
Rate of diffusion of CO ₂	= $\frac{v}{t} = \frac{100}{30} = 3.33 \text{ cm}^3\text{s}^{-1}$		

$$\frac{\text{Rate of diffusion of CO}_2}{\text{Rate of diffusion of NO}_2} = \sqrt{\frac{\text{Molecular mass of NO}_2}{\text{Molecular mass of CO}_2}}$$

$$\frac{3.33}{\text{Rate of diffusion of NO}_2} = \sqrt{\frac{46}{44}}$$

$$\frac{3.33}{\text{Rate of diffusion of NO}_2} = 1.0225$$

$$\begin{aligned}\text{Rate of diffusion of NO}_2 &= \frac{3.33}{1.0225} \\ &= 3.26\end{aligned}$$

$$\therefore \text{Rate of diffusion of NO}_2 = 3.26 \text{ cm}^3\text{s}^{-1}$$

Now the time taken by NO₂ to diffuse is calculated from the formula:

$$\text{Rate of diffusion of NO}_2 = \frac{\text{Volume of gas}}{\text{Time taken}}$$

$$3.26 \text{ cm}^3\text{s}^{-1} = \frac{150 \text{ cm}^3}{\text{Time taken}}$$

$$\therefore \text{Time taken} = \frac{150 \text{ cm}^3}{3.26 \text{ cm}^3\text{s}^{-1}} = 46 \text{ s}$$

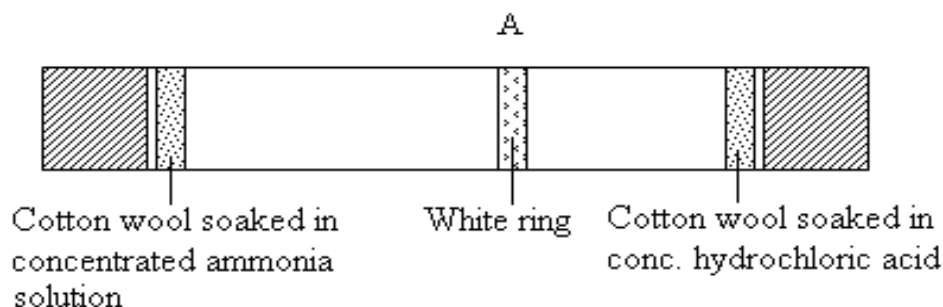
Points to recall:

- ❖ Diffusion is the “spreading” of molecules from a region of high concentration to a region of low concentration through a medium.
- ❖ Diffusion occurs faster in gases than in liquids.
- ❖ Gases with low densities diffuse faster than those with higher densities.

Examples

1. (1990 Q.6)

- (a) Smoke was put in a glass-cell and viewed under a microscope.
 (i) State what was observed.
 (ii) Explain the observation in (i).
- (b) One piece of cotton wool was soaked in concentrated ammonia solution and another in concentrated hydrochloric acid. The two pieces of cotton wool were placed in a glass tube as shown in figure below.



- (i) Write the formula of the substance that formed the white ring.
 (ii) Explain why the white ring is formed in position A and not in the middle of the tube.

Solution:

- (a) (i) *Smoke particles were seen moving randomly in the glass cell.*
 (ii) *The smoke particles collided with air molecules, hence the random motion.*
- (b) (i) NH_4Cl
 (ii) *Ammonia is less dense than hydrogen chloride gas and therefore diffuses faster than hydrogen chloride hence the formation of the white ring near the cotton wool soaked in concentrated hydrochloric acid.*

2. (1998 Q.3)

A crystal of potassium manganate (VII) was placed at the corner in a trough of water as shown in figure 1 below and the experiment was allowed to stand for about 30 minutes.

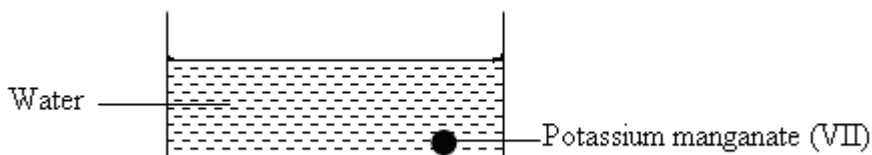


Fig. 1

- (a) State what was observed after 30 minutes.
 (b) Name the process that occurred.
 (c) State the purpose of the experiment.

Solution

- (a) *Purple colour of potassium manganate (VII) spread through out the liquid.*
 (b) *Diffusion.*
 (c) *To show that water molecules are in a constant random motion. And so collide with the particles of potassium manganate (VII).*

The Mole Concept and Volumetric Analysis

Stoichiometry:

- Deals with ratios in which substances combine with one another when chemical reactions occur.
- It also deals with composition of chemical substances i.e. ratios by mass in which elements are combined in a given compound.
- The fundamental unit of *Stoichiometry* is the *mole*.

1. Volumetric Analysis:

Volumetric analysis is the analytical process to determine quantities of compounds (usually acids and alkalis) which involves measurement of volumes of solutions, using pipette, burette and (for approximate measurements) measuring cylinder.

(a) Uses of Volumetric Analysis:

The object of volumetric analysis is to:

- (i) Measure concentration of a given compound either in g/l or mol/l.
- (ii) Determine the relative atomic mass of an element in a compound.
- (iii) Determine the number of moles of water of crystallization in a hydrated compound.
- (iv) Determine the percentage purity or impurity of a compound in an impure sample.
- (v) Determine the basicity of an acid.
- (vi) Determine the reaction ratio.

(b) Terms Used.

(i) Concentration:

Concentration is a measure of the amount of a substance dissolved in a known volume of a solution.

Units: It is measured in:

- mol/dm³ (moles per cubic decimeter)
- mol/l (moles per litre)
- g/dm³ (grams per cubic decimeter)
- g/l (grams per litre)
- g/cm³ (grams per cubic centimeter)

Note: 1 dm³ = 1 litre
1 dm³ = 1,000 cm³
1 l = 1,000 cm³

(ii) Standard Solution:

Standard solution is a solution of known concentration. Or - is one for which the concentration is known.

(iii) Molar Solution:

Molar Solution is one, which contains one mole of a substance in a litre or a cubic decimeter.

(iv) Molarity:

Molarity is the number of moles per litre.

Or *Molarity is the number of moles per cubic decimeter.*

(v) Molar Mass or Relative Formula Mass (RFM)

Molar mass is the mass of one mole of a compound.

RFM = the sum of the relative atomic masses of the elements present in the compound.

The Table below shows the RFMs of the common compounds used in Volumetric analysis.

(H = 1, Na = 23, Cl = 35.5, O = 16, C = 12)

Compound	Name of Compound	Chemical formula	Molar Mass(RFM)/g
Acid	Hydrochloric Acid	HCl	36.5
	Sulphuric Acid	H ₂ SO ₄	98
	Anhydrous Oxalic Acid	H ₂ C ₂ O ₄	90
	Hydrated Oxalic Acid	(COOH) ₂ H ₂ C ₂ O ₄ .2H ₂ O (COOH) ₂ .2H ₂ O	126
Base	Sodium Hydroxide	NaOH	40
	Anhydrous sodium carbonate	Na ₂ CO ₃	106
	Hydrated Sodium Carbonate	Na ₂ CO ₃ .10H ₂ O	286

(b) Relationship Between Molarity, Concentration and RFM

$$\text{Molarity} = \frac{\text{Concentration in g/l}}{\text{RFM}}$$

$$\text{Or Molarity} = \frac{\text{Concentration in g/dm}^3}{\text{RFM}}$$

2. (a) Titration:

Titration is the estimation of concentration of an acid or alkali in solution by reacting it with a standard alkali or acid respectively.

- In the analysis, two aqueous solutions are used.
- The concentration of one solution is always known and the other unknown.

(b) Titration Process:

- The titration process involves running one solution from burette on to the other solution of a fixed volume (usually 20 or 25 cm³) in which 2-3 drops of indicator are added in a conical flask until the two have just reacted completely.
- This is noticed when the colour of the indicator just changes.. This shows that the neutral point (end point) of the titration has been reached and the volume of the solution used from the burette is accurately noted.
- The volume used from the burette is calculated by using the formula:
 $\text{Volume used} = (\text{Final burette reading}) - (\text{Initial burette reading})$
- The value is always written in two places of decimal.
- The procedure is repeated 3 - 4 times and the values are entered in a table usually drawn in advance on the question paper.

3. Indicators:

An indicator is substance that has different colour in acid and alkali solutions.

The common indicators used are:

- Phenolphthalein indicator.
- Methyl Orange.

The table below shows the colours of the indicators in the laboratory in different media.

Indicator	Neutral	Colour in	
		Acid	Alkali
Phenolphthalein	Colourless	Colourless	Red
Methyl Orange	Orange	Red	Yellow

4. CALCULATIONS:

Steps followed in the calculations:

The steps followed in the calculations involve the following:

- (i) Write the volume of the pipette used in two places of decimal in the space usually provided on the question paper.
- (ii) Find the average volume, using the consistent results, of the solution used from the burette.
- (iii) Write a balanced reaction equation or ionic equation for the reaction that took place. Or use the written equation of the reaction on the question paper. However, sometimes the equations of the reaction is written on the question paper. In such a case, refer to the equation.
- (iv) From the balanced equation, extract the mole ratio and express it in full.
e.g. 1 mole of acid : 2 mole of the base or vice versa.
- (v) Always start the calculation with the solution of known concentration and volume.
- (vi) Calculate the number of moles of the other solution using the mole ratio.

NB: Do NOT use the "titration equation".

$$\frac{M_a V_a}{n_a} = \frac{M_b V_b}{n_b} \quad \text{Or} \quad \frac{M_a V_a}{M_b V_b} = \frac{n_a}{n_b}$$

Where:

M_a and M_b	=	the molarities of the acid and the base respectively.
V_a and V_b	=	the volumes of the acid and the base respectively.
n_a and n_b	=	the number of moles of the acid and the base respectively.

EXAMPLES

1. Determination of Molarity of a Solution (Concentration mol/l)

You are provided with **BA1** and **BA2** solutions.

BA1 is a solution of Hydrochloric acid.

BA2 is a solution of anhydrous sodium carbonate that contains 7.2g/dm³.

(Na = 23, C = 12, O = 16)

You are required to determine the molarity of **BA1**.

Procedure:

Pipette 25 cm³ or 20 cm³ of **BA2** into a conical flask. Add two drops of the indicator and titrate it with **BA1** from burette. Repeat the procedure until you get consistent results.

Enter the results in the table below.

Sample results

Volume of pipette used **20.00**. cm³

Burette Readings	1	2	3	4
Final Burette Reading (cm ³)	33.00	31.80	32.30	30.80
Initial Burette Reading (cm ³)	1.50	2.00	2.50	1.00
Volume of BA1 used (cm ³)	31.50	29.80	29.80	29.80

Titre values used to find the average volume = 29.80, 29.80, 29.80

$$\begin{aligned}
 \text{Average volume BA1 used} &= \frac{29.80 + 29.80 + 29.80}{3} \\
 &= \frac{89.4}{3} \\
 &= \mathbf{29.80 \text{ cm}^3}
 \end{aligned}$$

Questions:

(a) Write the equation for the reaction.



Mole ratio: 2 mol of acid : 1 mol of base

(b) Calculate:

(i) the molarity of BA2.

$$\begin{aligned}
 \text{Concentration} = 7.2 \text{ g/l, RFM of Na}_2\text{CO}_3 &= (2 \times 23) + 12 + (3 \times 16) \\
 &= \mathbf{106 \text{ g}}
 \end{aligned}$$

$$\begin{aligned}
 \text{Molarity} &= \frac{\text{Concentration in g/l}}{\text{RFM}} \\
 &= \frac{7.2}{106} \\
 &= \mathbf{0.07M}
 \end{aligned}$$

(ii) *The number of moles of **BA2** that reacted with BA1.*

Solution: Molarity = 0.07M, Vol = 20 cm³, Number of moles = ?

0.07 moles of **BA2** are contained in 1000 cm³ of solution.
 The number of moles contained in 20 cm³ of **BA2** solution

$$\begin{aligned}
 &= \frac{0.07 \times 20}{1000} \\
 &= \mathbf{0.0014 \text{ moles}}
 \end{aligned}$$

(iii) *The number of moles of **BA1** that reacted with **BA2**.*

From the mole ratio, 2 moles of acid : 1 mole of base, the number of moles of BA1 that reacted with BA2

$$\begin{aligned}
 &= 2 \times \text{the number of moles of BA2} \\
 &= 2 \times 0.0014 \\
 &= \mathbf{0.0028 \text{ moles}}
 \end{aligned}$$

(iv) *The molarity of **BA1**.*

Solution: Molarity = ?, Number of moles = 0.0028, Vol = 29.80 cm³

0.0028 moles of BA1 are contained in 29.80 cm³ of solution.
 The number of moles contained in 1000 cm³ of BA1 solution

$$\begin{aligned}
 &= \frac{0.0028 \times 1000}{29.80} \\
 &= 0.09 \\
 &= \mathbf{0.1M}
 \end{aligned}$$

2. Determination of Atomic mass of an Element in a compound

You are provided with **BA1** and **BA2** solutions.

BA1 is a solution 0.1M Hydrochloric acid.

BA2 is a solution obtained by dissolving 5.3 g of a metal carbonate, M_2CO_3 , per litre.

(C = 12, O = 16)

You are required to determine the atomic mass of the element **M** in the carbonate.

Procedure:

Pipette 25 cm³ or 20 cm³ of **BA2** into a conical flask. Add two drops of methyl orange indicator and titrate it with **BA1** from burette. Repeat the procedure until you get consistent results.

Enter the results in the table below.

Sample results

Volume of pipette used **20.00**. cm³

Burette Readings	1	2	3	4
Final Burette Reading (cm ³)	21.70	30.20	21.90	31.30
Initial Burette Reading (cm ³)	1.50	21.70	2.50	21.90
Volume of BA1 used (cm ³)	20.20	19.50	19.40	19.40

Titre Values used to find the average volume= 19.50, 19.40, 19.40

$$\begin{aligned}\text{Average volume of BA1 used} &= \frac{19.50 + 19.40 + 19.40}{3} \\ &= \frac{58.3}{3} \\ &= \mathbf{19.40 \text{ cm}^3}\end{aligned}$$

Questions:

(a) Write the equation for the reaction.



(b) Calculate:

(i) The number of moles of **BA1** that reacted with **BA2**.

$$\begin{aligned}\text{Volume of BA1 (acid)} &= \text{average volume} \\ &= 19.40 \text{ cm}^3\end{aligned}$$

Molarity of BA1 = 0.1M, Number of moles = ?
0.1 moles of acid are contained in 1000 cm³ of solution.

The number of moles contained 19.40 cm³ of solution

$$= \frac{19.40 \times 0.1}{1000}$$

$$= \mathbf{0.00194}$$

(ii) The number of moles of BA2 that reacted with BA1.

From the mole ratio: 2 moles of acid : 1 mole of base

The number of moles of M₂CO₃ that reacted with BA1

$$= \frac{1}{2} \text{ of number of moles of the acid}$$

$$= \frac{1}{2} \times 0.00194$$

$$= 0.00097$$

$$= \mathbf{0.001}$$

(iii) The molarity of **BA2**.

Volume of BA2 (Base) = 20.00 cm³, Number of moles = 0.001,

Molarity of BA2 = ?,

0.001 moles of BA2 are contained in 20 cm³ of solution.

The number of moles contained 1000 cm³ of solution

$$= \frac{0.001 \times 1000}{20}$$

$$= \mathbf{0.05 \text{ M}}$$

(c) Determine:

(i) The formula mass of **BA2**.

Solution: Molarity of BA1 = 0.05M, Concentration = 5.3 g/l, RFM = ?

$$\text{Molarity} = \frac{\text{Concentration in g/l}}{\text{RFM}}$$

$$\text{RFM} = \frac{\text{Concentration in g/l}}{\text{Molarity}}$$

$$= \frac{5.3}{0.05}$$

$$= \mathbf{106}$$

(ii) The atomic mass of M .

$$2M + 12 + 3 \times 16 = 106$$

$$2M + 12 + 48 = 106$$

$$2M + 60 = 106$$

$$2M = 106 - 60$$

$$2M = 46$$

$$M = \frac{46}{2}$$

$$\therefore M = 23$$

3. Determination of the number of moles of water of crystallization in a hydrated compound

Hydrated compounds are compounds that contain water of crystallization.

The hydrated compounds used for this purpose are:

- $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ and

- $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$

In the calculations, the figures **2** and **10** and the symbols **Na** and **C₂O₄** in the above compounds are replaced by a letters. The common letters used for the figures are **x** and **n**.

Let us consider the following examples.

I. You are provided with **BA1** and **BA2** solutions.

BA1 is a solution of ethanoic acid, $\text{H}_2\text{X} \cdot n\text{H}_2\text{O}$ made by dissolving 12.6 g of the acid per litre.

BA2 is 0.2M solution of sodium hydroxide solution.

You are required to determine the value of **n** in the formula of the acid.

Procedure:

Pipette 25cm^3 or 20cm^3 of **BA1** into a conical flask. Add two drops of phenolphthalein indicator and titrate it with **BA2** from burette. Repeat the procedure until you get consistent results.

Enter the results in the table below.

Sample results

Volume of pipette used **20.00**. cm³

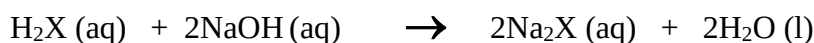
Burette Readings	1	2	3	4
Final Burette Reading (cm ³)	22.00	42.10	22.50	42.50
Initial Burette Reading (cm ³)	1.50	22.00	2.50	22.50
Volume of BA2 used (cm ³)	20.20	19.10	19.00	18.90

Titre values used to find the average volume= 19.10, 19.00, 18.90 cm³

$$\begin{aligned}\text{Average volume of BA2 used} &= \frac{19.10 + 19.00 + 18.90}{3} \\ &= \frac{57}{3} \\ &= \mathbf{19.00 \text{ cm}^3}\end{aligned}$$

Questions:

(a) Write the equation for the reaction.



(b) Calculate:

(i) The number of moles of **BA2** that reacted with **BA1**.

Volume of BA2 (base) = average volume = 19.00 cm³,

Molarity of BA2 = 0.2M, Number of moles = ?

0.2 moles of acid are contained in 1000 cm³ of solution.

The number of moles of BA2 contained 19.00 cm³ of solution

$$\begin{aligned}&= \frac{19 \times 0.2}{1000} \\ &= \mathbf{0.0038}\end{aligned}$$

(ii) The number of moles of **BA1** that reacted with **BA2**.

From the mole ratio: 2 moles of base : 1 mole of acid

$$\begin{aligned}0.004 \text{ moles react with} &= \frac{1}{2} \text{ of the number of moles of the BA1} \\ &= \frac{1}{2} \times 0.0038 \\ &= \mathbf{0.0019}\end{aligned}$$

(c) Determine:

(i) The molarity of **BA1**.

Volume of BA1 (acid) = 20.00 cm³, Molarity of BA1 =?,
Number of moles = 0.0019.

20.00 cm³ of BA1 solution contains 0.0019 moles,

$$1000 \text{ cm}^3 \text{ of solution contains } \frac{0.0019 \times 1000}{20} = 0.095$$

$$= 0.1 \text{ moles}$$

Therefore, the molarity of BA1 = **0.1M**.

(ii) The RFM of **BA1**.

$$\text{Molarity} = \frac{\text{Concentration in g/l}}{\text{RFM}}$$

$$\text{RFM} = \frac{\text{Concentration in g/l}}{\text{Molarity}}$$

$$= \frac{12.6}{0.1}$$

$$\therefore \text{RFM} = 126$$

(iii) The value of **n** (Given that the mass of H₂X = 90)

$$90 + 18n = 126$$

$$18n = 126 - 90$$

$$18n = 36$$

$$n = \frac{36}{18}$$

$$\therefore n = 2$$

II. You are provided with **BA1** and **BA2** solutions.

BA1 is a solution of 0.2M solution of Hydrochloric acid.

BA2 is a solution of a metal hydrated carbonate $M_2CO_3 \cdot nH_2O$, made by dissolving 4g of the hydrated carbonate in 250 cm³ of solution.

You are required to determine the value of **n** in the formula of the hydrated compound.

Procedure:

Pipette 25cm³ or 20 cm³ of **BA2** into a conical flask. Add two drops of methyl orange indicator and titrate it with **BA1** from burette. Repeat the procedure until you get consistent results.

Enter the results in the table below.

Sample results

Volume of pipette used **20.00**. cm³

Burette Readings	1	2	3	4
Final Burette Reading (cm ³)	13.50	24.70	35.90	13.10
Initial Burette Reading (cm ³)	1.00	13.50	24.70	2.00
Volume of BA1 used (cm ³)	12.50	11.20	11.20	11.10

Titre values used to find the average volume = 11.20, 11.20, 11.10 cm³.

$$\begin{aligned}\text{Average volume of BA1 used} &= \frac{11.20 + 11.20 + 11.20}{3} \\ &= \frac{33.60}{3} \\ &= \mathbf{11.20 \text{ cm}^3}\end{aligned}$$

Questions:

(a) Write the equation for the reaction.



(b) Calculate:

(i) The number of moles of **BA1** that reacted with **BA2**.

Volume of BA1 (acid) = average volume = 11.20 cm³

Molarity of BA1 = 0.2M, Number of moles = ?

0.2 moles of acid are contained in 1000 cm³ of solution.

Then 11.20 cm³ of BA1 solution contains $\frac{11.20 \times 0.2}{1000}$

= 0.00224 moles

(ii) The number of moles of **BA2** that reacted with **BA1**.

From the mole ratio: 2 moles of BA1 (acid): 1 BA2 mole of (base).

0.00224 moles of BA2 require $\frac{1}{2}$ the number of moles of the BA1.

$$= \frac{1}{2} \times 0.00224$$

$$= \mathbf{0.00112 \text{ moles}}$$

(c) Determine:

(i) The molarity of **BA2**.

Volume of BA2 (base) = 20.00 cm³, Number of moles = 0.00112,
Molarity of BA2 =?,

20.00 cm³ of BA2 solution contains 0.00112 moles.

$$1000 \text{ cm}^3 \text{ of solution contains } \frac{0.00112 \times 1000}{20} = 0.056$$

$$= 0.056 \text{ moles}$$

$$\text{Therefore, the molarity of BA2} = \mathbf{0.056M.}$$

(ii) The RFM of **BA2**.

250 cm³ of solution contains 4 g, then the mass contained in 1000 cm³

$$= \frac{4 \times 1000}{250}$$

$$= \mathbf{16 \text{ g/l}}$$

$$\text{From Molarity} = \frac{\text{Concentration in g/l}}{\text{RFM}}$$

$$\text{RFM} = \frac{\text{Concentration in g/l}}{\text{Molarity}}$$

$$= \frac{16}{0.056}$$

$$= 285.7$$

$$\therefore \text{RFM} = \mathbf{286 \text{ g}}$$

(iii) The value of n (Given that the mass of $M_2CO_3 = 106$)

$$106 + 18n = 286$$

$$18n = 286 - 106$$

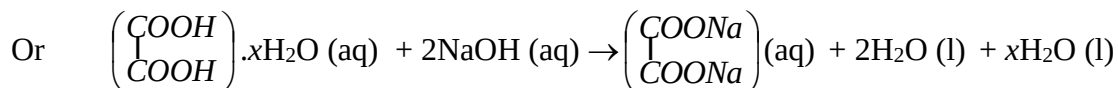
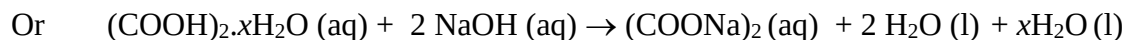
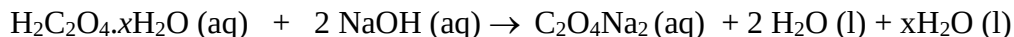
$$18n = 180$$

$$n = \frac{180}{18}$$

$$\therefore n = \mathbf{10}$$

III. A solution of ethanedioic acid was made by dissolving 10g of acid crystals, $\text{H}_2\text{C}_2\text{O}_4 \cdot x\text{H}_2\text{O}$ in water and the solution made up to 500 cm^3 . 25.0 cm^3 of the acid solution required 15.90 cm^3 of 0.5M sodium hydroxide solution for complete reaction.

(a) Write the equation for the reaction.



Mole ratio is 1 mole of acid: 2 moles of Sodium Hydroxide.

(b) Calculate:

(i) the concentration of acid in mol dm^{-3} .

Acid: Volume = 25.0 cm^3 , Conc. = $10\text{g}/500 \text{ cm}^3$ of water,
Molarity = ? No. of moles = ?

Base: Vol = 15.90 cm^3 , Molarity = 0.5M ; No. of moles = ?

Moles of Sodium hydroxide:

1000 cm^3 of sodium hydroxide solution contains 0.5 moles.

1 cm^3 of sodium hydroxide solution contains $\frac{0.5}{1000}$ moles.

15.90 cm^3 of sodium hydroxide solution contains $\frac{0.5 \times 15.90}{1000}$

= 0.00795 moles

Moles of acid:

From the mole ratio, 1 mole of acid: 2 moles of Sodium Hydroxide,
the number of moles of the acid is equal to half the number of moles of sodium hydroxide.

= $\frac{1}{2} \times 0.00795$

= 0.003975 moles

Molarity of the acid

25.0 cm³ of the acid solution contains 0.003975 moles.

1 cm³ of acid solution contains $\frac{0.003975}{25}$ moles

1000 cm³ of the acid solution contains $\frac{0.003975 \times 1000}{25} = \mathbf{0.159 \text{ mol dm}^{-3}}$

(ii) *The number of molecules of water of crystallization in the acid.*

Data: Molarity of acid = 0.159M, Conc. = 10g/500cm³(20g/l), RFM = ?

Relative Formula Mass of the Acid

$$\begin{aligned}
 \text{From Molarity} &= \frac{\text{Concentration in g/l}}{\text{RFM}} \\
 \text{RFM} &= \frac{\text{Concentration in g/l}}{\text{Molarity}} \\
 &= \frac{20}{0.159} \\
 &= 125.786 \\
 &= \mathbf{126 \text{ g}}
 \end{aligned}$$

The value of x:

$$\begin{aligned}
 \text{H}_2\text{C}_2\text{O}_4 \cdot x\text{H}_2\text{O} &= 126 \\
 (1 \times 2) + (2 \times 12) + (4 \times 16) + 18x &= 126 \\
 2 + 24 + 64 + 18x &= 126 \\
 90 + 18x &= 126 \\
 18x &= 126 - 90 \\
 \frac{18x}{18} &= \frac{36}{18} \\
 \mathbf{x} &= \mathbf{2}
 \end{aligned}$$

4. Determination of the percentage purity of a compound in an impure sample

You are provided with **BA1** and **BA2** solutions.

BA1 is a solution of 0.1M solution of Hydrochloric acid.

BA2 is a solution of an impure sample of anhydrous sodium carbonate, Na_2CO_3 , made by dissolving 6.2g of the carbonate per litre.

You are required to determine the percentage purity of the compound.

Procedure:

Pipette 25cm^3 or 20cm^3 of **BA2** into a conical flask. Add two drops of methyl orange indicator and titrate it with **BA1** from burette. Repeat the procedure until you get consistent results.

Enter the results in the table below.

Sample results

Volume of pipette used **20.00**. cm^3

Burette Readings	1	2	3	4
Final Burette Reading (cm^3)	13.50	24.70	35.90	13.10
Initial Burette Reading (cm^3)	1.00	13.50	24.70	2.00
Volume of BA1 used (cm^3)	20.40	19.00	19.00	19.00

Titre values used to find the average volume = 19.00, 19.00, 19.00 cm^3

$$\begin{aligned}\text{Average volume of BA2 used} &= \frac{19.00 + 19.00 + 19.00}{3} \\ &= \frac{57}{3} \\ &= \mathbf{19.00\text{ cm}^3}\end{aligned}$$

Questions:

(a) Write the equation for the reaction.



(b) Calculate:

(i) The number of moles of **BA1** that reacted with **BA2**.

Molarity = 0.1M, Vol. = 19.00 cm³, Number of moles = ?

0.1 moles of BA1 are contained in 1000 cm³ of solution.

The number of moles of BA1 contained 19.00 cm³ of solution

$$= \frac{19 \times 0.1}{1000}$$

$$= \mathbf{0.0019}$$

(ii) The number of moles of **BA2** that reacted with **BA1**.

From the mole ratio, 2 mol of acid: 1 mol of base,

the number of **BA2** hat reacted with BA1

$$= \frac{1}{2} \text{ the number of moles of BA1}$$

$$= \frac{1}{2} \times 0.0019$$

$$= \mathbf{0.00095 \text{ moles}}$$

(c) Determine:

(i) The molarity of BA2.

Solution:

Molarity = ?, Number of moles = 0.0095, Vol = 20.00 cm³

20 cm³ of BA2 solution contains 0.0095 moles.

$$1000 \text{ cm}^3 \text{ of BA2 contains } \frac{0.00095 \times 1000}{20} = 0.0475$$

$$= \mathbf{0.05M}$$

(ii) The concentration of **BA2** in g/l. (Na = 23, C =12, O = 16)

Molarity = 0.05M, RFM = 106, Concentration in g/l = ?

$$\text{Molarity} = \frac{\text{Concentration in g/l}}{\text{RFM}}$$

$$0.05 = \frac{\text{Concentration}}{106}$$

$$\text{Concentration} = 0.05 \times 106$$

$$\therefore \text{Concentration} = \mathbf{5.3 \text{ g/l}}$$

(iii) The percentage purity.

$$\begin{aligned}\text{Percentage purity} &= \frac{\text{Concentration in g/l of pure salt}}{\text{Concentration in g/l of impure salt}} \times 100 \\ &= \frac{5.3}{6.2}\end{aligned}$$

$$\therefore \text{Percentage purity} = 85.48\%$$

5. Determination of basicity of an acid

You are provided with **BA1** and **BA2** solutions.

BA1 is a solution of sodium hydroxide obtained by dissolving 8 g of a metal carbonate M_2CO_3 per litre. (Na = 23, C = 12 O = 16)

BA2 is a solution 0.1M acid of formula H_nX .

You are required to determine the basicity of the acid.

Procedure:

Pipette 25cm^3 or 20cm^3 of **BA2** into a conical flask. Add two drops of phenolphthaein indicator and titrate it with **BA1** from burette. Repeat the procedure until you get consistent results. Enter the results in the table below.

Sample results

Volume of pipette used **20.00**. cm^3

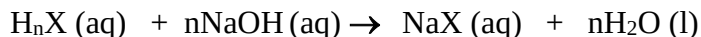
Burette Readings	1	2	3	4
Final Burette Reading (cm^3)	22.00	42.10	22.50	42.50
Initial Burette Reading (cm^3)	1.50	22.00	2.50	22.50
Volume of BA1 used (cm^3)	21.50	20.10	20.00	19.90

Titre Values used to find the average volume= 20.10, 20.00, 19.90 cm^3

$$\begin{aligned}\text{Average volume of BA1 used} &= \frac{20.10 + 20.00 + 19.90}{3} \\ &= \frac{60.00}{3} \\ &= 20.00\text{ cm}^3\end{aligned}$$

Questions:

(a) Write the equation for the reaction.



Mole ratio = 1 mol of acid: n mol of base

(b) Calculate:

(i) Molarity of sodium hydroxide.

Molarity of BA1 = ?, Concentration = 8 g/l, RFM = 40

$$\begin{aligned}\text{Molarity} &= \frac{\text{Concentration in g/l}}{\text{RFM}} \\ &= \frac{8}{40} \\ &= \mathbf{0.2\text{ M}}\end{aligned}$$

(ii) The number of moles of **BA1** that reacted with **BA2**.

$$\begin{aligned}\text{Volume of BA1 (acid)} &= \text{average volume} \\ &= \mathbf{20.00\text{ cm}^3}\end{aligned}$$

Molarity of BA1 = 0.2M, Number of moles = ?

0.2 moles of acid are contained in 1000 cm³ of solution.

The number of moles of BA1 contained 20.00 cm³ of solution

$$\begin{aligned}&= \frac{20 \times 0.2}{1000} \\ &= \mathbf{0.004}\end{aligned}$$

(iii) The number of moles of **BA2** that reacted with **BA1**.

Volume of BA2 = 20.00 cm³. Molarity of BA1 = 0.1M, Number of moles = ?

0.1 moles of acid are contained in 1000 cm³ of solution.

The number of moles of BA1 contained 20.00 cm³ of solution

$$\begin{aligned}&= \frac{20 \times 0.1}{1000} \\ &= \mathbf{0.002}\end{aligned}$$

(c) Determine the basicity of the acid.

From the reaction equation, 1 mole of BA2 requires n moles of BA1.

i.e 1 mole of BA2 : n moles of BA1.

0.002 moles : 0.004 moles

$$0.002 n = 0.004$$

$$n = \frac{0.004}{0.002}$$

$$\therefore n = 2$$

Therefore, the basicity of the acid = 2

SELF-CHECK 1.1

1. (2005 Q.26)

The concentration, in grams per litre, of a 0.05 M sodium carbonate solution is [N=23; O=16; C=12]

A. 0.05×83

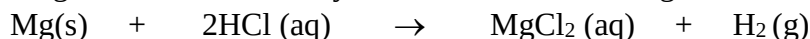
B. 0.05×106

C. $\frac{106}{0.05}$

D. $\frac{83}{0.05}$

2. (2004 Q.6)

Magnesium reacts with hydrochloric acid according to the following equation:



The volume of hydrogen formed at s.t.p. when 2.32 g of magnesium reacts completely with dilute hydrochloric acid is

[Molar gas volume at s.t.p. is 22.4 dm^3 , $\text{mg} = 24$]

A. $\frac{22.4 \times 2.32}{24} \text{ dm}^3$

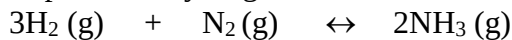
B. $\frac{22.4 \times 24}{2.32} \text{ dm}^3$

C. $\frac{2.32 \times 24}{22.4} \text{ dm}^3$

D. $2.32 \times 24 \times 22.4 \text{ dm}^3$

3. (2004 Q.10)

Under a certain temperature and pressure, hydrogen reacted with nitrogen according to the equation below:



The volume of nitrogen required to react with 150 cm^3 of hydrogen under the same temperature and pressure is

A. 15.0 cm^3

B. 50.0 cm^3

C. 300.0 cm^3

D. 450.0 cm^3

4. (2004 Q.29)

What mass in grams, of sodium carbonate -10- water ($\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$) is contained in 50cm^3 of 0.1 m solution? [H = 1, C = 12, O = 16, Na = 23]

A. $\frac{106 \times 0.1 \times 100}{50}$

B. $\frac{106 \times 0.1 \times 50}{1000}$

C. $\frac{286 \times 0.1 \times 1000}{50}$

D. $\frac{286 \times 0.1 \times 50}{1000}$

5. (2003 Q.9)

Hydrochloric acid reacts with sodium hydroxide according to the equation: HCl (aq)



25.0 cm^3 of 0.1M hydrochloric acid reacted completely with 20 cm^3 of sodium hydroxide. What is the molarity of sodium hydroxide?

A. $\frac{25 \times 0.1}{20}$

B. $\frac{20 \times 0.1}{25}$

C. $20 \times 0.1 \times 25$

D. $\frac{25 \times 0.1}{20}$

6. (2003 Q.11)

100 cm^3 of nitrogen were reacted with 300 cm^3 of hydrogen at s.t.p. What was the volume of ammonia produced?

A. 100 cm^3

B. 200 cm^3

C. 300 cm^3

D. 400 cm^3

7. (2003 Q.18)

The volume of carbon dioxide evolved when 6.0 g of carbon are burnt completely in air at s.t.p. is [C = 12, Molar volume = 22400cm^3 at s.t.p.]

A. $\frac{12 \times 22400}{6}\text{ cm}^3$.

B. $\frac{12 \times 6}{22400}\text{ cm}^3$.

C. $\frac{6 \times 22400}{12}\text{ cm}^3$.

D. $6 \times 12 \times 22400\text{ cm}^3$.

8. (2003 Q.21)

20.0 cm^3 of 0.1 M sodium hydroxide solution reacted with 0.1M of solution Y. The volume of solution Y that reacted completely with the alkali is [Mole ratio NaOH: Y = 2:1]

A. 40 cm^3 .

B. 30 cm^3 .

C. 20 cm^3 .

D. 10 cm^3 .

9. (2003 Q.26)

20 cm^3 of an acid HX neutralized 25 cm^3 of 0.05 M sodium carbonate solution. The molarity of the acid is

A. $\frac{25 \times 0.05}{20}$

B. $\frac{2 \times 25 \times 0.05}{20}$

C. $\frac{2 \times 20 \times 0.05}{25}$

D. $\frac{188}{1.88 \times 80}$

10. (2002 Q.19)

Calculate the relative molecular mass of gas **P** if 8.4 dm³ of the gas has a mass of 0.93 g. (1 mole of gas occupies 22.4 dm³ at s.t.p.)

A. $\frac{0.93 \times 22.4}{8.4}$

B. $\frac{22.4 \times 8.4}{8.4}$

C. $\frac{0.93 \times 8.4}{22.4}$

D. $\frac{0.93}{22.4 \times 8.4}$

11. (2002 Q.25)

The mass of silver nitrate, AgNO₃ in 0.2 M solution of the salt is (Ag = 108, O = 16, N = 14)

A. 17.0

B. 34.0

C. 85.0

D. 170.0

12. (2002 Q.35)

The mass in g of OH⁻ ions in 0.25M NaOH solution is (H = 1, O = 16)

A. $\frac{0.25}{17}$

B. 17 x 0.25

C. $\frac{0.25 \times 4}{17}$

D. $\frac{17 \times 0.25}{4}$

13. (2002 Q.36)

The minimum volume of 1M HCl (aq) required to produce 0.25 g of hydrogen with excess magnesium is

A. 25 cm³.

B. 100 cm³.

C. 250 cm³.

D. 1000 cm³.

14. (2001 Q.26)

25 cm³ of a 0.25M acid required 25 cm³ of 0.5M sodium hydroxide solution for neutralisation. The basicity of the acid is

A. 1

B. 2

C. 3

D. 4

15. (2000 Q.19)

What is the mass of sulphuric acid (Mr = 98) in 5 cm³ of a 0.2 M solution of the acid?

A. $\frac{98 \times 5}{0.2 \times 1000}$

B. $\frac{98 \times 0.2 \times 5}{1000}$

C. $\frac{98 \times 0.2}{5 \times 1000}$

D. $\frac{98 \times 5 \times 1000}{0.2}$

16. (2000 Q.21)

25 cm³ of M sodium carbonate required 22.70 cm³ of hydrochloric acid for complete neutralisation. The molarity of the acid is given by

A. $\frac{0.00125 \times 000}{2 \times 22.7}$

B. $\frac{0.00125 \times 1000}{2 \times 25}$

C. $\frac{0.00125 \times 2}{22.7}$

D. $\frac{0.00125 \times 2 \times 1000}{25}$

17. (1999 Q.13)

Calcium carbonate reacts with hydrochloric acid according to the following equation:



The mass of CaCO_3 that will react completely with 50 c.c of 2M hydrochloric acid is ($\text{CaCO}_3 = 100$)

- | | |
|---|---|
| A. $\frac{2 \times 50 \times 100}{2 \times 1000}$ | B. $\frac{2 \times 1000}{50 \times 2 \times 100}$ |
| C. $\frac{50 \times 100}{200 \times 1000}$ | D. $\frac{100 \times 50 \times 2}{1000}$ |

18. (1999 Q.30)

What is the molarity of sodium hydroxide solution if 30 cm^3 of 0.2 M hydrochloric acid just neutralises 20 cm^3 of the alkali?

- | | |
|-------------------------------|-------------------------------|
| A. $\frac{20}{0.2 \times 30}$ | B. $\frac{20 \times 0.2}{30}$ |
| C. $\frac{30}{0.2 \times 20}$ | D. $\frac{30 \times 0.2}{20}$ |

19. (1998 Q.4)

20 cm^3 of 0.2 M HCl reacts with 25 cm^3 of sodium hydroxide solution. The molarity of the hydroxide is

- | | |
|-------------------------------|-------------------------------|
| A. $\frac{25 \times 0.2}{20}$ | B. $\frac{20 \times 0.2}{25}$ |
| C. $\frac{25}{20 \times 0.2}$ | D. $\frac{20}{25 \times 0.2}$ |

20. (1997 Q.10)

0.02 moles of calcium chloride (CaCl_2) is dissolved to make 200 cm^3 of solution. What is the concentration of chloride ions in moles per litre, in this solution?

- | | |
|------------|-----------|
| A. 0.05 M. | B. 0.1 M. |
| C. 0.2 M. | D. 0.3 M. |

SELF-CHECK 1.2

1. A gaseous hydrocarbon X contains 20% hydrogen by mass. 7.5 g of X occupies 5.6 dm^3 at stp.

- | | | |
|----------------|-------|------------------------------|
| (a) Calculate: | (i) | The empirical formula of X. |
| | (ii) | The molar mass of X. |
| | (iii) | The molecular formula of X. |
| (b) Write: | (i) | The name of X. |
| | (ii) | The structural formula of X. |

(C = 12, H = 1)

CHAPTER TWO

EFFECT OF ELECTRICITY ON SUBSTANCES (ELECTROLYSIS)

Electrolysis is divided into two parts, namely:

- Part I. Qualitative Electrolysis
- Part II. Quantitative Electrolysis

QUALITATIVE ELECTROLYSIS

Experiment:

To investigate the conduction of Electricity by a variety of substances

Apparatus:

- (a) (i) Solids:
Small pieces of: -
- Metals e.g. Magnesium, Copper, Lead, etc.
 - Non-metallic substances: e.g. Rubber, Plastic, etc.
 - Compounds e.g. salts, sugar etc.
- (ii) Liquids:- Water, alcohol (ethanol), urea, urine
- (iii) Aqueous solutions of:
- *Salts*: E.g. Sodium chloride, Copper (II) sulphate, Silver nitrate etc.
 - *Strong acids (Mineral Acids)*.
E.g. Hydrochloric acid, Sulphuric acid,
Nitric acid and Phosphoric acid.
 - *Weak acids*: E.g. Ethanoic acid, Oxalic acid etc.
 - *Strong alkalis*: E.g. Sodium hydroxide, Potassium hydroxide, Aluminium hydroxide etc.
 - *Weak alkalis*: E.g. Ammonia Solution.
- (b) A beaker, 2 graphite rods, 2 Crocodile clips, 6V D.C. supply, 1 6V lamp bulb and 1 switch.

Procedure:

- Arrange the apparatus as shown in the diagram below, in which a 6V D.C. supply is connected in series with a 6V lamp to a pair of crocodile clips.

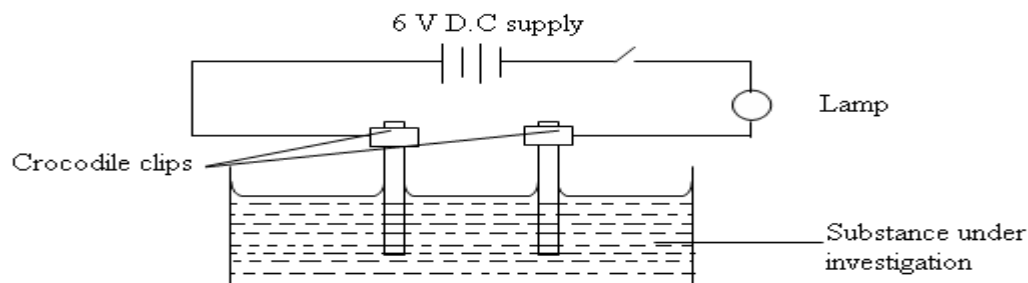


Fig. 1

- Test the conductivity of the substances in the list above in turn and in each case, observe the bulb.

Notes:

- If :
- (i) the bulb lights or glows, it indicates that the circuit is complete, which means that the substance under investigation is a good conductor of electricity.
 - (ii) the bulb does not light or glow, it indicates that the circuit is incomplete, which means that the substance under investigation is a bad conductor of electricity.

NB: *For the solid substances simply attach the crocodile clips to the ends of the spacemen while for the liquids/solutions attach the crocodile clips to the graphite rods and then insert them into the spacemen under investigation.*

Observations:

- (i) The bulb glows, brightly, for some solids especially all metals and some aqueous solutions and dimly, for some solutions.
- (ii) While for some solids e.g. rubber and some liquids the bulb did not glow at all.

Results:

The results showed that:

- (i) All metals are good conductors of electricity.
- (ii) Ionic compounds in a solid state e.g. salts do not conduct electricity.
- (iii) Ionic compounds in solution or molten state are good conductors of electricity.
- (iv) Covalent compounds are bad conductors of electricity.
- (v) Weak acids and alkaline solutions and water conduct electricity weakly.

Terms Used:

1. Conductors and Insulators(Non-conductors)

(a) Conductors:

A conductor is a substance, which allows electric current to pass through it.

E.g. - All metals and carbon in form of graphite.

(b) Insulators (Non-conductors):

These are substances, which do not allow electricity to pass through. Examples of insulators include: rubber, plastic, wood etc.

2. Electrolytes and Non-electrolytes:

(a) Electrolytes:

An electrolyte is a compound which, when in solution or molten, conducts an electric current and is decomposed by it.

Examples of electrolytes include:

- Molten ionic compounds.
- Aqueous solutions of ionic compounds (salts, acids and alkalis).

NB:

- ❖ All electrolytes are ionic compounds. I.e. are composed of ions.
- ❖ The above compounds only conduct electricity in molten (fused) and aqueous state. They **DO NOT** conduct electricity in solid state.

Reason:

- *In solid state, the ions are rigidly held in regular positions by strong forces as such are unable to move freely.*
- *In molten and aqueous state, the ions move freely hence conduct electricity.*
- ❖ The passage of electricity through the electrolyte is accompanied by a chemical decomposition.
- ❖ The decomposition of an electrolyte, when electric current passes through it is called *electrolysis*.

(a) Non-electrolyte:

A non-electrolyte is a compound which, when in aqueous or molten does not conduct electricity.

Examples include: - all covalent compounds. E.g. alcohol, sugar etc.

Differences between Electrolyte and Non-electrolyte

Electrolyte	Non-electrolyte
- Conducts electricity in molten or solution form.	- Do not conduct electricity in molten or solution form.
- Consists of free moving oppositely charged ions.	- Consists of covalent molecules and therefore no free moving ions.

(c) Types of Electrolytes:

There are two types of electrolytes, namely:-

- Strong Electrolytes and
- Weak Electrolytes.

(i) Strong Electrolyte:

A strong electrolyte is a compound, which ionizes or dissociates, completely in dilute solution or in the molten state into ions.

Examples of strong electrolytes are shown in the table below.

Compound	Name of Compound	State	Dissociation equation
Salts	Lead bromide	Molten	$\text{PbBr}_2(\text{s}) \rightarrow \text{Pb}^{2+}(\text{l}) + 2\text{Br}^-(\text{l})$
	Sodium Chloride	Aqueous	$\text{NaCl}(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
	Copper (II) Sulphate	Aqueous	$\text{CuSO}_4(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$
Acids	Hydrochloric acid	Aqueous	$\text{HCl}(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
	Sulphuric acid		$\text{H}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$
	Phosphoric acid		$\text{H}_3\text{PO}_4(\text{aq}) \rightarrow 3\text{H}^+(\text{aq}) + \text{PO}_4^{3-}(\text{aq})$
Alkali	Sodium Hydroxide	Aqueous	$\text{NaOH}(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$

(ii) Weak Electrolyte:

A weak electrolyte is a compound, which ionizes or dissociates, slightly in dilute solution or in the molten state.

Examples of weak electrolytes are shown in the table below.

Compound	Name of Compound	State	Dissociation equation
Alkali	Ammonium Hydroxide	Aqueous	$\text{NH}_4\text{OH}(\text{aq}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$
Acid	Ethanoic acid	Aqueous	$\text{CH}_3\text{COOH}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$
Water	Hydrogen Oxide	Liquid	$\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$

(d) Difference between Strong Electrolyte and Weak Electrolyte

Strong Electrolyte	Weak Electrolyte
- Ionizes completely in aqueous solutions	- Ionizes slightly in aqueous solutions
- Contains ions and no molecules	- Contains ions and molecules
- Produces a higher current thus causing a bulb to glow brightly.	- Produces small current thus causing a bulb to glow dimly.
- Causes large deflection in an ammeter	- Causes small deflection in an ammeter

3. Electrolysis:

Electrolysis is the decomposition of a substance (an electrolyte) by passing an electric current through it.

NB: Electrolysis is accompanied by chemical changes, which occur at the electrodes.

4. (a) Electrodes:

Electrodes are the two poles (conductors) through which electric current enters and leaves the electrolyte.

NB: (i) The electrodes are connected to the terminal of the D.C supply.
(ii) Before the switch is closed, the electrodes have no charges.
(iii) They acquire charges, on the closer of the switch, hence giving the types of electrodes.

(b) Types of electrodes:

There are two types of electrodes, namely:-

- anode and
- cathode.

(i) Anode:

Anode is the positively charged electrode through which electric current enters the electrolyte.

- It is electron deficient.
- It is connected to the positive terminal of D.C supply.
- It is usually tall in size, in the circuit symbol.

(ii) Cathode:

Cathode is the negatively charged electrode at which the electric current leaves the electrolyte.

- It is reach with electrons.
- It is connected to the negative terminal D.C supply
- It is usually short in size, in the circuit symbol.

The Ionic Theory (Put forward by Arrhenius about 1880)

States that:

An electrolyte consists of free positively and negatively charged particles called ions, which are responsible for the transmission of electricity through the electrolytes.

(a) Types of ions:

There are two types of ions, namely:-

- **Cations** - positively charged ions and
- **Anions** - negatively charged ions.

Examples of ions are shown in the table below.

Cations	Anions
(i) Metallic ions: K^+ , Na^+ , Ca^{2+} , Mg^{2+} , Al^{3+} , Zn^{2+} Fe^{2+} , Fe^{3+} , Pb^{2+} , (H^+) , Cu^{2+} , Ag^+	Non-metallic ions: Cl^- , Br^- , I^-
(ii) Radicals: NH_4^+	Radicals: SO_4^{2-} , NO_3^- , CO_3^{2-} , OH^-

(b) Experiment:

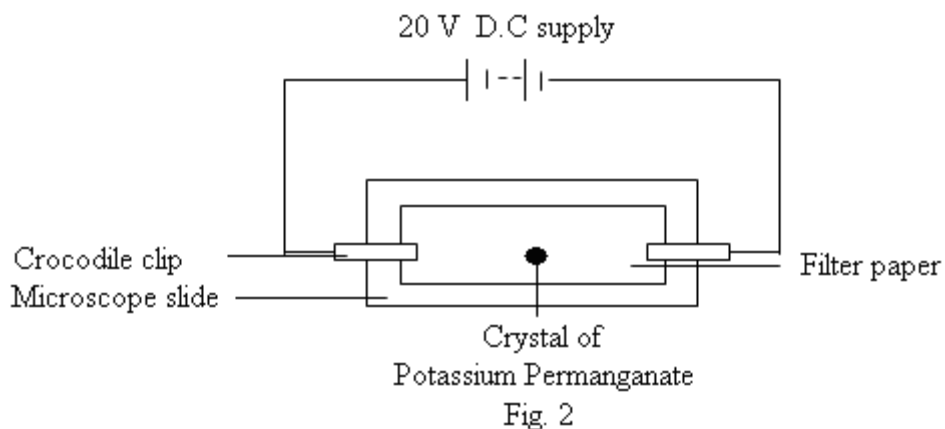
To investigate the movement of ions during electrolysis

(i) Apparatus:

- 20 V D.C supply,
- a microscope slide,
- 2 crocodile clips.
- A strip of filter paper,
- 4 pieces of connecting wire and
- a crystal of coloured salt e.g. Potassium permanganate or Copper (II) sulphate.

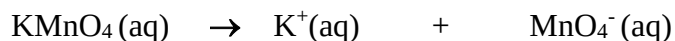
(ii) Procedure:

- Cut a piece of filter paper approximately the size of microscope slide.
- Attach the piece of the filter paper on top of the microscope slide using crocodile clips.
- Moisten the filter paper with water.
- Place a small crystal of Potassium permanganate mid-way between the points of attachment of the clips.
- Connect the slide and its contents in a circuit with a 20V D.C. supply as shown in figure 2 and observe the paper over a period of 10 - 15 minutes.



(iii) Observation:

- A purple colour is seen moving towards the anode.
- Since in aqueous state the Potassium permanganate ionizes according to the following equation:



The permanganate ions are attracted by the anode and therefore move towards the anode. This purple colour of permanganate is due to the MnO_4^- ions.

(c) **Explanation of Electrolysis using the ionic theory:**

Consider the electrolytic cell diagram in fig.3 below.

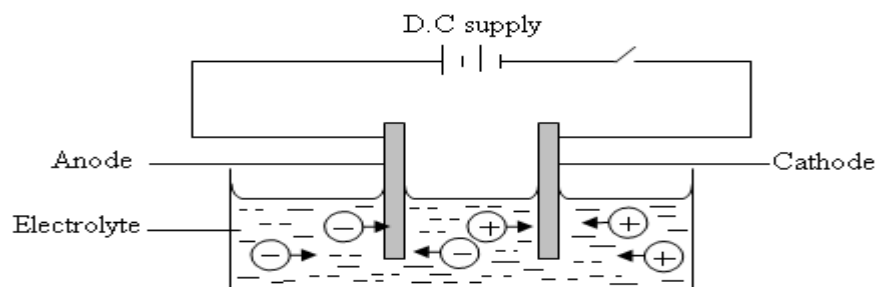


Fig. 3

- When the circuit is complete, the process of electrolysis begins.
- During the process, the law of electricity, which states that: **like charges repel and unlike charges attract** is obeyed and the ions migrate to their respective electrodes.
- That is the attraction between the electrodes and the ions begins.
- The cathode attracts cations (negatively charged ions) while the anode attracts anions (positively charged ions).
- Thus resulting to ion migration to the respective electrode.

Memory Aid:

You can remember which ion goes to which electrode by the statement:
"**Cat at Cat**" and "**An at An**"

(d) **The process of discharge:**

When ions arrive at their respective electrodes, they become discharged i.e. they lose their charges and form products.

The anions lose electrons to the anode and the cations gain electrons from the cathode.

(i) **Discharge at the Cathode:**

When cations arrive at the cathode, which is electron rich, they 'strike' the cathode and pick electrons from the cathode equal to their charges.

The electrons neutralize the positive charge on the ion and the ion turns back to atom(s).

Metal cations become metal atoms and are deposited on the cathode. For hydrogen, two atoms combine to form a molecule of hydrogen gas leading to the formation of colourless bubbles on the cathode.

(ii) Discharge at the Anode:

When anions arrive at the anode, which is electron deficient, they lose their excess electron(s) to the anode equal to their valences. Thus they become discharged and revert to atom(s).

Insoluble gaseous atoms form molecules and are evolved and bubbles of colourless gas are seen on the lower part of the anode.

(e) Selective/Preferential Discharge of Ions at the Electrodes

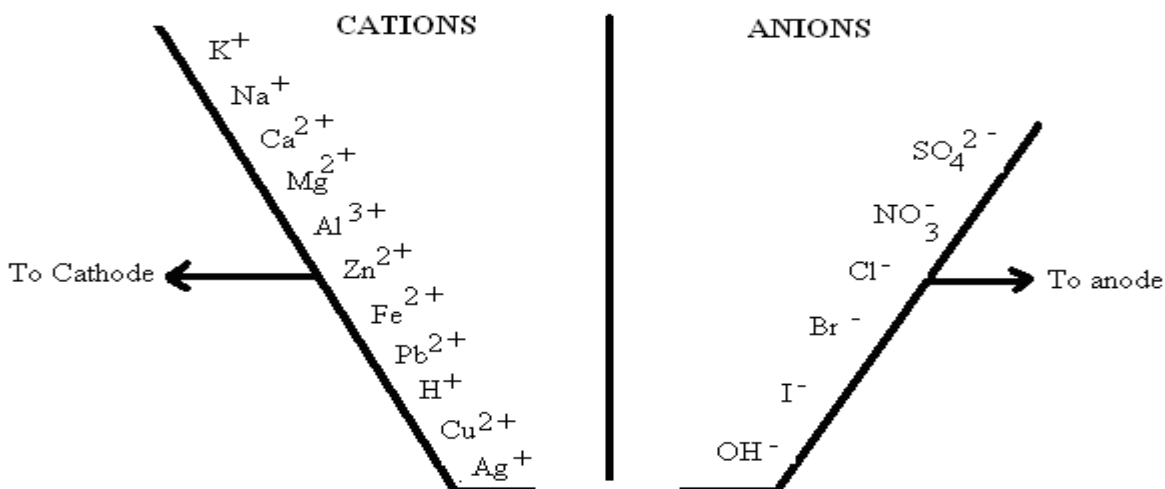
When two or more ions of the same charge arrive at an electrode, only one is selected preferentially discharged depending on the factors of preferential discharge.

The Factors which Determine Preferential Discharge of Ions at the Electrodes

1. Position of the ion in the Electro-Chemical Series:

When two or more ions of a similar charge are present in a dilute electrolyte, the one lower in the series is discharged in preference to that above it.

The electrochemical series



Highly electropositive and highly electronegative ions are not easily discharged unless there are no other less electropositive cations or less electronegative anions.

2. Concentration of the ion:

When ions of similar charge are present, the one in greater concentration is discharged preferentially even if it is higher in the E.C.S.

3. Nature of the electrode.

Different electrodes for a given electrolyte may cause the formation of different products despite the above factors.

For examples:

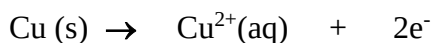
(i) *Electrolysis of Sodium Chloride.*

When Sodium chloride solution is electrolyzed:

- ❖ *using platinum cathode*, H^+ ions are discharged instead of Na^+ ions and
- ❖ when mercury is used as a cathode, Na^+ ions are discharged. The product formed is called *Sodium amalgam*.

(ii) *Electrolysis of Copper (II) Sulphate solution*

When Copper (II) Sulphate solution is electrolyzed using copper anode, neither OH^- nor SO_4^{2-} ions are discharged, but rather electrode ionization occurs according to the equation:



The copper ions formed are released into the electrolyte and the electrons are immediately passed to the cathode.

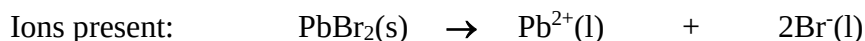
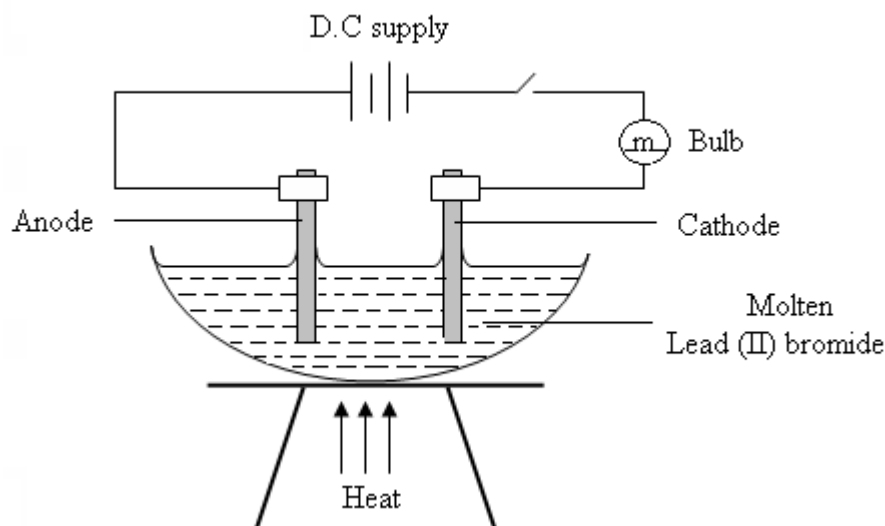
Note that:

- ❖ *A passive/un-reactive electrode does not affect the process of electrolysis.*
- ❖ *For this reason platinum or carbon (in form of graphite) which are passive are commonly used as electrodes.*
- ❖ *However, Platinum (Pt) is not used where chlorine is produced. This is because chlorine attacks platinum.*

EXAMPLES OF ELECTROLYSIS

1. Electrolysis of Salts:

(a) *Electrolysis of Molten Lead bromide using graphite electrodes.*

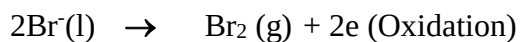
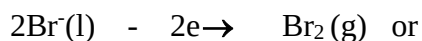


The ions formed migrate to the electrode of opposite charge, i.e. Pb^{2+} ions move to the cathode (negative electrode) and Br^{-} ions move to anode (positive electrode). On their arrival at the respective electrode, they become discharged.

Equations for the reactions at the electrodes

(i) At the anode:

The negatively charged bromide ions (Br^{-}) become discharged by loss of electrons to the anode according to the equation:

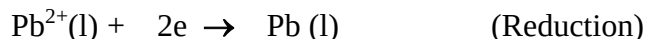


Observations:

- Bubbles of reddish brown gas with choking and irritating smell are formed.
- The gas is slightly soluble in water, forming a reddish brown solution.
- The bulb glows brightly.

(ii) At the cathode:

The positively charged lead ions (Pb^{2+}) become discharged by gain of electrons from the cathode one lead ion gains two electrons to become lead atom according to the equation:

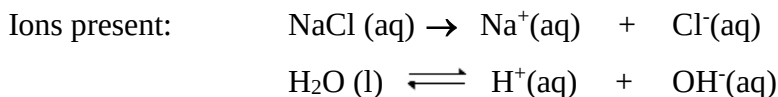


Observations:

A grey solid formed melts and sinks to the bottom of the crucible.

NB: *The electrolysis of molten Lead (II) bromide is comparatively easy to understand because only one type of cation and one type of anion are present in the electrolyte. In aqueous solutions, there are two or more cations and anions except for dilute sulphuric acid where there is only one cation (H^+ ion).*

(b) Electrolysis of Sodium chloride solution using graphite or platinum cathode and graphite anode (to resist attack by chlorine gas).



The ions formed migrate to the electrode of opposite charge, i.e. Na^+ and H^+ ions move to the cathode (negative electrode) and Cl^- and OH^- ions move to anode (positive electrode) and then a preferential discharge occurs.

Equations for the reactions at the electrodes:

(i) At anode:

The negatively charged Cl^- ions are preferentially discharged by loss of electrons to the anode according to the equation:

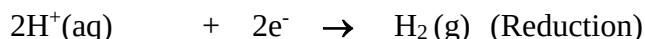


Observations:

- Bubbles of greenish-yellow gas with a choking, unpleasant and irritating smell are seen.
- The bulb glows brightly.

(ii) At the cathode:

The positively charged H^+ ions are preferentially discharged by gain of electrons from the cathode according to the equation:



Observations:

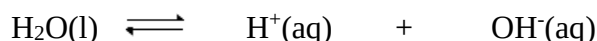
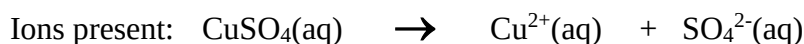
- Bubbles of colourless gas that burns with a pop sound are formed.

NB:

- ❖ Discharge of H^+ ions disturbs the ionic equilibrium of water and causes more water molecule to ionize to restore it.
- ❖ Excess OH^- ions so produced, with the presence of Na^+ ions in the solution, is equivalent to presence of sodium hydroxide.
- ❖ The ratio of the products indicates equal volumes. However, in practice the volume of chlorine produced is less than the expected because chlorine formed reacts with OH^- ions according to the equation:



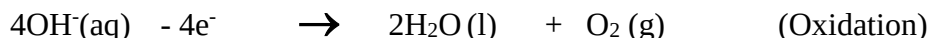
- ❖ At some concentration, both chlorine and oxygen are produced at the anode.

(d) Electrolysis of Copper (II) sulphate solution using graphite electrode or platinum anode

The ions formed migrate to the electrode of opposite charge, i.e. Cu^{2+} and H^+ ions move to the cathode (negative electrode) and SO_4^{2-} and OH^- ions move to anode (positive electrode) and then a preferential discharge occurs.

Equations for the reactions at the electrodes**(i) At the anode:**

The negatively charged OH^- ions are preferentially discharged by loss of electrons to the anode according to the equation:

**Observations:**

- Bubbles of colourless gas that rekindles a glowing splint are formed.
- The bulb glows brightly.

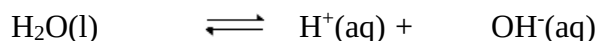
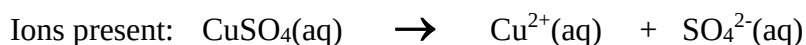
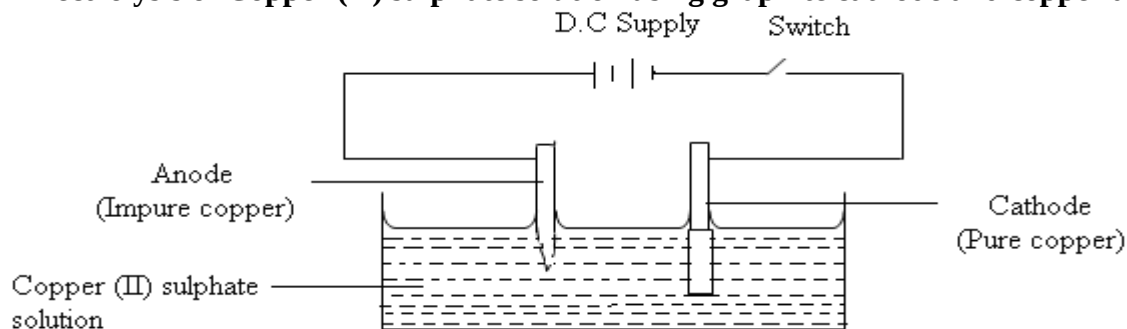
(ii) At the cathode:

The positively charged Cu^{2+} ions are preferentially discharged by gain of electrons from the cathode according to the equation:

**Observations:**

- A brown solid formed at the lower part of cathode.
 - The blue colour of the electrolyte fades. This is because the copper (II) ions which are responsible for the blue colour are being converted to copper.
- If the electrolysis is left for a long period of time, the solution becomes colourless.

(e) Electrolysis of Copper (II) sulphate solution using graphite cathode and copper anode

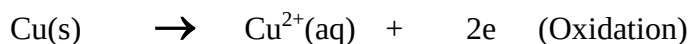


The ions formed migrate or move towards the electrode of opposite charge, i.e. Cu^{2+} and H^+ ions move to the cathode (negative electrode) and SO_4^{2-} and OH^- ions move to anode (positive electrode) and then a preferential discharge occurs.

Equations for the reactions at the electrodes

(i) At the anode:

Although both OH^- and SO_4^{2-} ions arrive at the anode, none of them is discharged. Instead, the copper anode dissolves into the solution according to the equation:

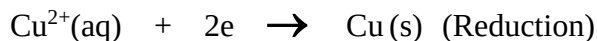


Observations:

- The copper anode dissolves into the solution thus reducing in size.

(ii) At the cathode:

The positively charged Cu^{2+} ions are preferentially discharged by gain of electrons from the cathode according to the equation:

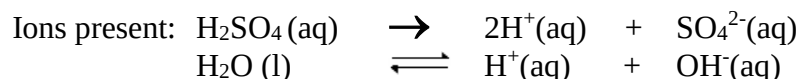


Observations:

- A brown solid formed at the lower part of cathode.
- The blue colour of the electrolyte remains.
- This is because the copper ions discharged are being replaced by the ones formed from the anode. That is copper ions are being effectively transferred from the anode to the cathode. Hence, the concentration of the copper (II) ions remains constant.

2. Electrolysis of Acids

(a) Electrolysis of Sulphuric acid using platinum electrodes.

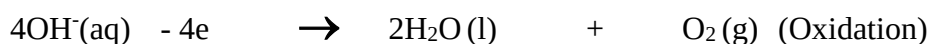


The ions formed migrate to the electrode of opposite charge, i.e. H^+ ions the only cations move to the cathode (negative electrode) and SO_4^{2-} and OH^- ions move to anode (positive electrode) and then a preferential discharge occurs.

Equations for the reactions at the electrodes

(i) At the anode:

The negatively charged OH^- ions are preferentially discharged by loss of electrons to the anode according to the equation:



Observations:

– Bubbles of colourless gas that relights a glowing splint are formed.

(ii) At the cathode:

The only positively charged H^+ ions are discharged by gain of two electrons from the cathode according to the equation:



Observations:

– Bubbles of colourless gas that burns with a pop sound are formed.

NB: *Total acidity remains constant.*

Explanation:

The discharge of H^+ ions at the cathode is equivalent to decrease of the concentration of the acid. But the discharge of OH^- ions disturbs the ionic equilibrium of water causing more water molecules to ionize to restore it according to the equation:



Excess H^+ ions so produced plus the SO_4^{2-} ions is equivalent to increased concentration of sulphuric acid. Therefore, at the cathode the acidity decreases while at the anode it increases, hence the total acidity remains constant.

The following points about electrolysis of solutions should be remembered:

- Metals, if produced, are discharged at the cathode.
- Hydrogen is produced at the cathode only.
- Non-metals, apart from hydrogen are produced at the anode.
- Reactive metals are not formed at the cathode during electrolysis of aqueous solutions. An exception is during electrolysis of sodium chloride using a mercury cathode.
- The products produced depend upon the concentration of the electrolyte in the solution. For example, electrolysis of concentrated sodium chloride produces chlorine at the anode but electrolysis of dilute sodium chloride can produce oxygen at the anode.

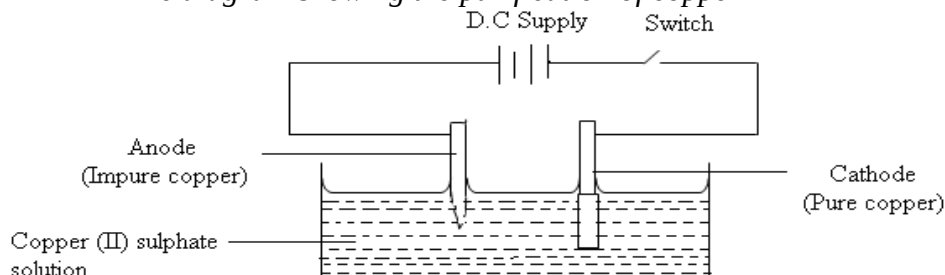
APPLICATIONS OF ELECTROLYSIS

1. **Purification of Copper.**

In the purification process, the anode is impure copper rod or plate and the cathode is pure copper plate or copper rod.

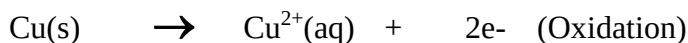
The electrolyte used must contain copper ions e.g. copper (II) sulphate solution.

The diagram showing the purification of copper



Explanation

During the electrolysis, copper dissolves from the anode, according to the equation.



The positively charged Cu^{2+} ions so formed move to the cathode and are preferentially discharged by gain of two electrons from the cathode according to the equation:



NB: *At the bottom of the cell, “anode slime” collects. The slime is rich in precious metals, e.g. Silver, Gold, and is formed from the impurities in the original impure copper.*

2. **Electroplating**

Electroplating is the coating of the surfaces of objects/metal surface with a thin layer of another metal.

It is done to:

- (i) Decorate ornaments and other goods.*
- (ii) Prevent corrosion.*

The process is the same as for the purification of copper.

Points to note:

- (i) The metal/object to be plated or coated is made the cathode and its surface must be clean and grease-free.*
- (ii) The metal to be used for coating is made the anode.*
- (iii) The electrolyte used should contain the cations of the metal used for the coating.*

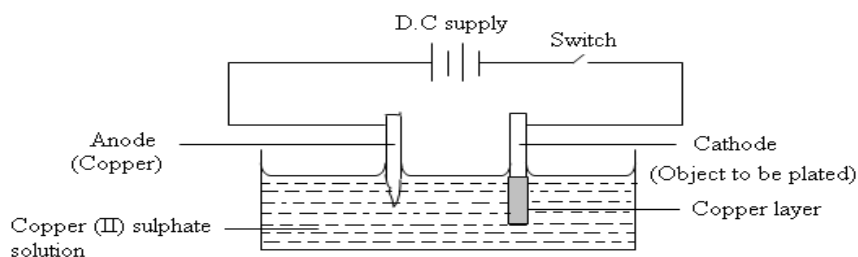
Observations

- The cathode surface is coated with a metal layer.*
- The anode dissolves and goes into solution as ions.*

Example UNEB 1992 Q.3 (b)

Draw a labeled diagram of an apparatus that can be used in the laboratory to copper plate an object.

Solution



3. Extraction of reactive metals

E.g. Potassium, Sodium, Calcium, Magnesium and Aluminium.

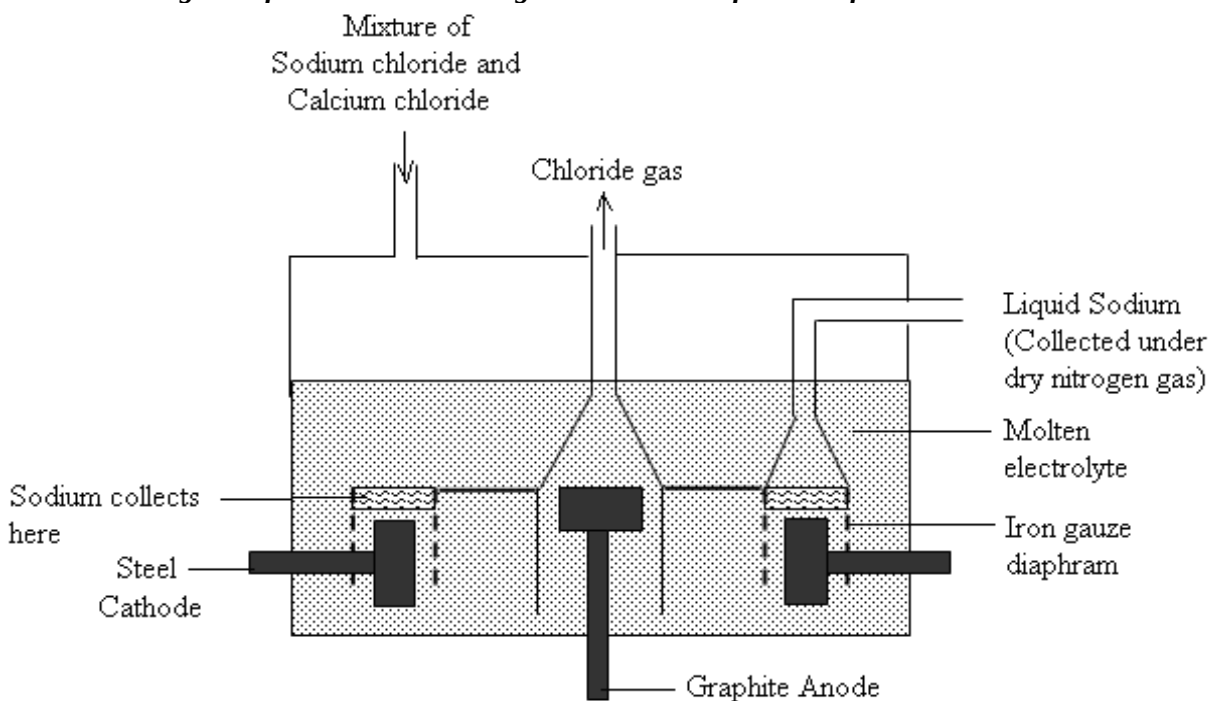
Extraction of reactive metals e.g Sodium from Sodium Chloride

Sodium is extracted by the electrolysis of molten sodium chloride in the Downs Cell.

The Downs Cell consists of a cylindrical steel cathode and graphite anode.

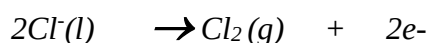
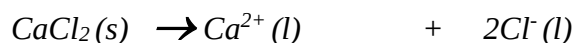
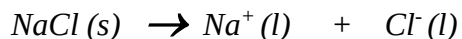
Sodium chloride is heated and melts at 801 °C. Calcium chloride is added as an impurity to lower the melting point to about 600 °C.

The Diagram of Downs Cell showing the Extraction of Sodium from Molten Sodium chloride



Explanation:

Ions present:



The products Sodium and Chlorine produced are kept apart to prevent them from reacting and reforming sodium chloride.

The liquid Sodium is collected under dry nitrogen gas. This prevents the metal from reacting with the atmosphere.

4. **Manufacture of Sodium Hydroxide.**

Sodium Hydroxide is produced industrially by the electrolysis of sodium chloride solution (brine) using either the:

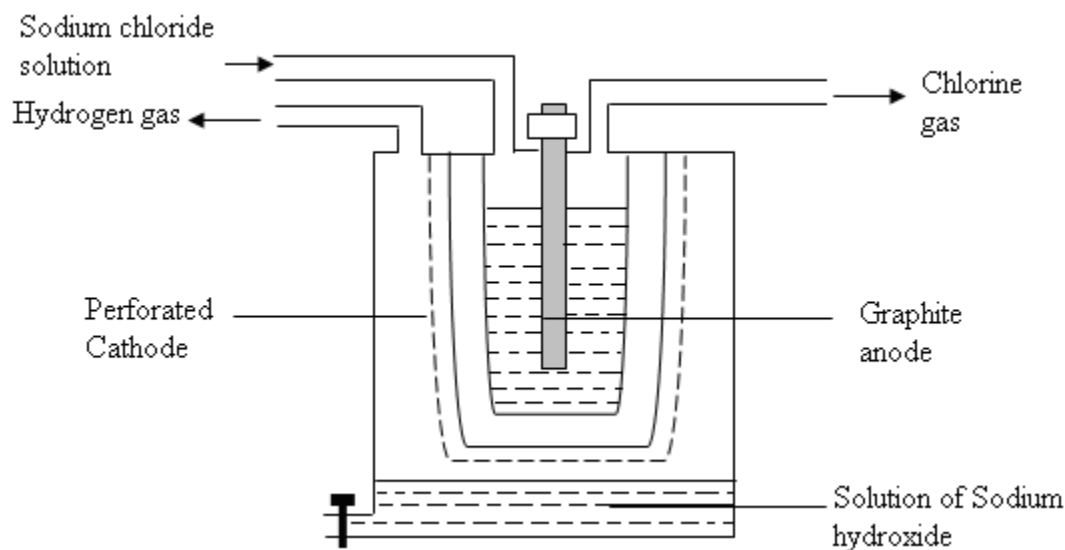
- Diaphragm cell or
- Kellner Solvay Cell.

In both cases the products are sodium hydroxide, hydrogen gas and chlorine gas.

In the diaphragm cell the electrolysis of nearly saturated solution of sodium chloride is carried out using a graphite anode and a steel cathode.

The cell is made of an asbestos substitute. It allows the solution to pass through but prevents the products chlorine and sodium hydroxide solution from coming into contact.

The Diagram of a Diaphragm Cell showing the manufacture of Sodium hydroxide



QUANTITATIVE ELECTROLYSIS

(a) The Laws of Electrolysis

The laws expressing the quantitative results of electrolysis were first stated by Faraday.

The laws assert that:

The amount (expressed in moles) of an element liberated during electrolysis depends upon:

- (i) The magnitude of the steady current passed
- (ii) The time taken for the current to pass through the electrolyte.
- (iii) The charge on the ion of the element.

(b) The Quantity of Electricity

The quantity of electricity, Q , (measured in units called coulombs) is given by the product of the current (measured in amperes) and the time (measured in seconds).

Thus:
$$\begin{array}{lcl} \text{Quantity of electricity} & = & \text{Current} \times \text{Time} \\ \text{(Coulombs)} & & \text{(Amps)} \times \text{(Seconds)} \end{array}$$

i.e.
$$Q = I t$$

(c) Faraday's Laws of Electrolysis:

Law I

Faraday's first law of electrolysis states that:

The mass of a given element liberated at (or dissolved from) an electrode during electrolysis is proportional to the quantity of electricity consumed or passed through the electrolyte.

Law II

The second law of electrolysis states that:

When the same quantity of electricity is passed through different electrolytes, the relative numbers of moles of the elements deposited are inversely proportional to the charges on the ions of each of the elements respectively.

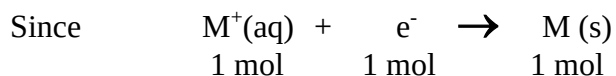
Or

The masses of the different elements liberated by the same quantity of electricity form simple whole number ratios when divided by their relative atomic masses.

Faraday's Constant:

The least quantity of electricity required to liberate one mole of a univalent element is the **Faraday**. Its approximate value is 96500C/mol. (96487C/mol).

Let the equation below represent chemical reaction at the cathode for a univalent metal.



From the equation, it follows that for one mole of uni-positive ion to be discharged, the charge carried by one mole of electrons will be required.

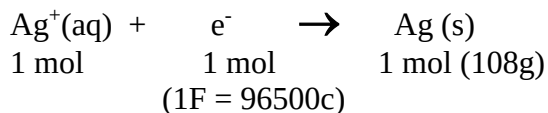
Hence,

One Faraday = One mole of electrons

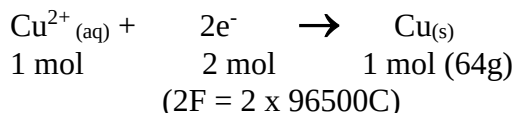
i.e. 1F = 1 mol of e⁻

This information may be used as follows:

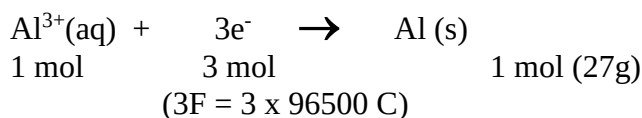
(i) Univalent cation



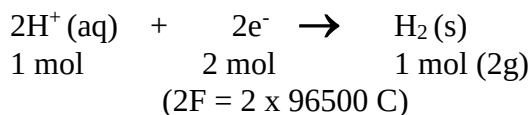
(ii) Divalent cation



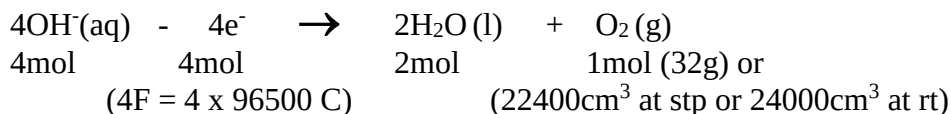
(iii) Trivalent cation



(iv) For a diatomic gas (except oxygen) e.g. Hydrogen or Chlorine.



(v) For oxygen:



From the above, it can be seen that the higher the charge on the ion of an element, the more the quantity of electricity required to liberate one mole of the element.

EXAMPLES

1. A steady current of 2.50 A was passed through Copper (II) sulphate solution for 40 minutes and 50 seconds.

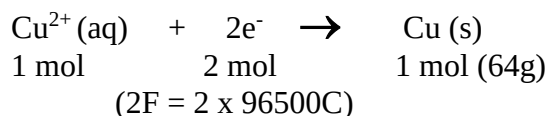
Calculate

- the quantity of electricity passed.
- the mass of copper deposited at the cathode.
- the volume of oxygen gas liberated at the anode in cm^3
($1\text{F} = 96500\text{C}$, $\text{Cu} = 64$, 1 mole of gas at rt = 24l)

Solution: $I = 2.50 \text{ A}$, $t = 40 \times 60 + 50 = 2450 \text{ s}$, $Q = ?$

$$\begin{aligned} \text{(i)} \quad Q &= It \\ &= 2.50 \times 2450 \\ &= \mathbf{6125 \text{ C}} \end{aligned}$$

(ii) Equation at cathode.

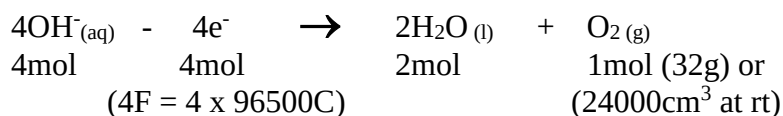


If (2 x 96500) C liberates 64g of copper,

1 coulomb liberates $\frac{64}{2 \times 96500}$ g of copper.

$$\begin{aligned} 6125\text{C liberates } \frac{64}{2 \times 96500} \times 6125 &= 2.031\text{g} \\ &= \mathbf{2.03\text{g}} \end{aligned}$$

(iii) Equation at anode:



If 4 x 96500 C liberates 24000 cm³ of oxygen gas at room temperature,

1 Coulomb of electricity liberates $\frac{24000}{4 \times 96500}$ of oxygen gas.

$$\begin{aligned} 6125\text{C will liberates } \frac{24000}{4 \times 96500} \times 6125 &= 380.975 \\ &= \mathbf{380.98 \text{ cm}^3} \end{aligned}$$

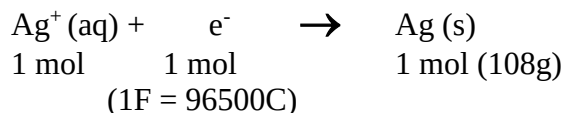
2. A steady current of 5.36 A was passed through an electrolyte of Silver nitrate solution for 30 minutes.

- Calculate
- (i) the quantity of electricity passed.
 - (ii) the mass of silver deposited at the cathode.
 - (iii) the volume of oxygen gas liberated at the anode at stp.
(1F = 96,500C, Ag = 108, 1 mole of gas at stp = 22,400cm³)

Solution: I = 5.36 A, t = 30 x 60 = 1800s, Q = ?

$$\begin{aligned} \text{(i)} \quad Q &= It \\ &= 5.36 \times 1800 \\ &= \mathbf{9648 \text{ C}} \end{aligned}$$

(ii) Equation at cathode.

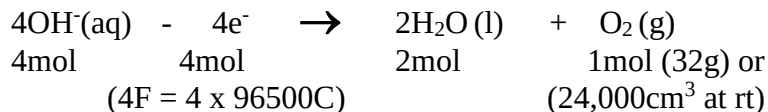


If 96500C liberates 108g of silver,

1 coulomb liberates $\frac{108}{96500}$ g of silver.

9648 C liberates $\frac{108}{96500} \times 9648 = 10.8\text{g}$

(iii) Equation at anode:



If (4×96500) C liberates $22,400 \text{ cm}^3$ of oxygen gas at stp,

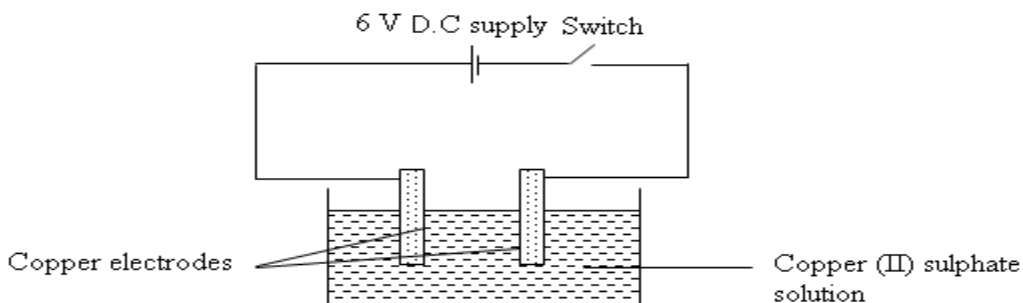
1 coulomb liberates $\frac{22400}{4 \times 96500} \text{ cm}^3$ of oxygen gas.

9648 C liberates $\frac{22400}{4 \times 96500} \times 9648 = 559.88 \text{ cm}^3$

SELF-CHECK 2.1

1. 1987 Q.10

A circuit was connected as shown in the diagram below and a steady current of 0.20 amperes was passed for 20 minutes. ($1F = 96500\text{Cmol}^{-1}$; $\text{Cu} = 64$)



- (a) Write equation for the reaction that took place at the cathode.
- (b) Calculate:
 - (i) The number of coulombs of electricity used.
 - (ii) The number of moles of coulombs of electricity.
 - (iii) The mass of the substance formed at the cathode.

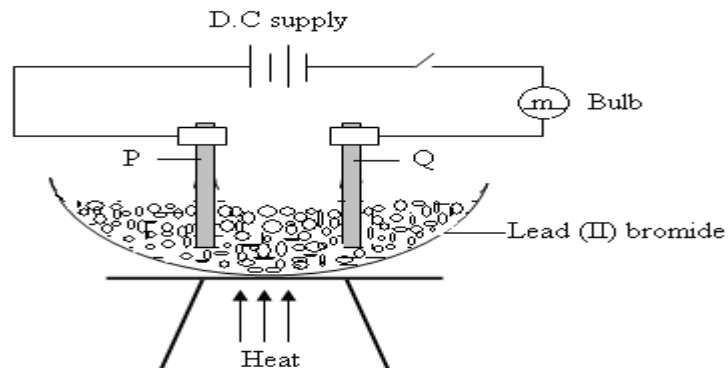
2. 1988 Q.8

A steady current of 0.65 A was passed for 35 minutes through acidified water to electrolyse it using carbon electrodes.

- (a) State the electrode at which oxygen was liberated.
- (b) Calculate the mass of oxygen liberated. ($1F = 96500 \text{ coulombs}$)

3. **1990 Q.8**

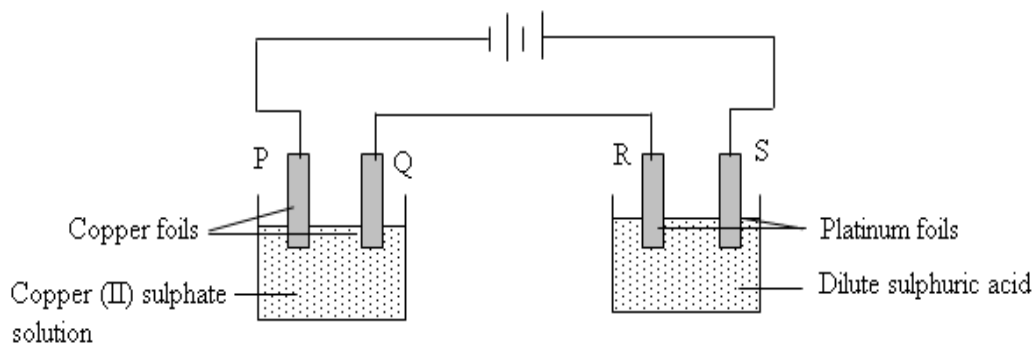
The circuit shown in the diagram in figure 3 was used in an experiment to study the effect of electricity on lead (II) bromide.



- (a) State what was observed.
 - (i) Before lead (II) bromide had melted.
 - (ii) After lead (II) bromide had completely melted.
- (b) Explain your answer in (a).
- (c) Write equation for the reaction that took place at:
 - (i) P
 - (ii) Q

4. **1992 Q.5**

When a steady current was passed through the circuit shown in the diagram below, 0.02 moles of a substance was deposited at P.



- (a)
 - (i) Name the electrodes: **P, Q, R and S.**
 - (ii) State what was observed at **R.**
- (b) Write equation(s) for the reaction(s) that took place at:
 - (i) P.
 - (ii) S
- (c) Calculate the number of moles of the gaseous substance formed at S.

5. **1993 Q.4**

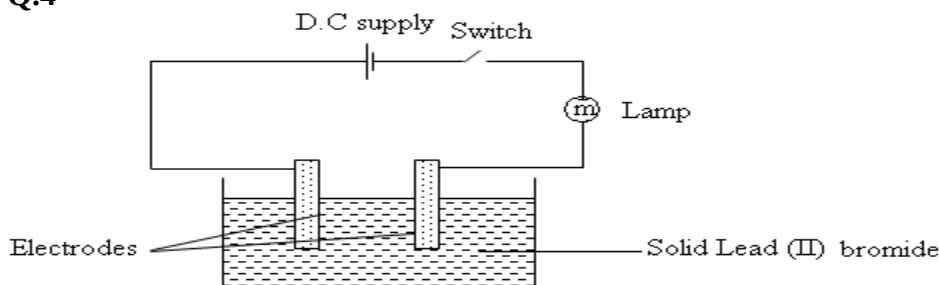
Copper (II) sulphate was electrolysed using carbon electrodes.

(a) *State what was observed at the:*

- (i) *Anode.*
- (ii) *Cathode.*

(b) *Write equation(s) for the reaction(s) that took place at the anode.*

6. **1994 Q.4**



(a) *State what was observed when:*

- (i) *The switch was turned on.*
- (ii) *Lead (II) bromide was melted and the switch turned on.*

(b) *Explain your observations in (a) (i) and (ii).*

7. **1997 Q.3**

Molten lead (II) bromide was electrolyzed between carbon electrodes.

(i) *State what was observed at the:*

Anode.

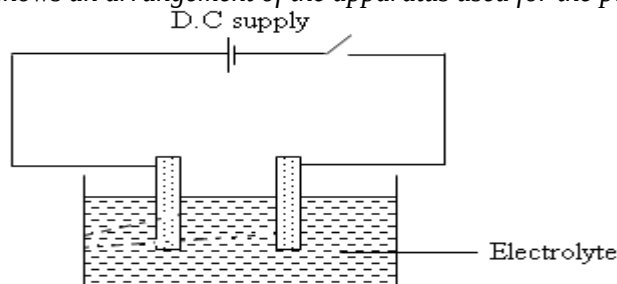
Cathode

(ii) *Write an equation for the reaction that took place at each electrode.*

(b) *Calculate the mass of the product formed at the cathode when a current of 2 amps is passed for 1 hour and 30 minutes.*

8. **2000 Q.2**

The diagram below shows an arrangement of the apparatus used for the purification of copper.



(a) *Name the substance used as:*

- (i) *anode.*
- (ii) *cathode.*

- (b) Name the electrolyte.
- (c) Write equation for the reaction that took place at:
 - (i) anode.
 - (ii) cathode.

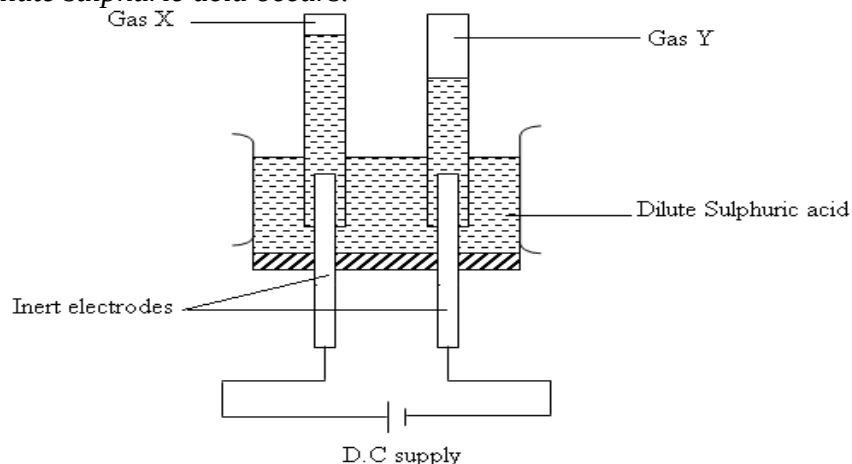
9. 2003 Q.7

Molten lead (II) bromide was electrolyzed between two carbon electrodes.

- (a) Explain why lead (II) bromide was electrolyzed in the molten state and not in the solid state.
- (b) State what was observed at the:
 - (i) anode.
 - (ii) cathode.
- (c) Write equation for the reaction that took place at the anode.

10. 2003 Q.13

- (a) The diagram below shows an electrolytic cell in which electrolysis of dilute sulphuric acid occurs.



- (i) Name the gases X and Y that are evolved during electrolysis.
- (ii) Give equation for the reaction occurring at the anode.
- (iii) Indicate the direction of electron flow in the circuit.
- (iv) Calculate the volume of gas X produced when a current of one ampere flows for 10 minutes through the electrolyte.
 [1F = 9.6×10^4 coulombs, 1 mole of a gas occupies $2.4 \times 10^4 \text{ cm}^3$ at room temperature and pressure]
- (v) State two industrial applications of electrolysis other than the manufacture of sodium hydroxide.
- (c) Sodium hydroxide can be manufactured using mercury cell. How would this manufacturing process affect the environment?

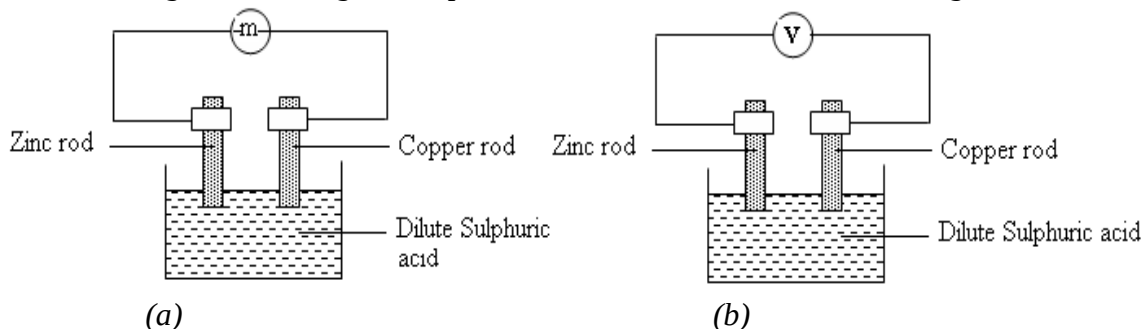
ELECTROCHEMICAL CELL

(a) **Definition:**

Electrochemical cell is a cell which changes chemical energy into electrical energy.

(b) **Structure of a Simple cell**

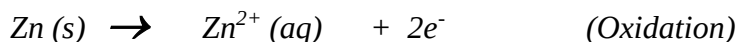
It consists of two metal strips dipped into an electrolyte and connected by pieces of conducting wires through a lamp bulb or voltmeter as shown in the diagrams below.



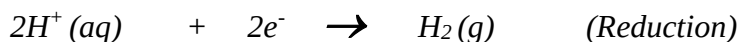
NB: In this cell, the metal higher in the electrochemical series (zinc) is the negative electrode but acts as the anode while metal lower in the electrochemical series (copper) is the positive plate of the cell and acts as the cathode.

(c) **Mechanism of the Cell**

The metal (zinc) which is higher in the electrochemical series dissolves leaving electrons on the strip and releasing Zn^{2+} ions into the electrolyte according to the equation:



The electrons travel through the external circuit and arrive at the metal lower in the electrochemical series (copper) where they are picked by hydrogen ions to form molecules of hydrogen gas according to the equation:



Observations:

In diagram (a);

- Bubbles of a colourless gas which burns with a pop sound are produced.
- The bulb lights (glows) brightly and then gradually dims until eventually it dies off.

In diagram (b);

- Bulbs of a colourless gas which burns with a pop sound are produced.
- The pointer of the voltmeter deflects through a large angle and then the angle of deflection gradually decreases until eventual it becomes zero.

Explanation:

The cell deteriorates due to polarization. **Polarization** is the building up of hydrogen bubbles on the copper strip. This prevents the hydrogen ions from being discharged quickly. The effect is that the quantity of electric current flowing is reduced. Shaking the electrode or cleaning its surface can overcome this effect.

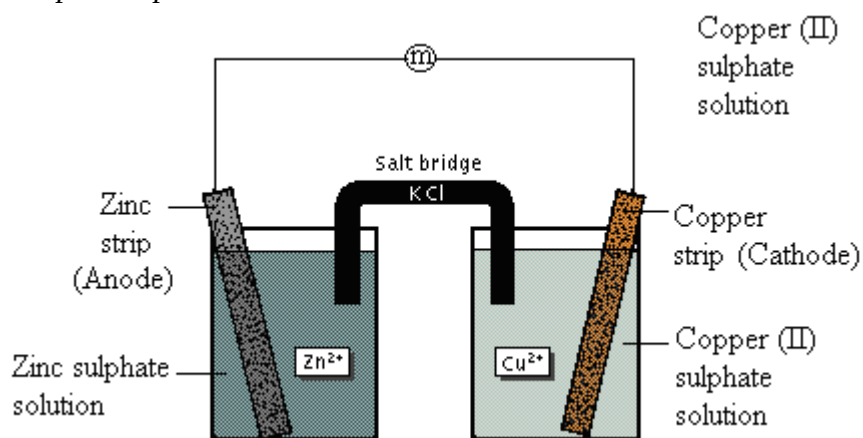
NB: The problem of polarization is best solved by dipping each metal in its own solution and the electrical contact between the electrolytes is completed by using a salt bridge, which is an inverted U-tube containing saturated potassium chloride, as seen in the Daniel Cell.

(d) **Convention for current flow**

The flow of electrons in the external circuit is from the element higher (i.e zinc) in the electrochemical series towards the one lower (i.e copper). However, since electrons are negative, current is conventionally considered to flow in the opposite direction to the direction of the flow of electrons. Current therefore flows from the metal lower in the ECS (cathode) to the metal higher in the ECS (anode) in the external circuit and vice versa in the salt bridge.

The Daniel Cell

- A Daniel Cell is a modified simple (voltaic) cell.
- It comprises electrodes immersed in solutions of their own ions and connected with a salt bridge or a porous partition.

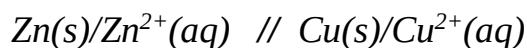


Voltaic Cell: A Chemical Battery

The zinc and copper strips act as electrodes, and the salt bridge (in this case potassium chloride) allows electrons to flow between the beakers without allowing the solutions to mix. When the circuit joining the two systems is completed (as shown on the right), the reaction generates electric current. Note that the zinc strip is used up (oxidation), and the strip appears eaten away. The copper strip is built-up as additional electrons react with the copper sulphate solution to produce additional metal (reduction).

NB: Replacing the light bulb with a battery would reverse the reaction, creating an electrolytic cell.

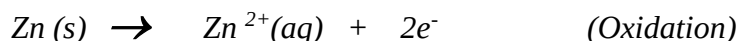
(b) Cell Representation:



- NB:**
- The metal higher in the ECS is always placed to the left hand side and the one lower in the ECS is placed to the right hand side.
 - // represents the salt bridge.

(c) Half Reaction Equations at the Electrodes:

At the anode [i.e the metal higher in the E.C.S (Zinc)]

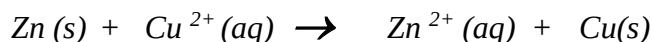
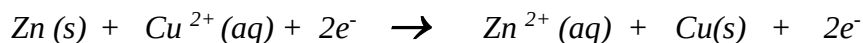


At the cathode: [i.e the metal lower in the E.C.S (Copper)]



(d) The Cell Reaction/Overall Reaction:

The overall reaction equation is obtained by putting together the reactants and the products of the half reaction equations and the electrons which appear on both sides canceled.



(e) Position of the electrodes in the E.C.S.

If zinc/zinc sulphate solution is replaced by magnesium/magnesium sulphate solution, a greater voltage is obtained. This is because the difference between the electrode potentials is greater for Mg/Mg^{2+} and Cu/Cu^{2+} than for Zn/Zn^{2+} and Cu/Cu^{2+} .

Electrode potential

Electrode potential of an element is the potential difference between the element and a solution of its own ions. The electrode potentials are negative for metals above hydrogen and positive for metals below hydrogen in the electrochemical series.

Energy Supplied by the Cell (Not so important at this level)

The energy supplied by the cell is given by the equation:

$$\begin{array}{lll} \text{Energy} & = & \text{Voltage} \times \text{Charge} \\ (\text{Joules}) & & (\text{Volts}) \quad (\text{Coulombs}) \end{array}$$

But **Coulombs = nF**

Where; n = Number of electrons and
 F = Faraday (96500C).

If voltage = E , then:

$$\text{Energy} = \text{EnF}$$

Example

A voltmeter in a Daniel cell reads 1.20V for a Zn/Cu couple. Find the electrical energy supplied by the cell. ($1F = 96500C$)

Solution: $E = 1.20 \text{ V}$, $n = 2$ (Valence of Zn and Cu = 2)

$$\begin{array}{l} \text{From:} \quad \text{Energy} = \text{EnF} \\ \quad \quad = 1.20 \times 2 \times 96500 \\ \quad \quad = \mathbf{231.6 \text{ KJ}} \end{array}$$

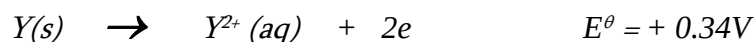
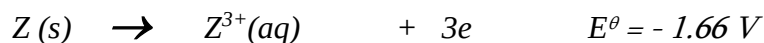
Electromotive Force (e.m.f) generated by the Cell

The e.m.f generated by the cell is given by the formula:

$$\text{E.m.f} = E^{\theta}_{\text{right}} - E^{\theta}_{\text{left}}$$

Example:

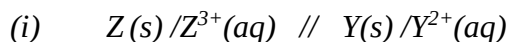
The following are half reactions for metals Y and Z.



- (i) Draw a chemical cell for the metals Z and Y.
- (ii) Calculate the e.m.f of the cell.

Recall that:

- ❖ Reactive metals above H have negative electrode potential and are always placed on the left in the cell representation.
- ❖ Less reactive metals below H have positive electrode potential and are always placed on the right hand side of the cell.

Solution:

$$(ii) \quad E.m.f = E^\theta_{\text{right}} - E^\theta_{\text{left}}$$

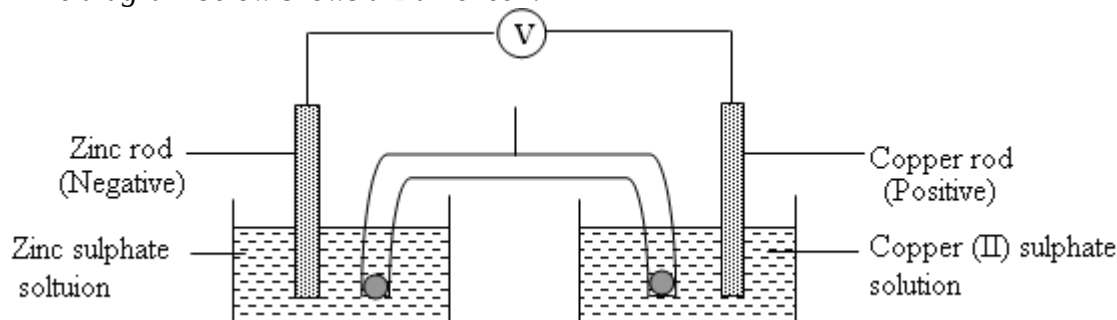
$$= +0.34 - (-1.66)$$

$$= 0.34 + 1.66$$

$$= 2.00 \text{ V}$$

SELF-CHECK 2.2

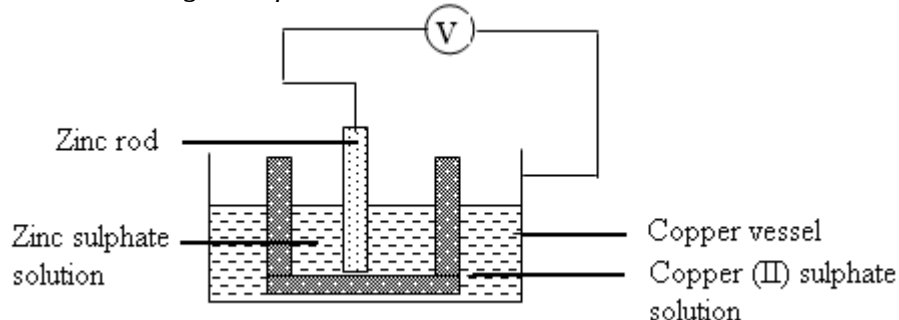
1. The diagram below shows a Daniel cell.



- (a) Indicate with an arrow on the diagram the direction of flow of electrons.
- (b) What is the purpose of the salt bridge?
- (c) Write an equation for the reaction taking place at:
 - (i) Zinc rod
 - (ii) Copper rod.
- (d) The voltmeter connected across the cell showed a reading of 0.90 V. Calculate the energy supplied by the cell. ($1F = 96500C$)
- (e) If the Zn/Zn^{2+} in the left hand side beaker is replaced by Mg/Mg^{2+} , would the reading on the voltmeter be less than, equal to or greater than 0.90 volts?
- (f) Give a reason for your answer in (e).

2. **UNEB 1991 Q.4**

Figure 1 shows a diagram of an electrochemical cell.



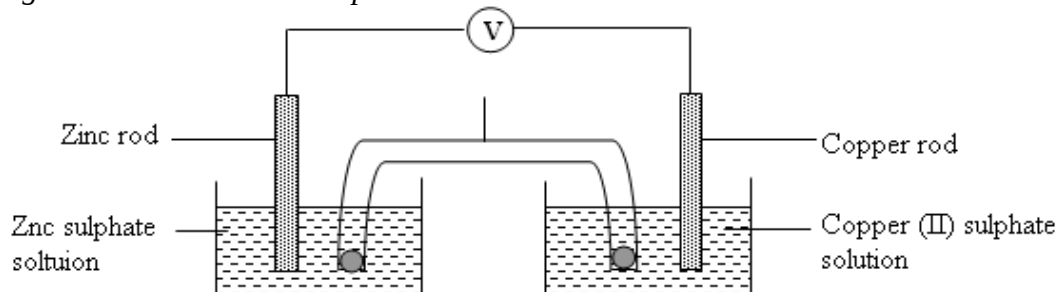
- (a) (i) Write an equation for the overall reaction.
 (ii) State what would be observed if the reaction is allowed to continue for a long time.
- (b) The reading on the voltmeter, V , was 1.10 V. Calculate the energy in KJ produced.

3. **UNEB 1997 Q.14**

- (a) Draw a diagram of Daniel cell consisting of a zinc rod dipped in zinc sulphate solution and a copper rod dipped in copper (II) sulphate solution; the solutions separated by a porous wall and the rods are connected by conducting wire.
- (b) Indicate: (i) the charges on each electrode.
 (ii) the direction of electron movement in the wire.
- (c) Write: (i) equations at each electrode.
 (ii) an equation for the overall reaction.

4. **(1998 Q.5)**

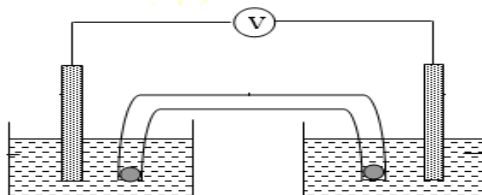
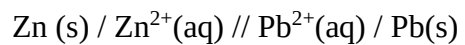
Figure 2 below shows a simple voltaic cell.



- (a) (i) Write equation for the reaction taking place at:
 – the cathode
 – the anode
 (ii) Write the overall equation of the cell reaction.
- (b) Draw an arrow on the diagram to show the direction of flow of electrons.

5. (2002 Q.4).

The cell conversion for an electrochemical cell is shown below.



- (a) Name **two** substances that could be used as electrolytes.
- (b) State which one of the electrodes is the anode.
- (c) Write equation for the reaction at.
 - (i) The anode.
 - (ii) The cathode.
- (d) Write equation for the overall cell reaction.

CHAPTER THREE

RATE OF REACTION

Rate Reaction of reaction - is the amount of reactant(s) used up per unit time.
Or - is the amount of product(s) formed per unit time.

i.e.
$$\text{Rate} = \frac{\text{Amount}}{\text{Time}}$$

The units of rate are: - $\text{mol l}^{-1}\text{s}^{-1}$
- $\text{g l}^{-1}\text{s}^{-1}$

Determination of rate of reaction

Consider the reaction between calcium carbonate and dilute hydrochloric acid.

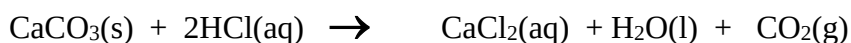
The determination of rate of reaction can be done by using one of the following methods:

- (i) Measuring the volume of a gas evolved with time.
- (ii) Measuring the decrease in mass of the reactants with time.
- (iii) Measuring the rate at which a colourless substance changes to coloured or the rate at which a coloured substance changes to colourless.

Consider the reactions below.

(i) Reaction between Calcium carbonate and Hydrochloric acid

The equation for the reaction is:



Facts about the reaction:

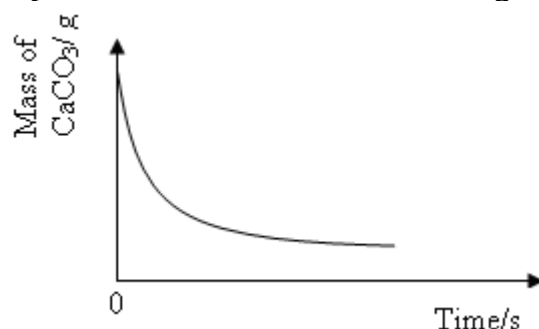
- If the reaction takes place in an open container e.g. a beaker and the reaction vessel is placed on a weighing machine, the total mass of reaction decreases with time. This is due to the loss of carbon dioxide to the atmosphere.
- If the reaction takes place in a closed system where the gas collects in a syringe, the volume in the syringe is seen increasing with time.

If the results are entered in a table and then plotted against time, the graph obtained is called the **Rate Curve**.

The Rate Curves

- (a) When the mass of reactant is plotted against time, the shape of the graph below is obtained.

The graph of mass of calcium carbonate against time

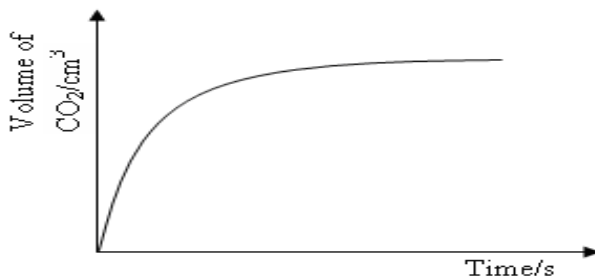


Interpretation of the shape of the graph

At time $t = 0$, the mass of the reactants is high. As time increases, the calcium carbonate is used up and the mass of the reactants drops very fast at first and the drop reduces as time increases. When all the calcium carbonate is used up, the mass remains constant.

- (b) When the volume of gas evolved is plotted against time, the shape of the graph below is obtained.

The graph of volume of carbon dioxide against time



Interpretation of the shape of the graph

A rate curve shows that the rate of chemical reaction is high at the beginning, then decreases with time and finally becomes zero.

Explanation

At the beginning of the reaction, the concentration of the reactants is high. As a result the frequency of collision of the reacting particles is high. Hence the reaction is fast. But as the reactants get used up the rate of reaction decreases with time (I.e. slows down) until finally it becomes zero. The reaction stops when one of the reactants is used up, hence the volume remains constant

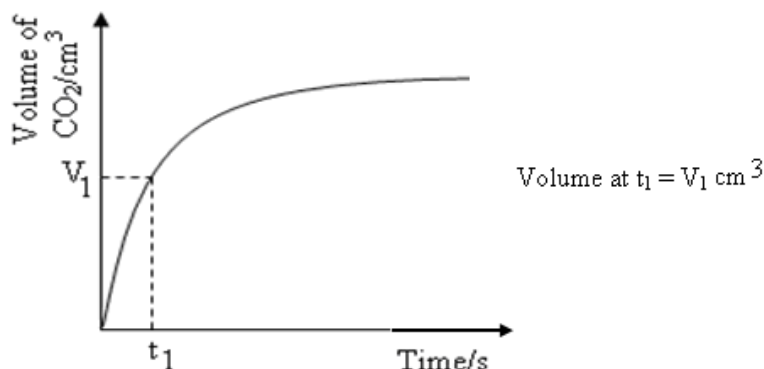
Reading of volume of gas from the graph

(a) Reading volume of a gas at a given time

To find the volume of a gas produced at a given time, say, t_1 from the graph, draw a perpendicular line to the time axis at the time given, to meet the curve.

Read off the volume, v_1 , where the perpendicular line meets the rate curve.

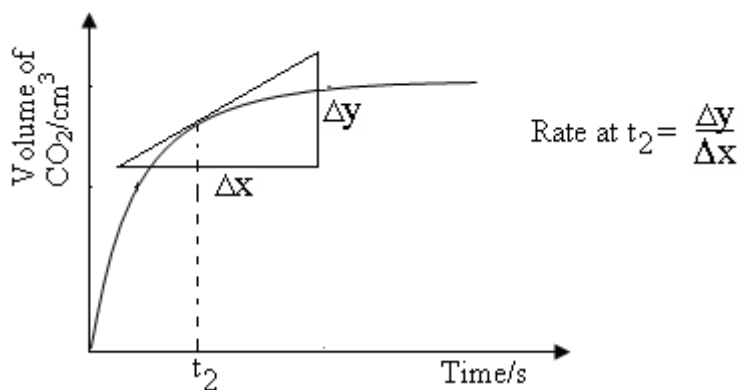
The graph of volume of carbon dioxide against time



(b) To determine the rate at a given time

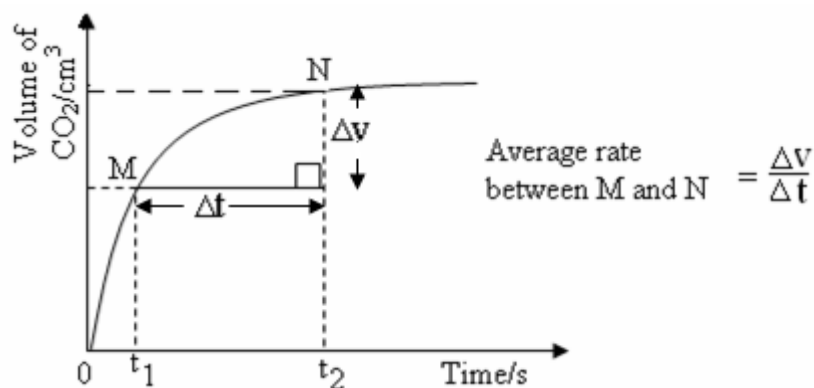
To obtain the rate of reaction at time, t_2 draw a tangent to the curve at the time, t_2 , given. Complete the tangent to obtain a right angle triangle.

Find the gradient of the tangent, $\frac{\Delta y}{\Delta x}$, gives the rate of the reaction at time, t_2 , given.

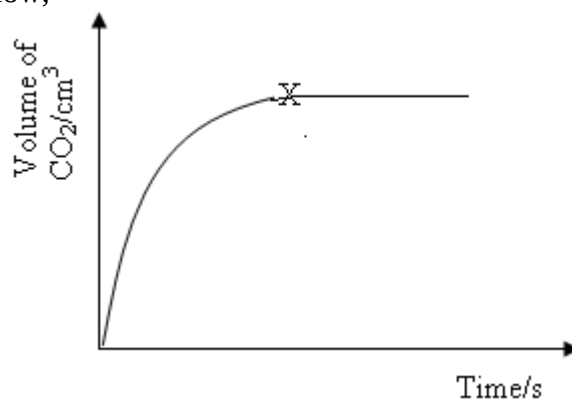


(c) To determine the average rate between two points

To find the average rate between points M and N, divide the change in volume by the change in time i.e. $\frac{\Delta v}{\Delta t}$. Units are cm^3s^{-1} .



- (d) **End of the reaction** - is the point at which the graph starts to level off. i.e point X in the graph below,



NB:

- The other substances that may be used instead of Calcium carbonate and dilute hydrochloric acid are:
- Magnesium and Hydrochloric acid or dilute sulphuric acid.
- Zinc and dilute Hydrochloric or dilute sulphuric acid.
- Hydrogen peroxide and Manganese (IV) oxide.

Factors affecting rates of reaction

The factors that affect the rate of Chemical reactions include the following:

- (i) Concentration of reactants.
- (ii) Temperature.
- (iii) Pressure.
- (iv) Catalyst.
- (v) Surface area of the reactants.
- (vi) Light.

The study of rates of reactions is known as **Chemical Kinetics**. The explanations for many of the effects of the above factors are based on the Kinetic Theory.

1. The effect of concentration of reactants on the rate of reaction

Experiment 1

To investigate the effect of concentration on the rate of reaction

Sodium thiosulphate reacts with hydrochloric acid to form sulphur which appears as yellow colouration according to the following equation:



Since this reaction produces a precipitate from two colourless solution, the intensity of the precipitate at any given moment in time represents the extent of the reaction.

Procedure

- Place 50 cm³ of sodium thiosulphate solution (containing 40 g dm⁻³ of the compound) in 100 cm³ beaker.
- Add 5 cm³ 2M hydrochloric acid and at the same time start a stop-clock, or note the time on watch reading in seconds.
- Swirl the beaker carefully a couple of times to mix it and place it on a white piece of paper with a cross drawn on it.
- Observe the cross by looking down through the solution from above the beaker and stop the clock (or record the time) at the moment the cross is invisible.
- Take 40, 30, 20 and 10 cm³ of the original sodium thiosulphate solution and make the total volume up to 50 cm³ by adding distilled water in different beakers. (This is to ensure the same depth of solution in the beaker for each reaction, so that the time of disappearance of the cross is made at the same stage in each of the reactions.)
- For each, solution, repeat the procedure (b) to (d).
- Enter your results in the table below.

Volume of original Sodium thiosulphate solution (cm ³)	10	20	30	40	50
Time (s)					

NOTE: *You should notice from your table of results that as the concentration of thiosulphate solution increases the time taken for the disappearance of the cross decreases.*

Conclusion: The rate of reaction increases with increase in concentration.

Explanation

The rate of reaction therefore depends upon the frequency with which the particles collide, which in turn depends on their concentration.

With increase in concentration, there are more reacting particles so the chances of collision and hence the rate of reaction is high. I.e. the more crowded the particles are the more often they bump into one another.

2. The effect of temperature on the rate of reaction

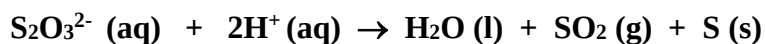
Experiment 2 To investigate the rate of reaction by using sodium thiosulphate

You are provided with the following:

BA1, which is a sodium thiosulphate solution.

BA2, which is a dilute hydrochloric acid.

Sodium thiosulphate reacts with hydrochloric acid to form sulphur which appears as yellow colouration according to the following equation:



The rate of the reaction at a particular temperature can be followed by noting the time taken for the yellow coloration to appear at that temperature.

You are required to investigate how the rate of the reaction varies with the temperature for the reaction.

Procedure:

- Mark a small cross (X) with a blue/black pen on a sheet of white paper provided and place it on the table.
- Place a 250 cm³ conical flask right onto the cross.
- Using a measuring cylinder, transfer 50 cm³ of **BA1** into the conical flask which is on the cross.
- Using another measuring cylinder, measure 5.0 cm³ of **BA2** and add at once to the solution of **BA1** in the conical flask, and **at the same time start the stop clock (watch)**. Shake to mix and place the flask over the cross.
Note and record the temperature of the mixture. View the cross from above through the mixture. **Note and record** the time taken for the yellow coloration to just make the cross invisible. (This is the time, **t** (in seconds) for the reaction to occur at room temperature.)
- Transfer a fresh 50 cm³ of **BA1** into the conical flask, and heat the solution to 30 °C.

- (f) Add 5.0 cm³ of **BA2** to the hot solution and the same time start the Stop-clock (watch).
- (g) Shake to mix and place the flask over the cross.
- (h) Look at the cross from above through the mixture.
- (i) **Note and record** the time taken for the yellow coloration to just make the cross invisible. (This is the time, **t** (in seconds) for the reaction to occur at 30 °C temperature.)
- (j) Repeat procedures (e) to (i) for the reaction at 40, 50 and 60 °C respectively. Record your results in the table below.

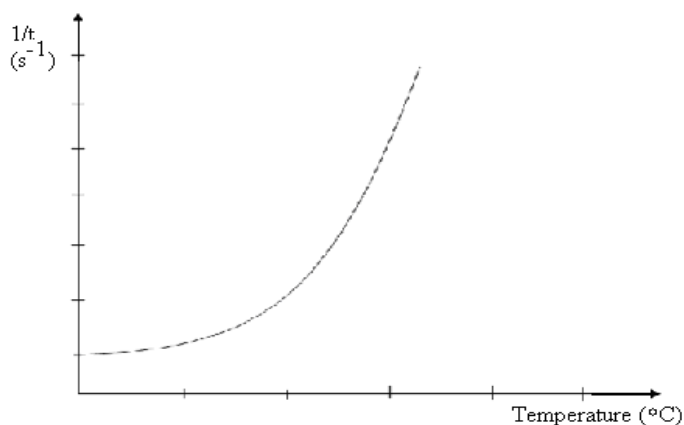
Spacemen results:

Temperature (°C)	Room temperature 25	30	40	50	60
Time, t, for yellow coloration to cover the cross (s)	46	45	25	16	10
$\frac{1}{t}$ (s ⁻¹)	0.02	0.02	0.02	0.06	0.1

Questions:

- (k) Calculate the reciprocal of the time ($\frac{1}{t}$) for each reaction temperature and record the values in the table above.
- (l) Plot the graph of $\frac{1}{t}$ (along the vertical axis) against temperature (along the horizontal axis).

The sketch of the Graph of $\frac{1}{t}$ against temperature



- (m) State how the rate of reaction varies with temperature.
The rate of reaction increase with temperature.

Explanation

When the temperature of a reaction is increased, heat energy is supplied to the reacting particles involved in the reaction. As the particles acquire K.E they move faster and thus collide with one another at more frequent intervals, hence the reaction proceeds at a faster rate.

Or

As temperature increases, the total energy of molecules increases so the number of molecules having the activation energy required for the reaction also increases.

3. The effect of Surface area of a reactant on the rate of reaction

For Heterogeneous reaction (reaction where the reactants are at the different states), particle size influences the rate of reaction. For example the reaction between marble chips and dilute hydrochloric acid shows that, the smaller the pieces of marble the greater the rate of reaction. This is attributed at the greater surface area of the marble for attack by the acid in the case of the smaller pieces.

In other words, a powder has a larger total surface area for reaction than granules. So more particles react at any given time.

4. The effect of Catalyst on the rate of reaction

A catalyst is a substance which alters the rate of a chemical reaction and remains chemically unchanged at the end of the reaction.

*A catalyst usually increases the rate of reaction by lowering the activation energy so, although energy remains the same, the energy of a greater number of molecules now exceeds the new activation. This is called **positive catalysis**.*

5. The effect of light on the rate of reaction

Light is source of energy and can influence the rate of some chemical reactions by energizing some of the molecules involved.

For example the reaction between chlorine and hydrogen at ordinary pressure is negligible in darkness. Slow in diffuse light but explosive in sun light (at room temperature).

Light is also vital factor in photosynthesis.

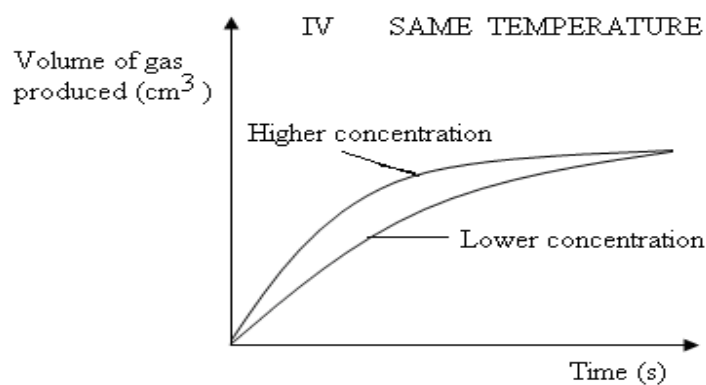
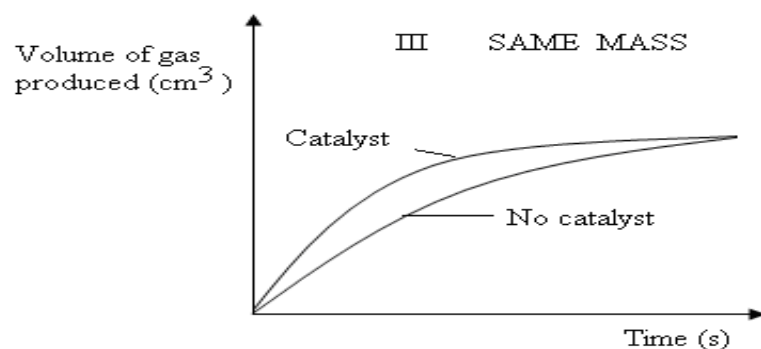
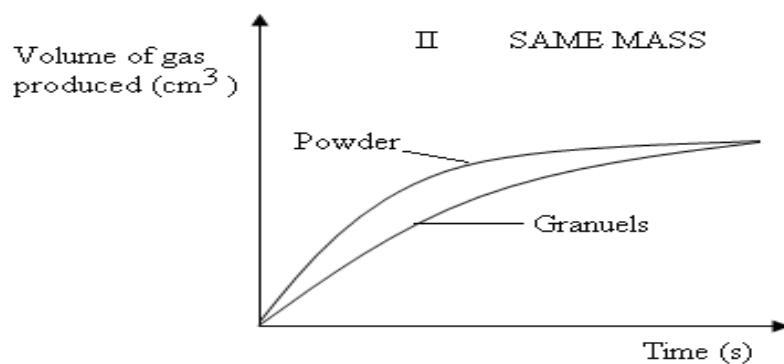
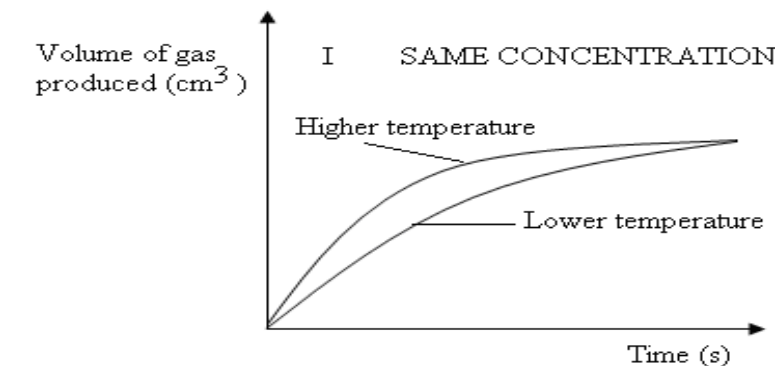
One of the commonest examples of the effect of light on the rate of chemical reaction is in photography.

6. The effect of Pressure on the rate of reaction

Pressure influences the reactions in which the reactants are gases.

High pressure brings gas molecules closer so they collide more often, leading to faster chemical reaction.

Typical results of graphs obtained by taking the volume of gas produced against time for different factor which determine the rate of reaction.



REACTION RATES

SELF-CHECK 3.1

- (2001 Q.15).** A catalyst is a substance which
 - controls the concentration of reactants.
 - stops a reaction from becoming violent.
 - increases the rate of a chemical reaction.
 - neutralises a base.
- (2000 Q.14).** The rate of the chemical reaction between calcium carbonate and hydrochloric acid can be determined by the
 - concentration of carbon dioxide produced.
 - temperature of carbon dioxide produced.
 - volume of carbon dioxide produced.
 - pressure of carbon dioxide produced.
- (2000 Q.28).** Increasing concentration increases the rate of a reaction because the particles
 - move faster.
 - collide more often.
 - have more energy.
 - collide with more force.
- (1999 Q.27).** The graph in figure 1 shows the variation in the volume of hydrogen evolved with time when Zinc reacts with dilute sulphuric acid using Copper (II) sulphate as a catalyst.

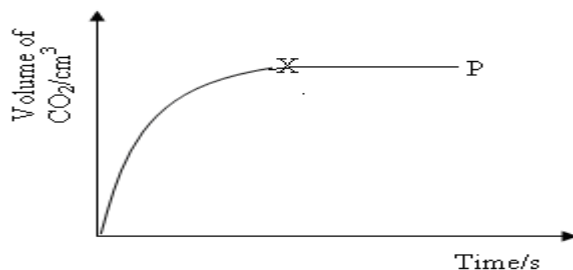
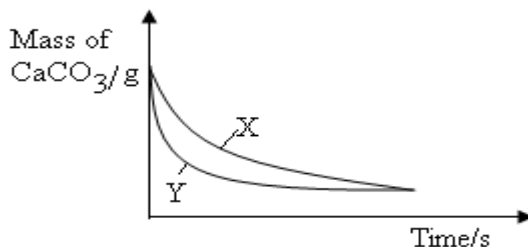


Fig. 1

- The best explanation for the shape of the graph between X and P is
- the zinc is used up.
 - the products stopped the reaction.
 - sulphuric acid is used up.
 - the catalyst is used up.
- (1996 Q.21).** The graph X below shows the variation in mass with time when 50 g of calcium carbonate powder was reacted with excess 1M HCl at 25°C.

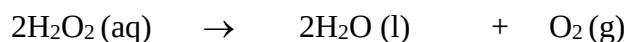


To obtain graph Y for the same reaction, one would keep all other conditions of the reaction the same but

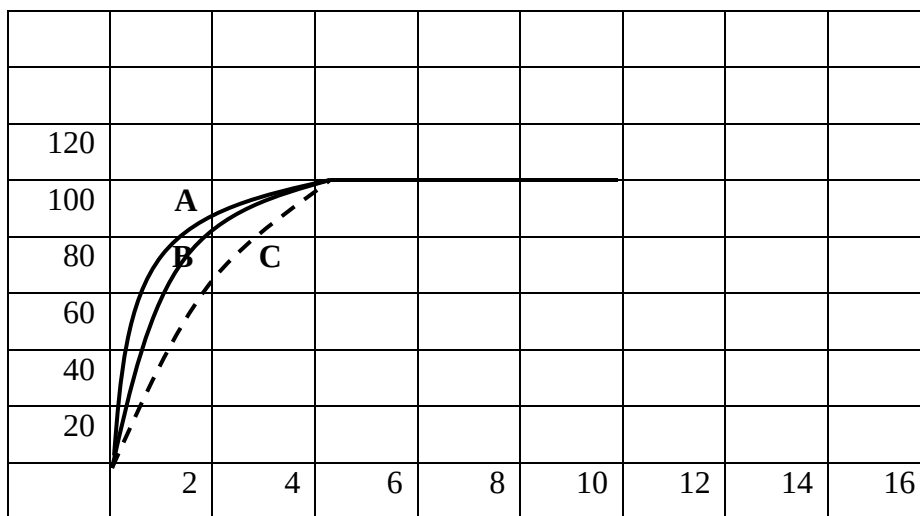
- use 2M hydrochloric acid.
- use 50 g of calcium carbonate lumps.
- reduce the temperature to 12,5°C.
- use 25 g of calcium carbonate powder.

SELF-CHECK 3.2

1. (a) Hydrogen peroxide, H_2O_2 , decomposes in aqueous solution to give oxygen gas. The equation for the reaction is:



The reaction is catalyzed by manganese (IV) oxide. In two separate experiments 100 cm^3 of solution of hydrogen peroxide was stirred with the same quantity of manganese (IV) oxide at 20°C . In the first experiment coarse particles of the catalyst were used, while in the second experiment a finely powdered form of the catalyst was used. The volume of oxygen produced was measured at 1 minute intervals. The results obtained are shown on graphs labeled **A** and **B** below.



- (i) State which graph corresponds to the coarse particles of catalyst. Give reason for your answer.
- (ii) In both experiments, the volume of oxygen evolved per minute decreased with time. Give an explanation to this.
- (iii) If experiment B were carried out at 10°C instead of 20°C , sketch on the graph the type of curve you would expect to obtain (label it **C**) and give an explanation to support it.
- (iv) Why do all the graphs level at the same value (100 cm^3)?

EXTRA STUDY QUESTIONS

(For better results plot the required graphs)

2. (2003 Q.11).

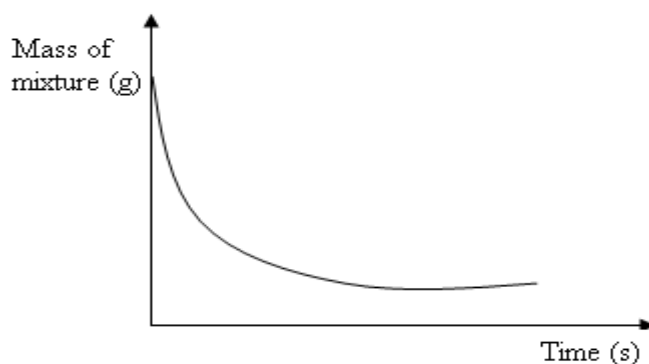
- (a) What is meant by **rate of a chemical reaction**?
- (b) State how the following factors affect the rate of a chemical reaction:
- (i) Temperature.
 - (ii) Surface area of the reactants.
- (c) The table below shows the volume of hydrogen collected at various time intervals when magnesium was reacted with a 2M hydrochloric acid.

Time (s)	0	1	2	3	4	5	6	7
Volume of hydrogen collected (cm ³)	0	25	45	60	70	75	77	77

- (i) Plot a graph of volume of hydrogen versus time.
 - (ii) Determine the rate of the reaction at 3 seconds.
 - (iii) Determine the volume of hydrogen evolved at 3.5 seconds.
- (d) State how the rate of the reaction at 3 seconds would be affected if a 1M hydrochloric acid was used.

3. (1999 Q. 2).

- (a) State the factors that can affect the rate of a chemical reaction.
- (b) A mixture of a known mass of magnesium and a certain volume of 2M hydrochloric acid was put in a conical flask and the mass of the mixture was recorded at various intervals. The result of the experiment is shown in the graph below.



On the same axes, draw a graph that would be obtained when same mass of magnesium was reacted with the same volume of 1 M hydrochloric acid.

- (c) 5.0 g of calcium carbonate was reacted with 20 cm³ of 2 M hydrochloric acid.
- (i) Write the equation for the reaction between hydrochloric acid and calcium carbonate.
 - (ii) Calculate the mass of calcium carbonate that was left.

(Ca = 40; O = 16; C = 12)

4. (1999 Q.11).

- (a) (i) What is the rate of *reaction*?
(ii) How does particle size affect rate of reaction?
Explain your answer.
- (b) The table of results below shows the time taken for Sulphur to form when various concentrations of sodium thiosulphate were used.

Concentration of $\text{S}_2\text{O}_3^{2-}$ (M)	0.2	0.6	0.8	1.2	1.6
Time for sulphur to form (Sec)	60	20	15	10	7.5
$1/t$ (Sec^{-1})	0.017	0.05	0.07	0.10	0.13

Plot the graph of $1/t$ (sec^{-1}) vertical against concentration of thiosulphate.

- (c) (i) Explain the relationship between a rate of the reaction and $1/t$.
(ii) Deduce from the graph, how the rate of reaction varies with the concentration of thiosulphate.
- (d) Name **one** reagent that you would use to test for sulphur dioxide and state what would be observed if the reagent was used.

5. (1999 Q.12).

In an experiment to determine the rate of reaction between zinc and sulphuric acid, dilute sulphuric acid was reacted with zinc granules to which some copper (II) sulphate solution was added. The volumes of hydrogen gas evolved at various times were measured. The results are shown in the table below.

Time in minutes	0	5	10	15	20	25	30
Volume of gas in cm^3	0	10	20	25.5	29.5	32	32

- (a) (i) What is the role of copper (II) sulphate solution?
(ii) Write an ionic equation for the solution in the reaction above.
(iii) Explain what would happen to the reaction if zinc granules were replaced with zinc powder.
- (b) (i) Plot the graph of volume of hydrogen evolved (vertical) against time.
(ii) Describe how you would determine the rate of the reaction at 12 minutes.
(iii) Compare the rate of reaction at 12 minutes with that of 20 minutes. Give a reason for your answer.
(iv) What happens to the shape of the graph after 25 minutes? Explain your answer.

CHAPTER FOUR

ORGANIC CHEMISTRY

Organic chemistry is a branch of chemistry which deals with the study of carbon and its compounds.

The compounds referred to here exclude:

- the oxides of carbon, Carbon dioxide (CO_2) and carbon monoxide, (CO),
- the metallic carbonates and hydrogen carbonates(bicarbonates).

Modern chemists consider organic compounds to be those containing carbon and one or more other elements, most often: hydrogen, oxygen, nitrogen, sulphur, or the halogens, but sometimes others as well.

Uniqueness of Carbon

Carbon is a unique element. Its atoms have the ability to join and form straight chains, branched chains and rings. The unique properties of carbon are manifest in the simplest class of organic compounds, the aliphatic, called *hydrocarbons*.

Hydrocarbons

Hydrocarbons are organic compounds that contain carbon and hydrogen only.

The general formula of hydrocarbons is C_xH_y

Where x and y are whole numbers.

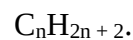
Classes of Hydrocarbons

Hydrocarbons are classified into:

- Alkanes,
- Alkenes and
- Alkynes.

Alkanes

Alkanes are saturated hydrocarbons with the general formula



Where $n = 1, 2, 3 \dots$ for successive members of the group.

The table below shows the molecular formula and names of the first ten members of the alkanes class

Value of n	Molecular formula	Name
1	CH ₄	Methane
2	C ₂ H ₆	Ethane
3	C ₃ H ₈	Propane
4	C ₄ H ₁₀	Butane
5	C ₅ H ₁₂	Pentane
6	C ₆ H ₁₄	Hexane
7	C ₇ H ₁₆	Heptane
8	C ₈ H ₁₈	Octane
9	C ₉ H ₂₀	Nonane
10	C ₁₀ H ₂₂	Decane

Structure of Alkanes

A saturated compound is a compound in which all the valency electrons are used in bonding.

In alkanes, carbon atom(s) exercise a valency of four (4). I.e. all the valency electrons are shared with other atoms to form covalent bonds. Therefore each carbon atom is surrounded by four covalent bonds, thus making alkanes saturated compounds.

Molecular formula, Structural formula and Displayed formula

(i) Molecular formula

The molecular formula of a compound is the formula that indicates the number of each kind of atom in a molecule of that substance. E.g. The molecular formula of butane (C₄H₁₀), shows that, there are four carbon atoms chemically bonded to 10 atoms of hydrogen.

(ii) Structural formula

The structural formula of a compound is the formula that shows the sequence and arrangement of the atoms in a molecule of the compound.

Name	Molecular formula	Structural formula
------	-------------------	--------------------

Ethane	C_2H_6	CH_3CH_3
Propane	C_3H_8	$CH_3CH_2CH_3$
Butane	C_4H_{10}	$CH_3CH_2CH_2CH_3$

(iii) Displayed (Graphic) formula

The displayed formula of a compound is the formula that shows the sequence and arrangement of the atoms and their bonds in a plane.

Name	Molecular formula	Structural formula
Methane	CH_4	$ \begin{array}{c} H \\ \\ H - C - H \\ \\ H \end{array} $
Ethane	C_2H_6	$ \begin{array}{c} H \quad H \\ \quad \\ H - C - C - H \\ \quad \\ H \quad H \end{array} $
Propane	C_3H_8	$ \begin{array}{c} H \quad H \quad H \\ \quad \quad \\ H - C - C - C - H \\ \quad \quad \\ H \quad H \quad H \end{array} $
Butane	C_4H_{10}	$ \begin{array}{c} H \quad H \quad H \quad H \\ \quad \quad \quad \\ H - C - C - C - C - H \\ \quad \quad \quad \\ H \quad H \quad H \quad H \end{array} $

Homologous Series

Homologous Series Is the family of related carbon compounds.

Compounds of the same homologous series are referred to as members.

Properties of the Members of Homologous series

- (i) They have the same general formula.
- (ii) They have the same name ending.
E.g. - *ane* – for alkanes,
 - *ene* – for alkenes and
 - *yne* – for alkynes.
- (iii) They differ from each other by a methylene group (- CH_2 -).

- (iv) They show a gradation of physical properties due to successive increase in molecular mass.
- (v) They can be prepared in the laboratory by the same general method.
- (vi) They have similar chemical properties.

Isomers and Isomerism

(i) Isomers

Isomers are compounds having the same molecular formulae but different structural formulae.

(ii) Isomerism

Isomerism is the occurrence of two or more compounds with the same molecular formula but different structural formulae.

The table below shows some isomers and their formulae

Name	Molecular formula	Structural formula	Displayed formula
Butane	C ₄ H ₁₀	CH ₃ CH ₂ CH ₂ CH ₃	<pre> H H H H H - C - C - C - C - H H H H H </pre>
2-methylpropane	C ₄ H ₁₀	<pre> CH₃CHCH₃ CH₃ </pre>	<pre> H H H H - C - C - C - H H H H </pre>

Properties of Alkanes

(a) Physical Properties

- (i) The
 - lower (first four) members are gases.
 - next twelve (C₅H₁₂ to C₁₆H₃₄) are liquids.
 - rest are solid e.g. paraffin wax.
- (ii) All are insoluble in water but soluble in organic solvents e.g. ethanol, benzene, toluene etc.
- (iii) All have low densities, less than the density of water and the densities rise gradually with relative molecular mass.

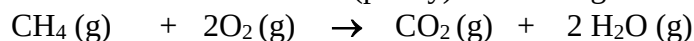
(b) Chemical properties

(i) Combustion.

Combustion is divided into *complete* and *incomplete combustion*.

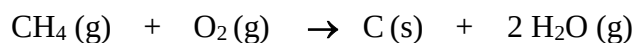
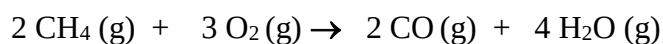
➤ **Complete Combustion:**

Alkanes burn in sufficient (plenty) air forming carbon dioxide and water.



➤ **Incomplete Combustion:**

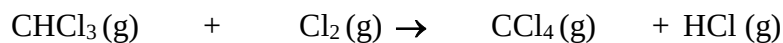
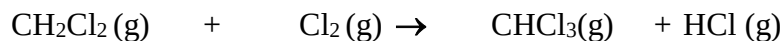
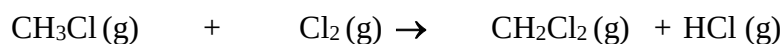
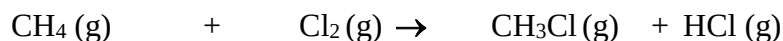
In insufficient (limited) air, alkanes burn forming carbon monoxide or carbon and water.



(ii) Substitution Reaction

Since alkanes are saturated compounds, they undergo *substitution reaction* with halogens. The halogen atoms such as chlorine and bromine substitute one or more of the hydrogen atoms.

E.g. Reaction with chlorine in diffused sunlight.



Uses of Alkanes

- They are used :
- (i) As fuel for cooking and lighting.
 - (ii) For making Lubricants.
 - (iii) For making Printer's ink.
 - (iv) For making Paints.
 - (v) For making Carbon paper.

Alkenes

Alkenes are unsaturated hydrocarbons with the general formula C_nH_{2n} .

Where $n = 2, 3, 4 \dots$ for successive members.

This family of hydrocarbons is characterized by one or more double bonds between two carbon atoms ($C=C$). Due to the double bond, alkenes are reactive. This is because the double bond is readily converted to a single bond by addition of other atoms.

Structure of Alkenes

The table showing the molecular formula and names of the first three members of alkenes class

Value of n	Name	Molecular formula	Structural formula	Displayed formula
2	Ethene	C_2H_4	$CH_2=CH_2$	<pre> H H \ / C=C / \ H H </pre>
3	Propene	C_3H_6	$CH_3CH=CH_2$	<pre> H H H H-C - C = C H H </pre>
4	But-1-ene	C_4H_8	$CH_3CH_2CH=CH_2$	<pre> H H H H H-C - C - C = C H H H H </pre>
	But-2-ene		$CH_3CH=CHCH_3$	<pre> H H H H H-C - C = C - C-H H H </pre>

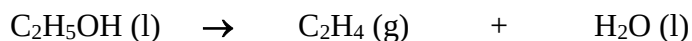
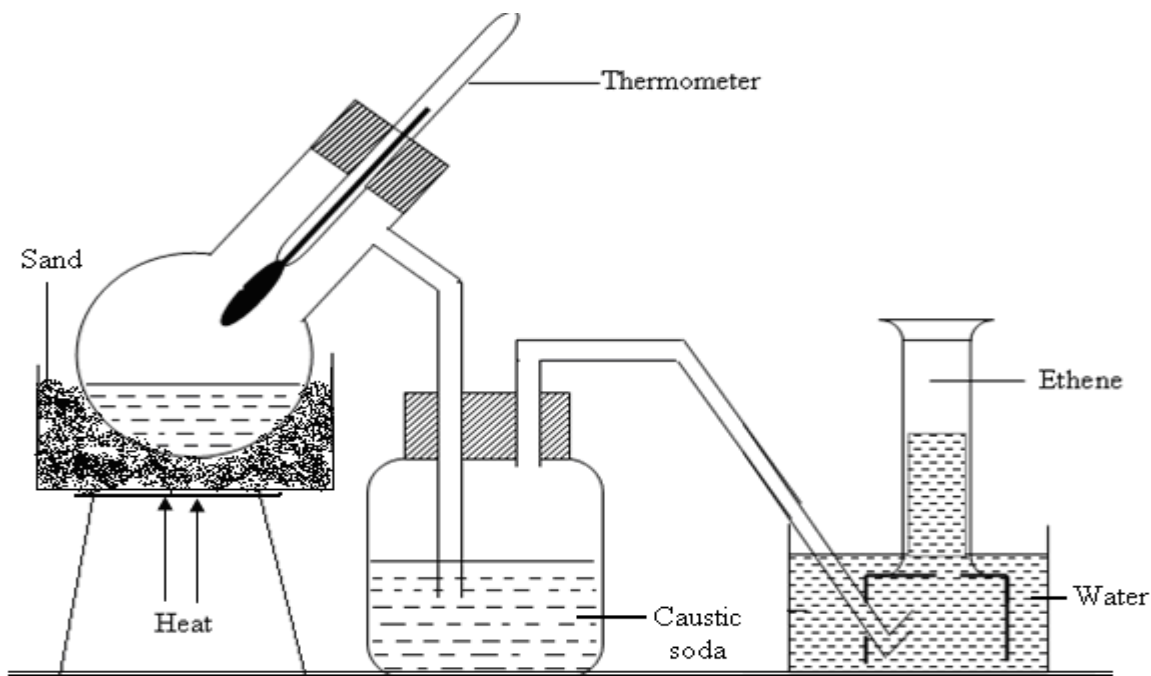
- NB:**
1. *But-1-ene and But-2-ene are isomers.*
 2. *The most important alkene is ethene.*

Ethene

Laboratory Preparation of Ethene

Ethene is prepared in the laboratory by dehydration of ethanol. A mixture of concentrated sulphuric acid and ethanol is heated to a temperature of 170 °C. The gas is passed through solution of caustic soda to remove sulphur dioxide and is collected over water.

The Diagram Showing Lab Preparation of Ethene



Properties of Ethene

(a) Physical properties

- (i) It is a colourless gas.
- (ii) It is slightly soluble in water.
- (iii) It has a faint smell.
- (iv) Its density is about the same as that of air.
- (v) It undergoes polymerization.

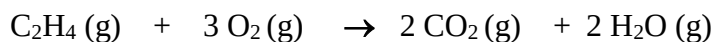
(b) Chemical Properties

(i) Combustion

Combustion is divided into complete and incomplete combustion.

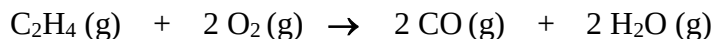
➤ **Complete Combustion:**

Ethene burns in sufficient (plenty) air forming carbon dioxide and water according to the equation.



➤ **Incomplete Combustion:**

In insufficient (limited) air, ethene burns forming carbon monoxide or carbon and water.

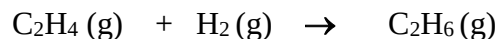


(ii) Addition reaction of ethene

Due to the double bond in ethene, it undergoes addition reaction with hydrogen and halogens.

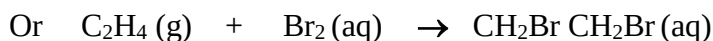
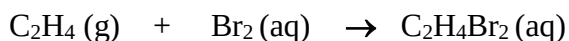
➤ **Hydrogenation:**

Ethene reacts with hydrogen at 200°C in presence of *nickel catalyst* to form ethane.



➤ **Bromination:**

When ethene is bubbled into bromine water, which is brown, the colour of bromine becomes decolourised.



Uses of ethene

Ethene is used in the manufacture of:

- plastics
- ethanol
- electrical insulators
- bottles and parts of some apparatus.

Test for Alkenes

- (a) Add two drops of a solution of bromine in trichloroethane (CHCl_3) to an alkene.

Observation: The brown colour becomes colourless.

- (b) Add a few drops of acidified potassium manganate (VII) to an alkene.

Observation: The pink colour of the solution becomes colourless.

Alkynes

Alkynes are unsaturated hydrocarbons with the general formula C_nH_{2n-2} .

Where $n = 2, 3, 4, \dots$

The functional group of the alkyne series is a triple bond between two carbon atoms ($C \equiv C$). Alkynes are highly unsaturated compared to alkenes as a result they are more reactive than alkenes.

Like alkenes, they undergo similar chemical reactions.

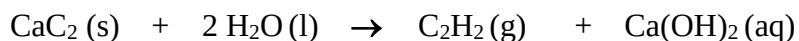
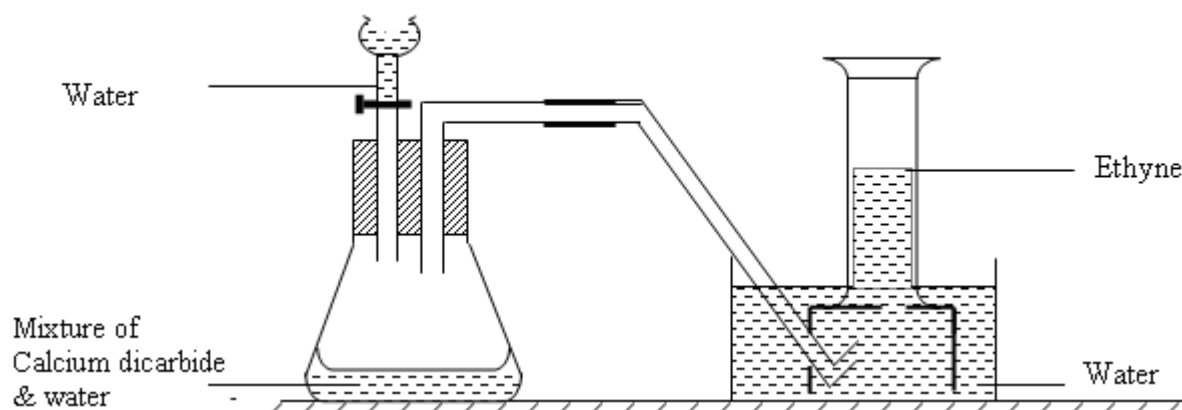
E.g. Combustion and addition reactions.

The most important member of alkynes is ethyne (acetylene), $HC \equiv CH$.

Laboratory Preparation of Ethyne (acetylene)

Ethyne is prepared in the laboratory by dropping water on to calcium dicarbide at room temperature and the gas is collected over water.

Diagram showing laboratory preparation of ethyne



Properties of Ethyne

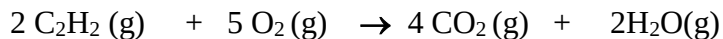
(a) Physical properties

- (i) It is colourless gas.
- (ii) It has a sweet smell.
- (iii) It is insoluble in water.

(b) Chemical Properties

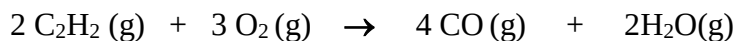
(i) Complete combustion.

Ethyne burns in sufficient air like any other hydrocarbon to form carbon dioxide and water.



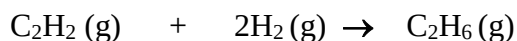
(ii) Incomplete combustion

In limited supply of air, ethyne burns to form carbon monoxide and water.



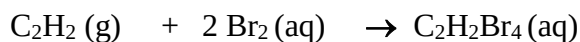
(iii) Addition reaction with hydrogen

Hydrogen combines in presence of nickel to form ethane.



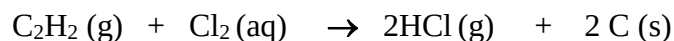
(iv) Addition reaction with bromine

When ethyne is bubbled through bromine water, the reddish-brown colour of bromine water or liquid bromine becomes colourless.

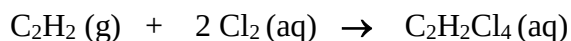


(v) Addition reaction with chlorine

It reacts explosively with chlorine, forming carbon.



NB: *Under special conditions, chlorine and ethyne combine to form tetrachloroethane.*



Uses of Ethyne

Ethyne is used in:

- (i) The industrial manufacture of compounds like adhesive and plastics.
- (ii) Oxy-acetylene flame, which is used in welding and metal cutting.

Petroleum (Crude Oil)

The word “petroleum” means ‘rock oil’. It is found in a layer of porous rock such as sandstone or between two layers of non-porous rock. It is a thick green or dark brown liquid.

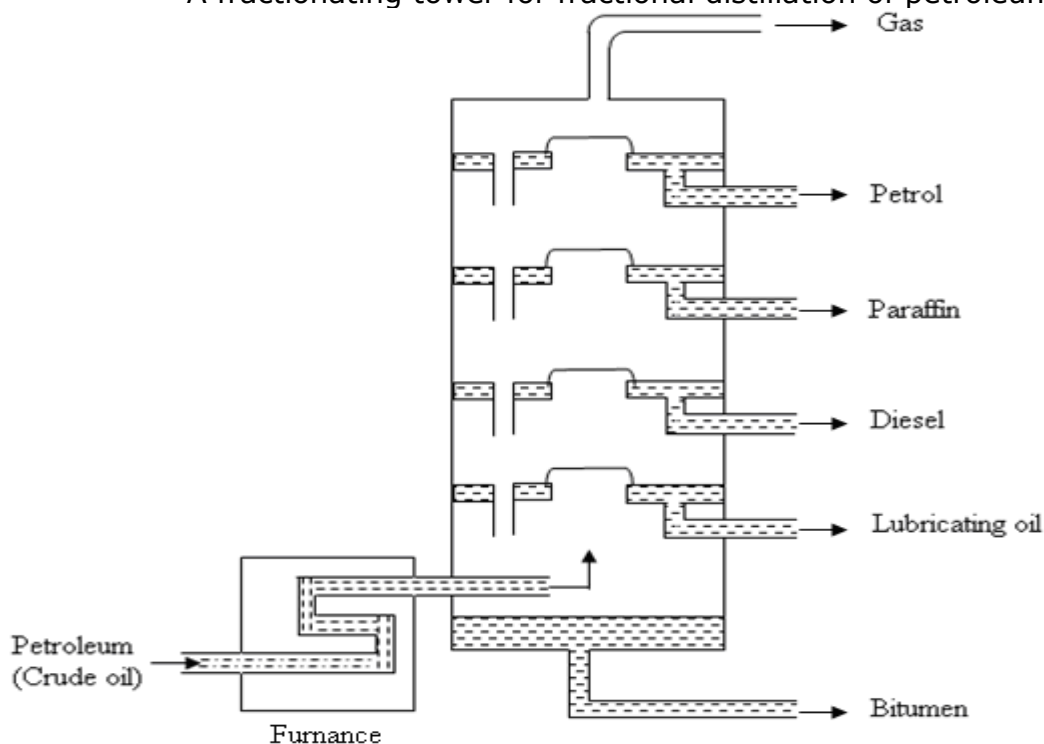
Above the oil is a natural gas. It is a mixture of many gases but contains mainly methane, CH_4 (between 80 – 99%).

Isolation of Alkanes from petroleum

Petroleum contains many different alkanes. The alkanes are separated by a process called *fractional distillation* into a number of different mixtures called *fractions*.

The products (or fractions) are not pure compounds. They are simpler mixtures, each mixture containing several alkanes with different range of boiling points.

A fractionating tower for fractional distillation of petroleum



NB: Petroleum refineries or distilleries are located in isolated areas near seaports rather than far inland. E.g. Mombassa – Kenya, Dar es Salaam – Tanzania.

This is to avoid:

- heavy transport costs of the bulky oil.
- the risks of fire out break.

The table below shows the fractions and their boiling point temperature range

Fraction	Distilling Temperature/°C
Gas	Below 40
Petrol (Naphtha)	40 - 175
Paraffin	175 - 250
Gas oil	250 - 300
Diesel	300 - 350
Lubricating oil	350 - 400
Waxes	Above 400
Bitumen	

NB: Note that:

As the boiling points of the fractions increase, the viscosity and intensity of colour also increases while inflammability decreases.

Uses of the fractions

Fraction	Uses
Gas	Fuel - for heating and lighting in gas cookers
Petrol (Naphtha)	Fuel - in motor aviation using petrol engine.
Paraffin	Fuel - for heating and lighting.
Gas oil	For making petrol, house warming.
Diesel	Fuel - in diesel engines
Lubricating oil	Lubrication of machine parts for smooth running.
Waxes	For making Vaseline, grease, candles.
Bitumen	For tar making roads and runways.

Polymers

A polymer is a complex molecule of high molecular mass formed when several molecules of simple compounds called monomers join up by a process called polymerization.

Polymerization

Polymerization is the joining together of simple molecules called monomers to form a complex molecule called polymer.

Types of polymerization

There are two types of polymerization, namely:

- (i) Addition polymerization
- (ii) Condensation polymerization

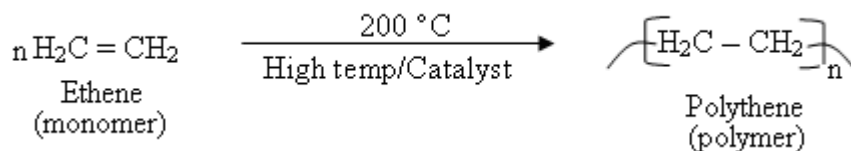
Addition polymerization

Addition polymerization is the formation of a polymer by the combination of monomers, of the same kind, without loss or gain of other atoms.

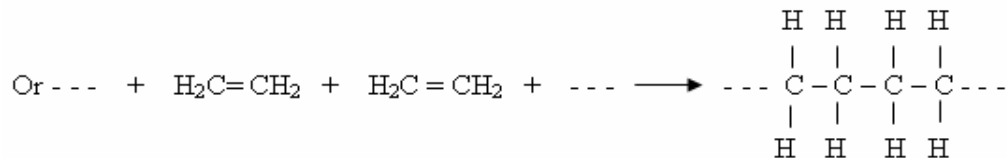
For example, under suitable conditions,

- (i) Ethene molecules (monomers) join to form polythene (polymer).

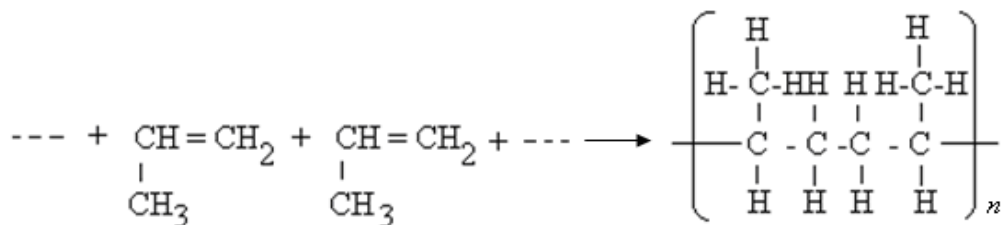
Equation showing addition polymerization



Where: ***n*** is in thousands.



- (ii) Propene molecules join to form polypropene.



Condensation polymerization

Condensation polymerization is the formation of a polymer by the combination of monomers with elimination of water molecules.

- E.g.
- Starch is a polymer formed by condensation polymerization of glucose molecules with the elimination of water molecules.
 - Nylon 6 and nylon 66 are condensation polymers

Types of polymers

There are two types of polymers, namely:

- Natural polymers.
- Synthetic polymers.

(a) Natural polymers:

Natural polymers are polymers that exist in nature. I.e. made by God.

They are formed by condensation polymerization.

Examples of natural polymers are

- Starch, proteins, Fats, Cellulose, Glycogen, Rubber, Wool, Silk, Etc.

(b) Synthetic Polymers

Synthetic Polymers are man made polymers.

Examples of synthetic polymers include:

- Polythene, Polypropene, Perspex, Nylon, Bakelite, Synthetic rubber, Poly Vinyl Chloride (PVC).

All the above polymers are plastic in nature, hence can be molded into plastics.

(b) Synthetic Rubber

It is elastic and not strong enough for use. However, its physical properties can be improved (modified) by heating it with sulphur. The process by which synthetic rubber is heated with sulphur is known as vulcanization of rubber.

Vulcanised rubber is strong, tough and durable.

Cracking

Cracking is the process by which long chain alkanes are broken down to produce shorter chain hydrocarbons.

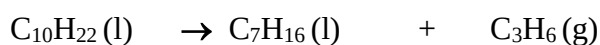
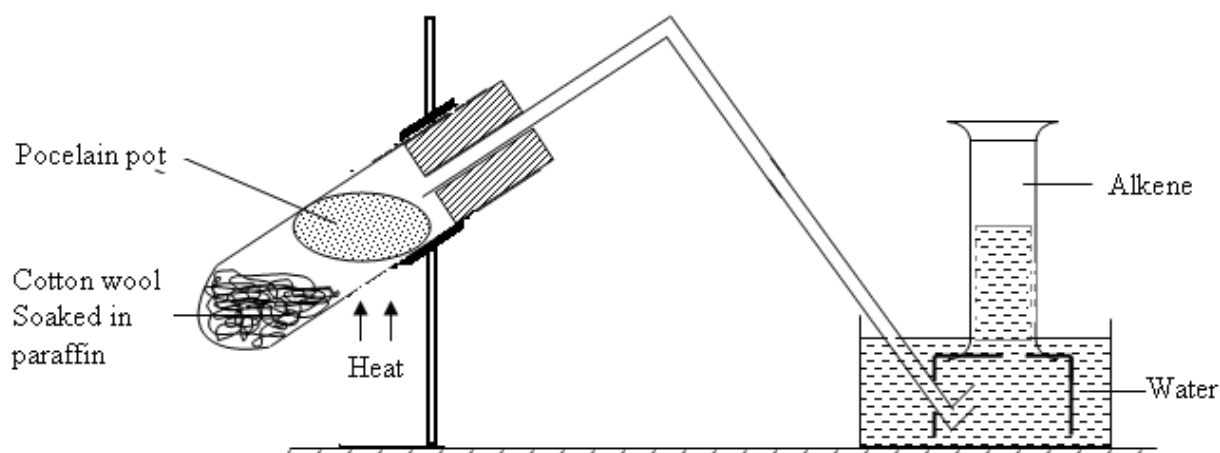
Types of Cracking

There are two types of cracking, namely:

- Thermal cracking
- Catalytic cracking.

(a) Thermal Cracking

This is simply carried out by heating. E.g. cracking of paraffin in the lab.



(b) Catalytic Cracking

Catalytic cracking is carried out using a catalyst.

Plastics

A plastic is a substance which when soft can be made into different shapes.

Classification of Plastics

Plastics are classified into two classes according to their behaviour when heated.

- These are:
- Thermoplastics (Thermo softening plastics)
 - Thermosetting plastics.

(a) Thermoplastics (Thermosoftening plastics)

Thermoplastics are plastics which soften on heating and harden on cooling.

Softening and hardening are reversible. As a result they can be molded or remolded under hot conditions.

E.g. Polythene, polypropene, Perspex, polyvinylchloroethene (PVC). They have low melting points. This is because they are made of molecules with long chains as a result the forces between the chains are weak.

(b) Thermo sets (Thermosetting plastics)

These are plastics, which soften or melt on heating during manufacture and take the shape of the mold in which they are processed on cooling.

They cannot be remolded by heating after manufacture. At high temperatures, they decompose.

E.g. Bakelite (plastic used for making electric switches and some types of plates).

The molecules in thermosetting plastics have cross linkage. As a result the forces between the molecules are strong and their structures are three-dimensional networks.

Uses of Plastics

They are used for making:

- house hold utensils e.g. jags, plates, cups, jericans, combs, shoes etc.

Some uses of specific plastics

Type of Plastic	Use
Polythene	For making bags, plates, cups, electric insulators, rain coats
Poly Vinyl Chloride (PVC)	For making suitcases, radio cases, television sets, computers.
Nylon	For making ropes, clothes, fishing nets.
Perspex	For making wind screens for cars and aircrafts.
Rubber	For making tyres, soles of shoes and tubings
Bakelite	For making electric switches, plugs and sockets.

Advantages and Disadvantages of Plastics

(a) Advantages

- (i) They are light and potable.
- (ii) They are good insulators.
- (iii) They are good thermal radiators.
- (iv) They are resistant to corrosion, hence are used to store acids and alkalis.
- (v) Thermosoftening plastics are easy to manufacture and recycle.
- (vi) They can easily be made coloured and therefore become attractive.

(b) Disadvantages:

- (i) They are poor heat resistant hence melt and burn easily.
- (ii) They are poor weather resistant. As a result they deteriorate on long exposure to high temperature; become brittle and break easily.
- (iii) When burnt, they produce pollutant gases, such as carbon monoxide and carbon dioxide.
- (iv) They are non-biodegradable (i.e. do not rot) hence litter the environment thus making it dirty.

Alcohols

Alcohols are saturated organic compounds which contain the hydroxyl group (OH) as a functional group.

They form a homologous series of general formula $C_nH_{2n+1}OH$ or $C_nH_{2n+2}O$.

Since the hydroxyl (OH) is characteristic of the alcohols, the first formula is usually preferred. The names of alcohols end with - **ol**.

The first two members of this series are:

- | | |
|------------|--------------|
| - methanol | - CH_3OH |
| - ethanol | - C_2H_5OH |

Both exist as liquids at room temperature.

Manufacture of Ethanol by fermentation method

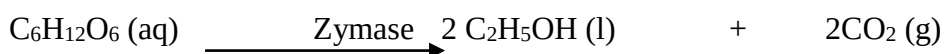
Starch is pressure cooked to release the starch granules and then treated with malt for an hour at 60°C. The starch is hydrolysed by an organic catalyst called diastase to sugar (maltose).



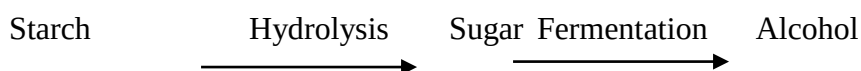
Yeast is added and the mixture is left at room temperature. The maltose is hydrolyzed to glucose by an enzyme called *maltase*.



Another enzyme of yeast, *zymase*, catalyses the decomposition of glucose to ethanol and carbon dioxide.



NB: The main steps are summarized as follows:



Concentrating Ethanol Solution:

The “wash” containing less than 11% of ethanol is fractionally distilled at 78 °C – 82°C to a liquid containing about 95% of ethanol and 5% water. The water can be absorbed by absorption using quicklime (CaO).

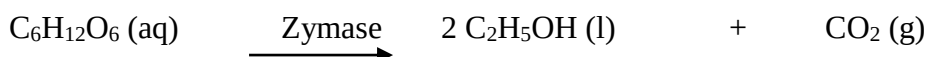
Local Brewing (Production) of Ethanol in Uganda

Maize, cassava and millet flour are mixed with water. The product is soaked in a tin or pot or buried in the ground for about a week to allow starch to be converted into sugar.

The mixture is then removed, roasted and dried. The product is then allowed to ferment for two to three days in suitable container. During this process, the sugar formed is converted to crude ethanol.

The formation of the crude ethanol from the carbohydrates (starch) involves the participation of enzymes.

The relevant equations for the main reactions involved are given below.



Properties of Ethanol

(a) Physical Properties

- (i) It is a colourless liquid.
- (ii) It has a boiling point of 78.5°C.
- (iii) It is highly soluble in water.
- (iv) It has a characteristic smell.

(b) Chemical Properties.

➤ Combustion.

Ethanol burns with a blue flame to form carbon dioxide and water.



➤ Reducing properties of ethanol

- (i) *With acidified potassium dichromate.*

When acidified potassium dichromate, orange in colour, is warmed with ethanol, it turns green.

- (ii) *With acidified potassium manganate (VII).*

When ethanol is added to a solution of acidified potassium permanganate, purple in colour, the purple colour becomes decolourised.

NB: *In each case, ethanol is oxidized to ethanoic acid.*

➤ Dehydration of ethanol

When excess concentrated Sulphuric acid is added to ethanol and the mixture is heated to 180 °C, ethene is formed. This reaction is called *dehydration*.



NB: *At temperatures below 140 °C, ether is predominantly formed.*

Uses of Ethanol

Ethanol is used as:

- (i) beverage.
- (ii) fuel.
- (iii) thermometric liquid in thermometers.
- (iv) A solvent for organic compounds.

Soaps and Detergents

(a) Soaps:

A soap is a sodium salt of a long chain carboxylic acid with general formula

$C_nH_{2n+1}COONa$. Where $n > 8$.

The common soap has $n = 17$, ($C_{17}H_{35}COONa$) and is called sodium stearate in short written as (NaSt).

Soap is manufactured by a process called *saponification*.

Manufacture of Soap (Saponification)

Soap is manufactured by boiling a vegetable oil with concentrated sodium hydroxide solution.

Oil or Fat + Sodium hydroxide $\rightarrow$ Soap + Glycerol

When the process of saponification is complete, brine (concentrated sodium chloride) is added to the mixture to precipitate (to solidify) the soap. Perfumes and colour are also added. The soap, which floats above the liquid, is removed and pressed into bars.

Fats and Oils are esters of long chain carboxylic acid and glycerol.

Oils occur naturally in plants while fats occur naturally in animals only.

Oils are got from plants such as Sunflower, Groundnut, simsim, corn etc.

Examples of fats include: lard in pigs, butter from milk.

(b) Detergents.

A detergent is a sodium salt of sulphuric acid or hydrogen sulphate.

Preparation of detergents

In the laboratory a detergent can be prepared by boiling a vegetable oil with concentrated sulphuric acid. The product, hydrogen sulphate, is then neutralized by adding sodium hydroxide solution. A precipitate forms. The mixture is evaporated on a water bath to leave a white solid detergent.

Advantages and disadvantages of soaps and detergents

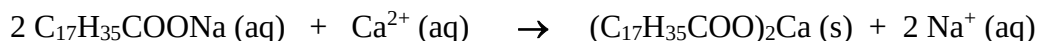
(a) Soaps

Advantages

- (i) Soaps are cheap (inexpensive).
- (ii) They are friendly to environment.
- (iii) They are used for bathing and washing.

Disadvantages

- (i) Soap does not form lather easily with hard water. This leads to waste of it.
- (ii) They react with calcium and magnesium ions in hard water to form scum according to the equation:



If the cloth is not rinsed properly, the scum leaves dirty stains on it.

Detergents Advantages

- (i) They dissolve in water and form lather easily.
- (ii) They do not form scum in hard water.
- (iii) They do not leave stains on clothes.

Disadvantages

- (i) They are expensive.
- (ii) The foam from detergents interferes with the bacterial breakdown of sewage. The sewage not only pollutes the rivers and lakes but also kills the aquatic life. E.g Fish.
- (iii) Detergent contains fertilizer materials which nourishes algae and other microorganisms. This reduces oxygen supply in the water. The result is that fish and other aquatic animals die.

SELF-CHECK 4.1

1. (2005 Q.2).

Which one of the following formulae represents an alkane?

- A. C_2H_4 . B. C_3H_4 . C. C_4H_8 . D. C_4H_{10} .

2. (2005 Q.18).

The reaction between ethanol and concentrated sulphuric acid to form ethene is called

- A. hydrogenation. B. catalysis. C. dehydration. D. hydration.

3. (2004 Q.17).

What name is given to the reaction leading to the formation of soap from oil?

- A. Hydrogenation. B. Polymerisation.
C. Saponification. D. Degradation.

4. (2003 Q.14)

Which one of the following substances decolourises bromine water?

- A. Ammonia. B. Ethene. C. Methane. D. Chlorine.

5. (2003 Q.28).

Which one of the following polymers is a synthetic?

- A. Wool. B. Cotton. C. Sisal. D. Nylon.

6. (2003 Q.29).

Which one of the following is **not** produced during the fermentation of sugar solution?

- A. Water vapour. B. Carbon dioxide. C. Methane. D. heat.

7. (2003 Q.32).

In which one of the following processes does ethene form a plastic?

- A. Polymerisation. B. Precipitation.
C. Neutralisation. D. Electrolysis.

8. (2000 Q.5).

Which one of the following is **not** used as a fossil fuel?

- A. Hydrogen. B. Charcoal.
C. Coal. D. Ethanol.

9. (2000 Q.12).

Concentrated sulphuric acid reacts with ethanol to form ethene. This shows that

- A. the acid is an oxidising agent. B. ethene can be converted to ethanol.
C. ethene has a higher affinity for the acid. D. the acid is dehydrating agent.

10. (2000 Q.37).

An example of a non-biodegradable substance is

- A. silk. B. wool.
C. polythene. D. paper.

11. (1999 Q.7).

The process by which property of rubber is improved by heating it with sulphur is called

- A. polymerisation. B. vulcanisation.
C. catalysis. D. dehydration.

12. (1999 Q.38).

A gas that when bubbled through bromine water changes the colour of bromine water from reddish-brown to colourless is

- A. methane. B. ethene. C. ethane. D. hydrogen.

13. (Q.40).

Which one of the following sets contains natural fibres only?

- A. Nylon, wool, cotton. B. Cotton, nylon, silk.
C. Silk, nylon, wool. D. Cotton, wool, silk.

14. (1998 Q.25).

One of the disadvantages of using detergents for washing is, they

- A. are precipitated in hard water to form scum.
- B. sometimes cause stains on clothes.
- C. cause water pollution in rivers and lakes.
- D. wash away the colour of clothes.

15. (1998 Q.26).

The colourless gas produced during fermentation of sugar is

- A. ammonia.
- B. carbondioxide.
- C. hydrogen.
- D. oxygen.

16. (1998 Q.28).

Which one of the following hydrocarbon contains multiple bonds?

- A. CH_4 .
- B. C_2H_2 .
- C. C_2H_4 .
- D. C_3H_4 .

17. (1998 Q.34).

A colourless gas which decolourises bromine water is

- A. chlorine.
- B. ethene.
- C. sulphur dioxide.
- D. ethane.

18. (1998 Q.40).

Which one of the following can be used to test for ethene?

- A. Lime water.
- B. Bromine water.
- C. Glowing water.
- D. Potassium dichromate paper.

19. (1997 Q.21).

Which one of the following is a synthetic polymer?

- A. Wool.
- B. Cotton.
- C. Silk.
- D. Nylon.

20. (1996 Q.2).

Which one of the following is a monomer of protein?

- A. Glucose.
- B. Amino.
- C. Isoprene.
- D. Ethene.

Each of the questions 21 to 24 consists of an assertion (statement) on the left-hand side and a reason on the right-hand side.

Select:

- A. *if both the assertion and the reason are **true** statements and the reason is a correct explanation of the assertion.*
- B. *if both the assertion and the reason are **true** statements but the reason is **not** a correct explanation of the assertion.*
- C. *if the assertion is **true** but the reason is **not** a correct statement.*
- D. *if the assertion is **not** correct but the reason is a correct statement.*

INSTRUCTIONS SUMMARISED:

Assertion	Reason
A. True	True (reason is a correct explanation.)
B. True	True (Reason is not a correct explanation.)
C. True	Incorrect.
D. Incorrect	True.

21. (1995 Q.43)

Soap can remove both dirt and oil from cloth

because

soap is made from cooking oil.

22. (1995 Q.46)

When ethene burns in air, it produces a smoky flame

because

of the presence of unburnt carbon.

23. (1996 Q.44)

Ethene readily decolourises bromine water

because

it is saturated.

24. (1995 Q.34)

Crude petroleum is refined by fractional crystallisation

because

its fractions have different boiling points.

In each of the questions 25 to 26 one or more of the answers given may be correct. Read each question carefully and then indicate the correct answer according to the following:

A. If 1, 2 and 3 only are correct.

B. If 1 and 3 only are correct.

C. If 2 and 4 only are correct.

D. If 4 only is correct.

25. (1997 Q.32)

How does soap differ from a detergent?

1. A detergent is made from vegetable oil but soap from animal oil.
2. A detergent is not efficient with hard water.
3. A detergent is in powder form only.
4. Soap is not efficient with hard water.

26. (1997 Q.41).

When a mixture of ethanol and concentrated sulphuric acid is heated, a gas is liberated.

Which of the following properties is/are shown by the gas?

1. It is soluble in water.
2. It decolourises bromine water.
3. It decolourises potassium manganate (VII).
4. It forms a white precipitate with limewater.

SELF-CHECK 4.2

1. (2004 Q.9) .

- (a) The molecular formula of ethene is C_2H_4 . Write the structural formula of ethene.
- (b) Bromine water is one of the reagents that can be used to test for the presence of ethene.
 - (i) State what would be observed if ethene is tested with bromine water and write an equation for the reaction.
 - (ii) Name **one** other reagent that can be used to test for the presence of ethene.
- (c) Name **one** compound from which ethene can be prepared.

2. (2003 Q.3)

During the manufacture of soap, sodium hydroxide was boiled with substance X.

- (a) Identify substance X.
- (b) What name is given to the process leading to the formation of soap?
- (c) Name a substance that can be used to precipitate the soap from the solution.
- (d) State what would be observed if soap solution was reacted with aqueous calcium hydrogen carbonate.

3. (2001 Q.13)

- (a) Name the raw materials used in your locality to make an alcoholic drink.
- (b) Briefly describe how ethanol can be obtained from the materials you have named in (a).
- (c) State how ethanol prepared in (b) can be concentrated and suggest one way of determining whether the ethanol is pure or not.
- (d) Ethene can be formed from ethanol.
 - (i) Write equation and state the conditions for the reaction leading to the formation of ethene.
 - (ii) Name two uses of ethanol apart from the preparation of ethene.

4. (1998 Q. 12)

- (a)
 - (i) State the difference between fats and oils.
 - (ii) Give one example of each.
- (b) Briefly describe how soap can be prepared.
- (c) State what would be observed if soap solution was shaken with a solution containing magnesium hydrogen carbonate.
- (d) Explain your answer in (c) above.
- (e) State what would be observed if a soap less detergent was used in (c) instead of soap solution.
- (f) Give one disadvantage of soap less detergents.

5. (1994 Q.7)

- (a) Ethanol can be prepared by fermentation of sugars.

Write an equation to show how ethanol can be prepared from glucose, $C_6H_{12}O_6$.
- (b) When ethanol is heated with concentrated sulphuric acid at $170^\circ C$, a substance W is formed.
 - (i) Name W.
 - (ii) Write the structural formula of W.
 - (iii) Name one reagent that can be used to identify W.

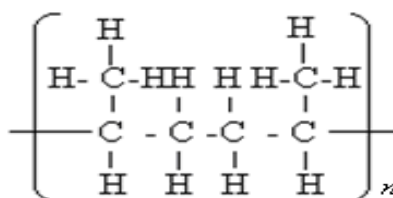
6. (1988 Q.2)

Glucose, $C_6H_{12}O_6$, can be converted to ethanol by a catalytic reaction caused by an enzyme produced from yeast.

- (a) Name; (i) the reaction in which yeast converts glucose into alcohol.
(ii) The enzyme produced by yeast during the reaction.
- (b) Write the equation for that leads to the formation of ethanol.
- (c) Briefly describe how the ethanol produced can be concentrated.

7. (1988 Q.14)

- (a) (i) Explain what is meant by the term polymerization.
(ii) Name two naturally occurring polymers and one synthetic polymer.
- (b) The structure of a polymer is shown below.

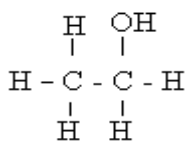


Write down the structural formula of the monomer of the polymer.

- (c) Distinguish between a thermoplastic and a thermosetting plastic.
- (d) Explain the term cracking.
- (e) Draw a fully labeled diagram of the apparatus that can be used to crack liquid paraffin in the laboratory.

8. (1987 Q. 3).

The structure of an organic substance A is shown below.



- (a) Name A.
- (b) A reacts with excess concentrated sulphuric acid at 170°C to form an organic product B.
- (i) Name B.
- (ii) Write the structure of B.
- (iii) Name one reagent that could be used to detect the presence of B.
- (iv) State what would be observed if the reagent named in (iii) was used.

CHAPTER FIVE

NITROGEN AND ITS COMPOUNDS

Nitrogen

Nitrogen, symbol N, with atomic number 7 and mass number 14, is a non-metal. It belongs to group V and period II in the Periodic Table.

Occurrence:

It occurs as uncombined gas in air forming 78.9% by volume of air.

It also exists combined in:

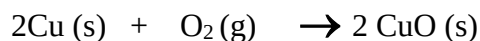
- minerals such as Nitrates and Nitrites,
- oxides e.g. NO, NO₂.

Preparation of Nitrogen from air

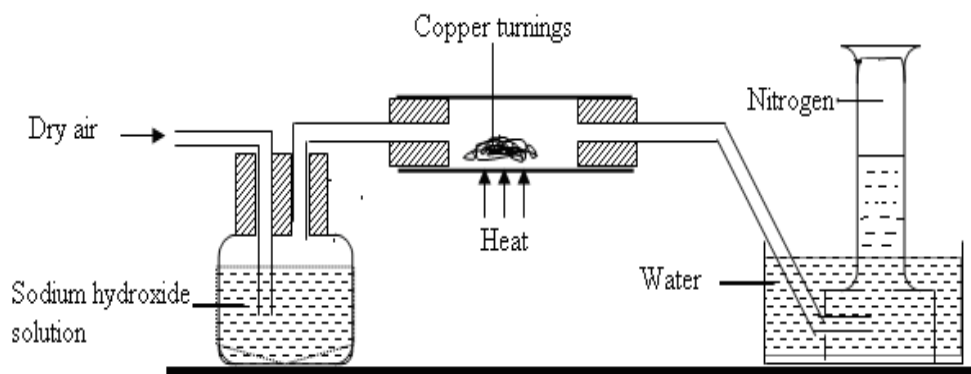
A sample of air is first filtered, dried by passing it through silcagel and then passed through a sodium hydroxide solution to remove carbon dioxide according to the equation:



The remaining air is passed over heated copper turnings in a glass tube to remove oxygen according to the equation:



The nitrogen left is then collected over water as shown in the diagram below.



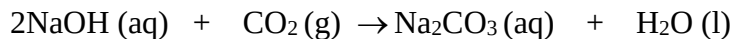
Observation:

The brown copper turnings turn black due to the formation of copper (II) oxide (CuO).

- NB:**
- The nitrogen prepared is not pure. This is because it contains noble gases as impurities.
 - Pure nitrogen gas is manufactured by Industrial distillation of liquid air.

Industrial manufacture of nitrogen

Pure nitrogen is manufactured by fractional distillation of liquid air. A sample of air is first filtered, dried by passing it through silica gel and then passed through a solution of sodium hydroxide to remove carbon dioxide according to the equation:



The resulting dry air is liquefied by successive compressing and cooling.

The liquid is then fractionally distilled. Nitrogen distills off at a boiling point of -196°C and is collected as pure nitrogen gas.

Properties of Nitrogen

(a) Physical Properties

- (i) Nitrogen gas is colourless.
- (ii) It is odorless.
- (iii) It is insoluble in water.
- (iv) It does not support combustion.

(b) Chemical Properties

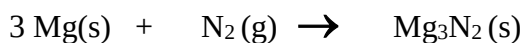
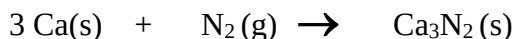
Nitrogen is generally inert. This is because of the triple bond between the two nitrogen ($\text{N} \equiv \text{N}$) atoms in the molecule of nitrogen. To break the triple bond requires a lot of energy. This makes Nitrogen stable. Therefore Nitrogen reacts under special conditions such as:

- Very high temperature,
- Pressure and
- Presence of a catalyst.

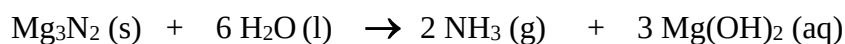
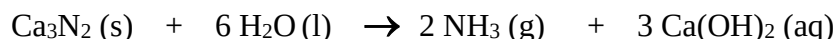
(i) Formation of metallic Nitrides

Nitrogen reacts with some metals at very high temperatures forming nitrides.

E.g. Calcium and Magnesium

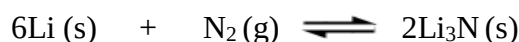


The Magnesium nitride and Calcium nitrides are white solids. They dissolve in water forming hydroxides and ammonia gas according to the equations below.



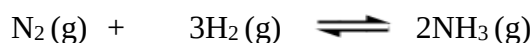
The gas is identified by its characteristic smell and action on red litmus paper.

Lithium also reacts with nitrogen similarly forming lithium nitride.



(ii) Formation of Hydride (Reaction with Hydrogen to form ammonia)

Nitrogen reacts with hydrogen in presence of a catalyst of finely divided iron heated at 450°C and a pressure of about 200 atmospheres to form ammonia according to the equation:

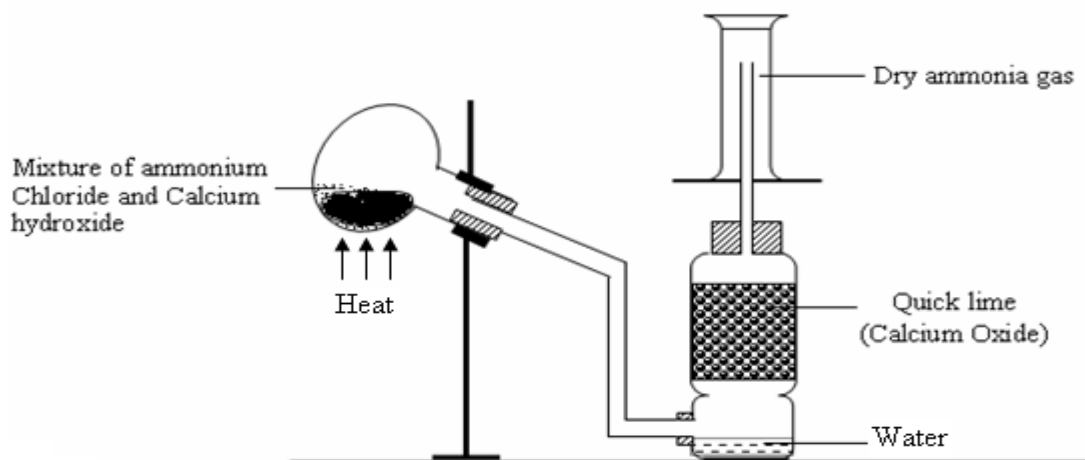


Ammonia Gas (NH₃)

Laboratory Preparation of Ammonia

Ammonia is prepared in the laboratory by heating a mixture of ammonium salt and an alkali. Usually, a mixture of ammonium chloride and calcium hydroxide. The gas is collected by upward delivery method since it is lighter (less dense) than air.

The diagram showing the laboratory preparation of dry Ammonia gas



Properties of Ammonia

(a) Physical Properties

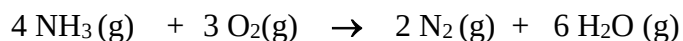
- (i) Ammonia is a colourless gas.
- (ii) It has a characteristics pungent smell.
- (iii) It is lighter than air.
- (iv) It is the most soluble of all gases (1cm³ in water dissolves about 800cm³ of ammonia at room temperature forming ammonia solution).
- (v) It liquefies at -33 °C.
- (vi) It extinguishes a burning splint.
- (vii) It does not burn in air.

(b) Chemical Properties

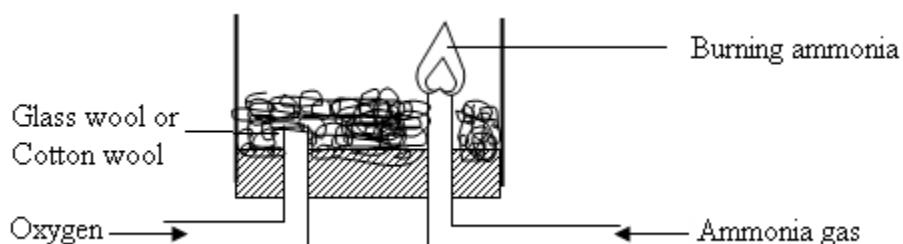
(i) Combustion.

➤ Aided combustion of ammonia.

Ammonia burns in oxygen or air enriched with oxygen giving a yellowish green flame forming nitrogen and water.

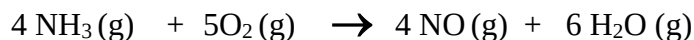


The diagram showing burning of ammonia in enriched air

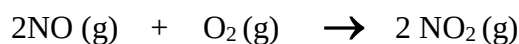


➤ Catalytic oxidation of ammonia

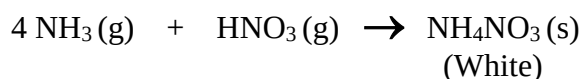
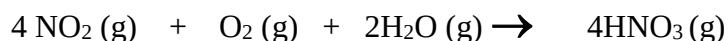
In presence of a catalyst platinum or copper, ammonia burns in oxygen forming nitrogen monoxide and water.



The oxygen oxidizes the nitrogen monoxide to nitrogen dioxide.



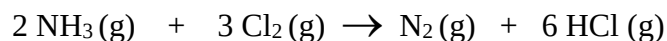
Later the gases turn white due to the formation of ammonium nitrate.



Observation: Reddish-brown fumes and white solid are formed as explained in the above equations.

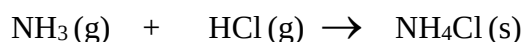
(ii) Reaction with Chlorine.

Ammonia burns in chlorine forming a mist of hydrogen chloride gas.



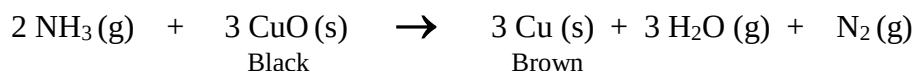
(iii) Reaction with Hydrogen Chloride

Ammonia reacts with hydrogen chloride gas forming dense white fumes of ammonium chloride according to the equation:



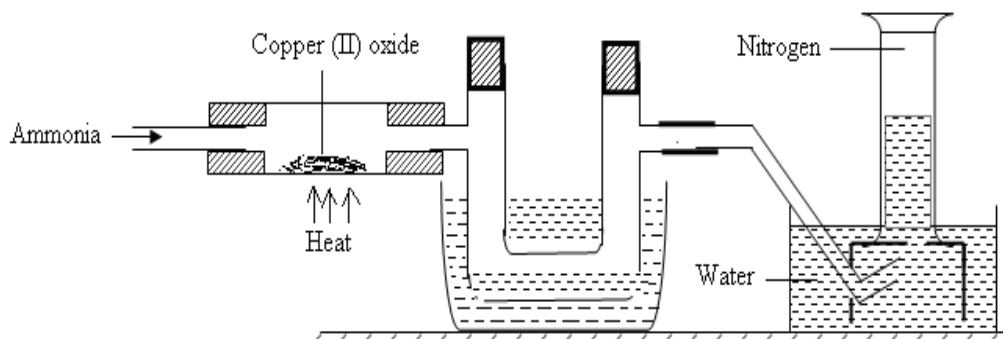
(iv) Reduction of Copper (II) oxide or Lead (II) oxide

Ammonia reduces heated oxide of Copper (II) oxide and Lead (II) oxide to their metals and it self is oxidized to nitrogen and water.

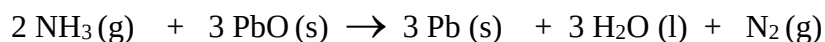


Observation:

- The black solid turned brown and
- A colourless liquid that turns white anhydrous copper (II) sulphate blue.

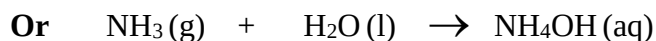
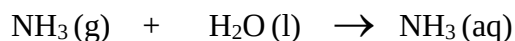


With lead (II) oxide, it reacts according to the equation:



(v) Solubility

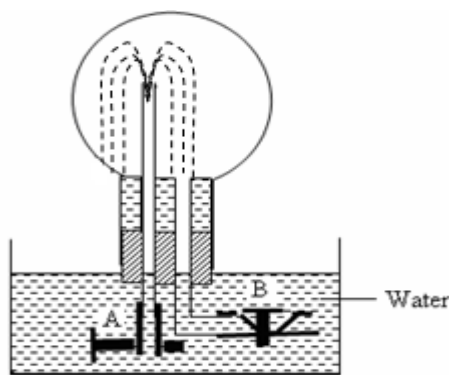
It is the most soluble of all gases. 1cm³ of water dissolves about 800 cm³ of ammonia at room temperature forming ammonia solution.



The high solubility of ammonia is seen in the fountain experiment.

Fountain Experiment

- Fill a large, round-bottomed flask with ammonia and cork it up with a rubber stopper carrying two tubes fitted with clips.
- Invert and clamp firmly the flask carrying the tubes such that its mouth is under water coloured with red litmus in a trough as shown in the diagram below.



- Open clip B for a moment, close it and allow the few drops of water that entered to run down into the round part of the flask.
- Then replace the flask and the tubes under the water and open clip A.

Observation:

- Water rises rapidly up the tube and it enters the flask as a fountain and continues until the flask is full.
- The red litmus solution turns blue.

Explanation:

The few drops of water, which entered when tube B was opened, dissolved so much of ammonia that there was a partial vacuum in the flask.

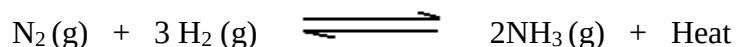
When clip A was opened, atmospheric pressure forced water rapidly up the tube, and the water entering dissolves more ammonia thus creating more vacuum hence a fountain is seen playing until the flask is full of water.

The turning of the red litmus solution to blue shows the alkaline nature of ammonia.

Manufacture of Ammonia (Haber process)

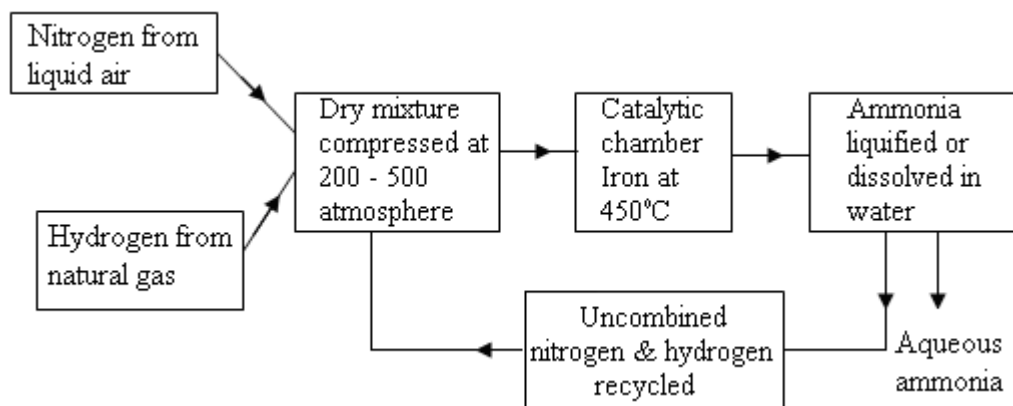
A mixture of dry nitrogen and hydrogen mixed in volume proportions of 1 : 3 respectively is compressed in a compressing chamber at a relatively high pressure usually between 200 and 500 atmospheres and are then passed over finely divided iron impregnated or mixed with alumina, aluminium oxide (a catalyst) at a temperature of 450°C in a catalytic chamber. (The aluminium oxide improves the performance of iron by making it porous)

The reaction is reversible and exothermic according to the equation:



The ammonia produced or formed is either liquefied by cooling it in a cooling chamber or dissolved in water and tapped off. The uncombined nitrogen and hydrogen are passed back to the compressing chamber for further recycling.

Flow chart for the manufacture of ammonia by Haber process



Test of Ammonia

Ammonia turns moist red litmus paper blue.

NB: *Ammonia is the only alkaline gas that turns red litmus blue.*

Uses of Ammonia

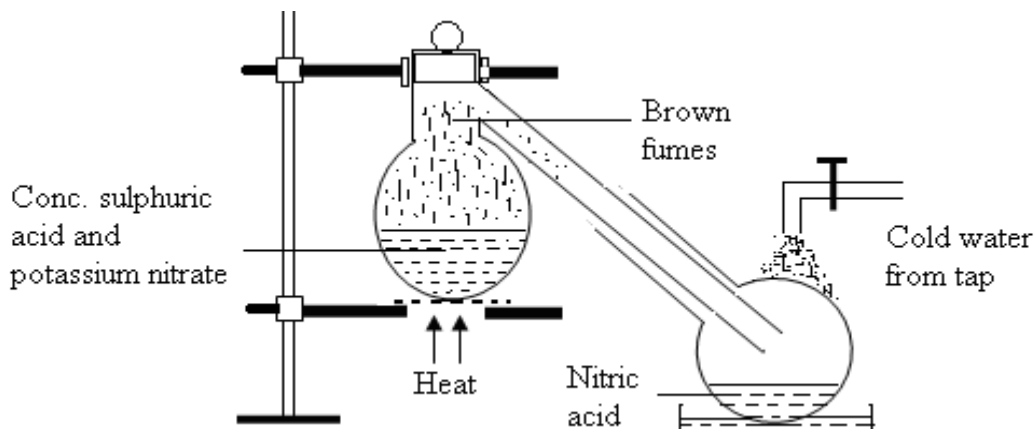
- (i) Manufacture of fertilizers.
- (ii) Making of explosives.
- (iii) In laundry, ammonia is used to soften water.
- (iv) Ammonia is used to prevent fainting and dizziness.
- (v) Liquid ammonia is a raw material for the manufacture of hydrogen.
- (vi) Manufacture of nitric acid.

Nitric Acid

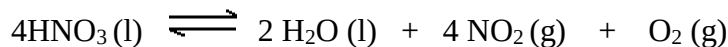
Laboratory Preparation of Nitric Acid

Nitric acid can be prepared in the laboratory by heating a mixture of metallic nitrate such as potassium nitrate and concentrated sulphuric acid as shown in the diagram.

The diagram showing the laboratory preparation of nitric acid



The nitric acid distils and collects in the condenser as a yellow liquid. The yellow colouration is due to brown fumes of nitrogen dioxide formed by slight decomposition of the nitric acid by heat and dissolving in it.



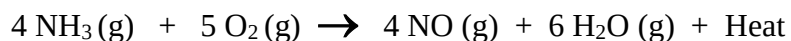
The yellow colour may be removed by bubbling air or oxygen into the acid.

Manufacture of Nitric Acid

There are three main stages in the manufacture of nitric acid. These include:

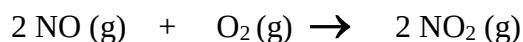
1. Oxidation of Ammonia

Ammonia from Haber process is mixed with excess air and the mixture is passed over platinum catalyst heated at about 800°C and 8 atmospheres pressure in the catalytic chamber. It is oxidized to nitrogen monoxide. The reaction is exothermic and once, started, maintains the temperature of the catalyst.



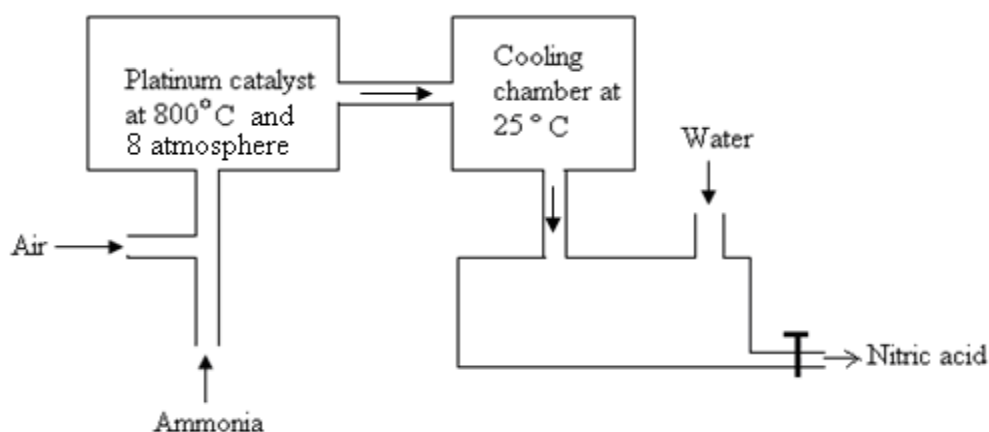
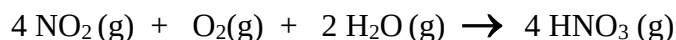
2. Oxidation of nitrogen monoxide to nitrogen dioxide.

The nitrogen monoxide formed is rapidly cooled and combines with the oxygen from excess air to form nitrogen dioxide.



3. Oxidation of nitrogen dioxide and absorption.

The nitrogen dioxide is oxidized by more air and then absorbed in hot water to form nitric acid.



The pure ordinary concentrated nitric acid as sold contains about 70% by mass of the pure acid and 30% of water.

Properties of Nitric Acid

(a) Physical properties

- (i) Nitric acid is a colourless, fuming liquid.
- (ii) It has a density of 1.5 g cm^{-3}
- (iii) It has a boiling point of 85°C at ordinary atmospheric pressure.
- (iv) The acid is corrosive and destroys organic matter very readily.

(b) Chemical properties

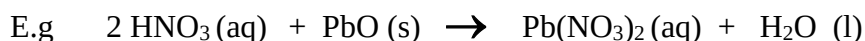
The chemical properties of nitric acid are divided under two main headings.

- Where nitric acid behaves as an acid (when dilute).
- Where nitric acid behaves as an oxidizing agent (when concentrated).
-

1. Nitric acid as an acid.

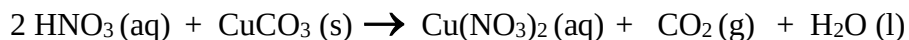
(i) Reaction with bases

It reacts with bases to form salt and water.



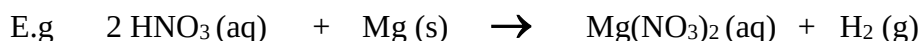
(ii) Reaction with carbonates

It reacts with carbonates to form carbon dioxide, metallic nitrate and water according to the equation:.



(iii) Reaction with metals

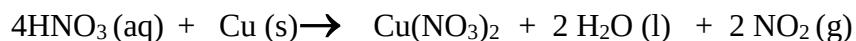
Unlike other dilute acids, dilute nitric acid does not release hydrogen when heated with metals, except with magnesium.



2. Nitric acid as an oxidizing agent (When concentrated)

(i) Reaction with Copper

When copper is dropped into concentrated nitric acid, it dissolves with effervescence of reddish brown fumes of nitrogen dioxide and a green solution containing copper (II) nitrate is formed.



On dilution with water, the solution turns from green to blue.

(ii) Reaction with sulphur

When sulphur is dropped into concentrated nitric acid and the mixture is warmed gently in presence of bromine as a catalyst, effervescence of reddish brown fumes of nitrogen dioxide are evolved according to the equation:



NB: *Whenever nitric acid is reduced, brown fumes of nitrogen dioxide are evolved.*

Uses of Nitric Acid

Nitric acid is used in the Manufacture of:

- (i) fertilizers
- (ii) explosives
- (iii) dyes
- (iv) drugs

Nitrates

(a) Nitrates

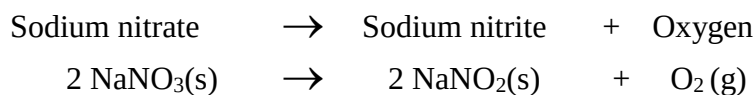
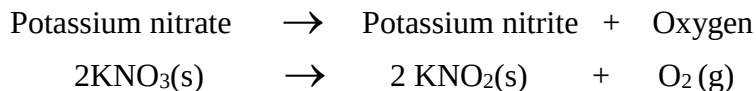
Nitrates are salts of nitric acid.

The properties of nitrates vary according to the position of the metal in the reactivity series. The lower the metal is in the series, the more readily and completely the nitrate of that metal decomposes.

(i) Metallic nitrates that decompose to form two products

Potassium nitrate (KNO₃) and Sodium nitrate (NaNO₃)

Potassium nitrate and Sodium nitrate melt to form a colourless liquid. On further heating, they slowly decompose to form a pale-yellow *metallic nitrite* and *oxygen* as shown in the following equations:



Ammonium nitrate (NH₄NO₃)

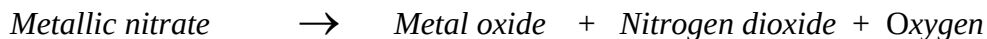
Ammonium nitrate, NH₄NO₃, decomposes on heating to form di-nitrogen oxide (N₂O) and water.



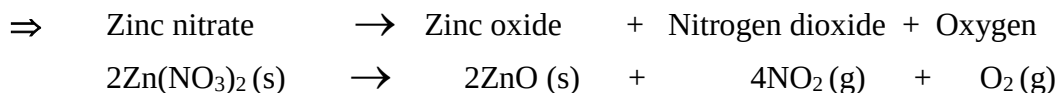
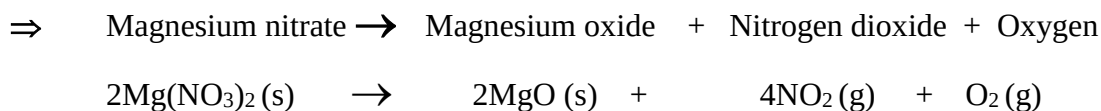
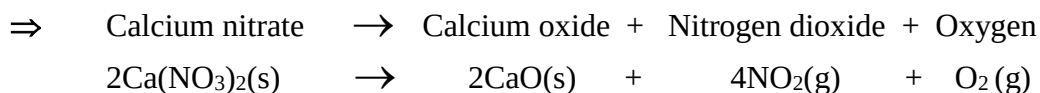
NB: *Samples of ammonium nitrate can be explosive and it is best done in Situ by using a mixture of ammonium chloride and potassium nitrate.*

(ii) Metallic nitrates that decompose to form three products

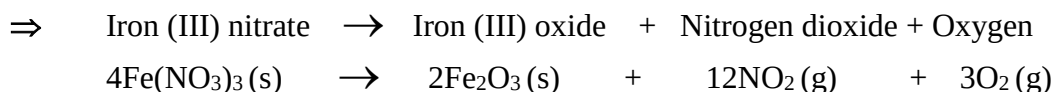
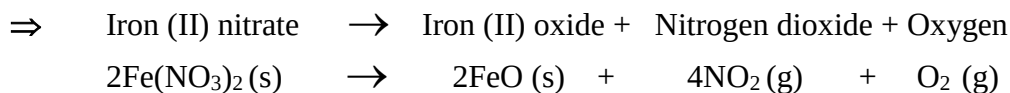
The metallic nitrates from *calcium* to *copper* decompose to form the *oxide of the metal*, brown fumes of *nitrogen dioxide* and *oxygen*.



Example:

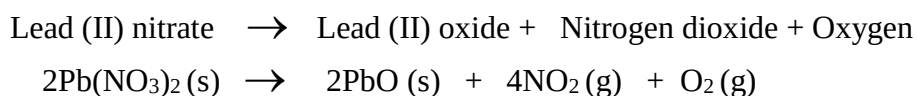


NB: The zinc oxide is yellow when hot and white when cold.

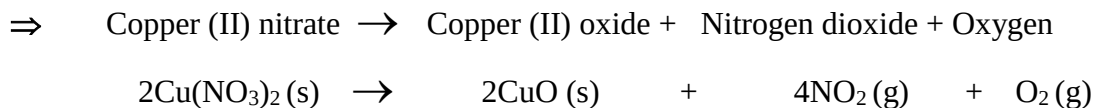


$\Rightarrow$ Lead (II) nitrate

When lead (II) nitrate is heated, it decrepitates. That is it decomposes with a **crackling** sound giving off brown fumes of nitrogen dioxide and leaving a reddish brown solid when hot and yellow when cold.



NB: The crackling sound is due to formation of gas molecules that form inside the crystals and splits them.

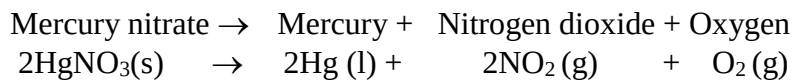
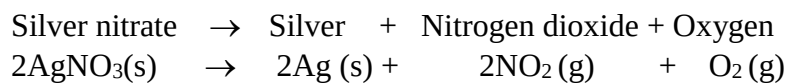


(iii) Metallic nitrates that decompose to form three (3) products, metal, nitrogen dioxide and oxygen are:

- Silver nitrate and
- Mercury

Why do these nitrates form metals when heated?

The nitrates of these metals decompose to form metal oxides, nitrogen dioxide and oxygen. But the oxides are unstable to heat and therefore they decompose to the metal and oxygen.



The summary of the effect of heat on nitrates

Metal	Effect of heat	Solubility
K Na	} Nitrates of these metals are decomposed by heat to give two products : - The <i>nitrite</i> and <i>oxygen</i> .	All Nitrates are soluble in water
Ca Mg Al Zn Fe Pb Cu	} Nitrates of these metals are decomposed on heating to give three products : - The <i>oxide of the metal</i> , - <i>Nitrogen dioxide</i> and - <i>Oxygen</i>	
Hg Ag	} Nitrates of these metals are decomposed on heating to give three products : - <i>Metal, nitrogen dioxide</i> and <i>oxygen</i> .	

NITROGEN AND ITS COMPOUNDS)

SELF-CHECK 5.1

1. (2005Q.14)

Which one of the following nitrates will produce nitrogen dioxide when strongly heated?

- A. Potassium nitrate.
- B. Sodium nitrate.
- C. Zinc nitrate.
- D. Ammonium nitrate.

2. (2005 Q.17)

The substance which is **most** suitable for drying ammonia is

- A. concentrated sulphuric acid.
- B. calcium chloride.
- C. phosphorus (V) oxide.
- D. calcium oxide.

3. (2005 Q.25)

The gas that turns brown when exposed in air from the following list is

- A. sulphur dioxide.
- B. hydrogen chloride.
- C. hydrogen sulphide.
- D. nitrogen monoxide.

4. (2005 Q.31)

The catalyst used during the manufacture of nitric acid by oxidation of ammonia is

- A. platinum.
- B. iron.
- C. nickel.
- D. copper.

5. (2004Q.10)

Under a certain temperature and pressure, hydrogen reacted with nitrogen according to the equation below: $3\text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$

The volume of nitrogen required to react with 150 cm^3 of hydrogen under the same temperature and pressure is

- A. 15.0 cm^3
- B. 50.0 cm^3
- C. 300.0 cm^3
- D. 450.0 cm^3

6. (2004 Q.24)

Which one of the following is **not** a property of ammonia? It is

- A. an alkaline gas.
- B. a reducing agent.
- C. soluble in water.
- D. is denser than air.

7. (2004 Q.28)

Which one of the following oxides can be reduced by dry ammonia?

- A. Copper (II) oxide.
- B. Magnesium oxide.
- C. Zinc oxide.
- D. Calcium oxide.

8. (2004 Q.34)

Ammonia burns in oxygen to yield

- A. nitrogen and water.
- B. nitric acid.
- C. nitrogen and hydrogen.
- D. nitric acid, nitrogen and water

9. (2004 Q.36)

Fuming nitric acid was heated and the gas evolved was collected over water.

This gas was

A. nitrogen dioxide.

B. oxygen.

C. nitrogen monoxide.

D. hydrogen.

10. (2004Q.39)

The process which does **not** require a catalyst is the manufacture of

A. nitric acid.

B. ammonia.

C. sodium hydroxide.

D. Sulphuric acid.

11. (2003Q.6)

Which one of the following metals reacts with cold dilute nitric acid?

A. Calcium.

B. Copper.

C. Silver.

D. Lead.

12. (2003 Q.10)

Which one of the following substances is used as a catalyst in the manufacture of nitric acid?

A. Vanadian (V) oxide.

B. Manganese (IV) oxide.

C. Platinised asbestos.

D. Finely divided iron.

13. (2003 Q.35)

Which one of the following hydroxides is normally used in the laboratory to prepare ammonia from ammonium chloride?

A. Sodium hydroxide.

B. Calcium hydroxide.

C. Copper (II) hydroxide.

D. Iron (II) hydroxide.

14. (2003 Q.38)

Which one of the following gases forms white fumes with ammonia?

A. Chlorine.

B. Carbon dioxide.

C. Sulphur dioxide.

D. Hydrogen chloride.

15. (2002Q.39)

When ammonium nitrate is heated, it produces

A. nitrogen dioxide.

B. ammonia.

C. dinitrogen oxide.

D. nitrogen monoxide

16. (2001Q.19)

Dilute nitric acid reacts with copper to produce

A. copper nitrate, water and nitrogen dioxide. C. copper nitrate, water and ammonia.

B. copper nitrate, water and nitrogen monoxide. D. copper nitrate, water and hydrogen.

17. (2001 Q.24)

The product formed when silver nitrate is heated until there is no further change is

A. silver oxide, nitrogen dioxide and oxygen. C. silver metal and nitrogen dioxide

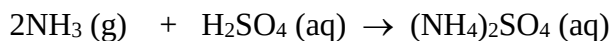
B. silver metal, nitrogen dioxide and oxygen. D. silver oxide and nitrogen dioxide.

18. (2000 Q.25). Which one of the following metals reacts with nitrogen when heated?
- | | |
|-------------|------------|
| A. Calcium. | B. Silver. |
| C. Copper. | D. Sodium. |
19. (1999 Q.9).
The catalyst used in the manufacture of nitric acid is
- | | |
|----------------------|-----------------------|
| A. iron. | B. platinum. |
| C. iron (III) oxide. | D. vanadium (V) oxide |
20. (1999 Q.23).
Which one of the following products are formed when concentrated nitric acid reacts with copper?
- A. Copper oxide, water and nitrogen dioxide.
B. Copper (II) nitrate, water and nitrogen monoxide.
C. Copper oxide, water and nitrogen monoxide.
D. Copper (II) nitrate, water and nitrogen dioxide.
21. (1999 Q.37).
Which one of the following nitrates does **not** decompose on heating?
- | | |
|--------------------|--------------------|
| A. Copper nitrate. | B. Lead nitrate. |
| C. Sodium nitrate. | D. Silver nitrate. |
22. (1998 Q.3).
Which one of the following nitrates does **not** produce nitrogen oxide when heated strongly?
- | | |
|-------------------------------|---------------------------------|
| A. KNO_3 . | B. $\text{Ca}(\text{NO}_3)_2$. |
| C. NH_4NO_3 . | D. $\text{Cu}(\text{NO}_3)_2$. |
23. (1998 Q.7).
Which one of the following gases is dried using calcium oxide?
- | | |
|---------------------|-----------------------|
| A. Sulphur dioxide. | B. Hydrogen chloride. |
| C. Ammonia. | D. Hydrogen. |
24. (1997Q.24).
The fountain experiment can be demonstrated with ammonia because ammonia
- | | |
|------------------------------------|-------------------------|
| A. reacts readily with water. | C. is lighter than air. |
| B. is a very soluble gas in water. | D. is denser than air. |
25. (1997 Q.45).
Which of the following does **not** occur when sodium nitrate is strongly heated?
- | | |
|----------------------------------|-------------------------|
| A. It melts. | C. It liberates oxygen. |
| B. It liberates nitrogen dioxide | D. It loses weight. |

SELF-CHECK 5.2

NITROGEN

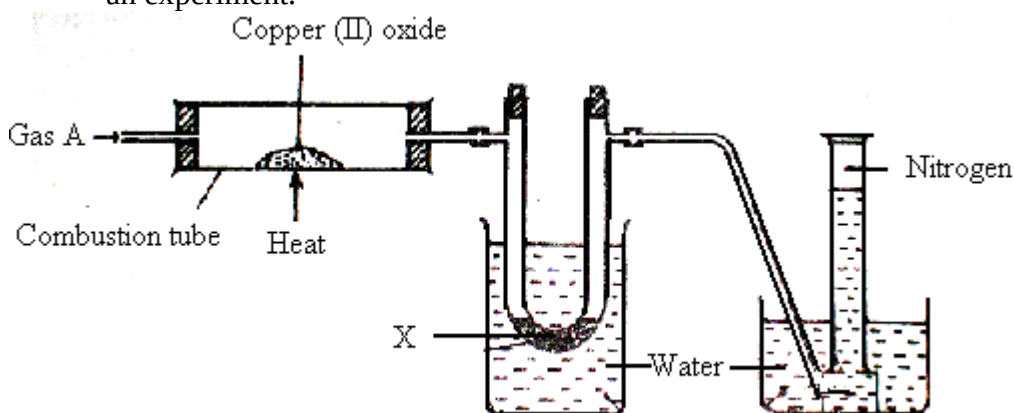
1. In an experiment, ammonia was prepared by heating an ammonium salt with an alkali. After drying, 240 cm³ of ammonia gas were collected at room temperature and pressure. All the ammonia was then reacted completely with 250 cm³ solution of Hydrochloric acid.
- (a) What is meant by the term *alkali*?
 - (b) Explain, using the physical properties of the gas, why ammonia is not collected:
 - (i) Over water
 - (ii) By downward delivery method.
 - (c) Ammonia turns wet red litmus paper blue. Which ion in ammonia solution is responsible for this reaction?
 - (d) Calculate the number of moles of ammonia gas that were collected in the above experiment given that one mole of gas occupies a volume of 24,000 cm³ at room temperature.
 - (e) The equation below shows the reaction between ammonia and sulphuric acid.



- (i) Explain how crystals of ammonium sulphate could be obtained in this experiment.
- (ii) Calculate the maximum mass of ammonium chloride that could be obtained in this experiment. (N = 14, H = 1, Cl = 35.5, S = 32, O = 16)

SELF-CHECK 5.3

2. (a) The diagram below shows a set-up that can be used to obtain nitrogen gas in an experiment.



- (i) Name gas A and liquid X.
- (ii) What observation would be made in combustion tube after heating for some time?
- (iii) Write an equation for the reaction that took place in the combustion.

(iv) If 320 cm³ of ammonia gas reacted completely with the copper?

Calculate:

(I) the volume of nitrogen gas produced.

(II) the mass of copper(II) oxide that reacted. (Cu = 63.5, O = 16.0, 1 mole of gas occupies 24 dm³ at rt and pressure).

(III) At the end of the experiment the pH of the water in the beaker was found to be about 10. Explain.

(b) In another experiment a gas jar containing ammonia was inverted over a burning splint. State what observation was made?

(c) Why is it advisable to obtain nitrogen from air instead ammonia?

3. (2005 Q.11).

(a) Describe how a dry sample of ammonia can be prepared in the laboratory. (Diagram **not** required.)

(b) Name a reagent that can be used to test for ammonia and state what would be observed if ammonia is tested with the reagent.

(c) (i) Draw a labelled diagram of the set up of the apparatus that can be used to show that ammonia can burn in oxygen.

(ii) Write an equation for the combustion of ammonia in oxygen.

(d) Dry ammonia was passed over heated copper (II) oxide.

(i) State what was observed.

(ii) Write an equation for the reaction.

4. (2002 Q.13).

(a) (i) Draw a labelled diagram to show how a dry sample of ammonia can be prepared from ammonium chloride in the laboratory.

(ii) Write equation for the reaction leading to the formation of ammonia.

(b) Dry ammonia gas was passed over-heated lead (II) oxide.

(i) State what was observed.

(ii) Write equation for the reaction that takes place.

(c) Describe how ammonia can be converted to nitric acid. Use equation to illustrate your answer.

5. (2000 Q.11).

- (a)
 - (i) Name the raw materials used for the manufacture of ammonia.
 - (ii) Write equation for the reaction leading to the formation of ammonia.
- (b) Explain how formation of ammonia is affected by:
 - (i) Pressure.
 - (ii) Temperature.
- (c) State another factor that affects the formation of ammonia.
- (d) Dry ammonia was passed over heated copper (II) oxide until there was no further change. State what was observed and explain your answer.

6. (1993 Q.12).

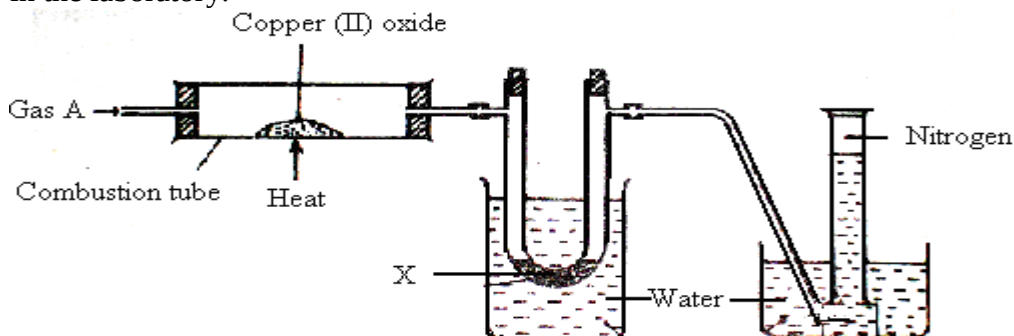
- (a) Draw a labelled diagram of the apparatus that can be used to prepare ammonia in the laboratory.
- (b) Describe an experiment that can be used to show that ammonia is a soluble alkaline gas.
- (c) A copper coil was heated strongly and held over a concentrated solution of ammonia in a beaker. Oxygen was then bubbled into the ammonia solution.
 - (i) State what was observed.
 - (ii) Explain the observations in (i).
- (d) Ammonia reacts with lead (II) oxide according to the equation.

$$2\text{NH}_3(\text{g}) + 3\text{PbO}(\text{s}) \rightarrow \text{N}_2(\text{g}) + 3\text{Pb}(\text{s}) + 3\text{H}_2\text{O}(\text{l}).$$
 Calculate the volume of ammonia at room temperature that would be required to completely react with 2.5 g of lead (II) oxide.

(N = 14, H = 1, Pb = 207, O = 16)

7. (1995 Q.4).

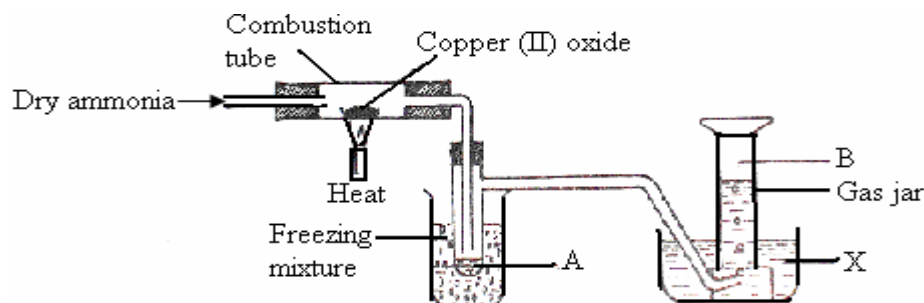
The diagram in the figure shows the apparatus which can be used to prepare nitrogen in the laboratory.



- (a) Name gas A.
- (b)
 - (i) State what would be observed in the combustion tube during the reaction.
 - (ii) Write an equation for the reaction.
- (c) Name
 - (i) substance X.
 - (ii) One reagent that can be used to identify X.

8. (1991 Q.7).

Substances A and B were obtained from a reaction between ammonia gas and copper oxide using the apparatus shown in the diagram below.



- (a) Name substance
 - (i) A
 - (ii) B
 - (iii) X
- (b) Write an equation for the reaction that takes place in the combustion tube.
- (c) State why it is not possible to collect excess ammonia in the gas jar.
- (d) Name one other oxide that can be used instead of copper (II) oxide.

9. (1992 Q.14).

- (a) Describe the industrial preparation of nitric acid from ammonia (Diagram of the plant is not required). Your description should include equations for the reactions that occur.
- (b) Explain what happens when concentrated nitric acid is added to copper.
- (c) Describe one chemical test that can be used to confirm the presence of nitrate.

EXTRA STUDY QUESTIONS

1. (2005 Q.1).

Although nitrogen is generally un-reactive, it readily reacts with a burning magnesium ribbon.

- (a) Give a reason why nitrogen is generally inert.
- (b) Burning magnesium reacts with nitrogen.
 - (i) Give a reason for the reaction.
 - (ii) Write an equation for the reaction.
- (c) Water was added to the product in (b).
Write an equation for the reaction.

2. (1994 Q.8). (a)

- (a) A piece of burning magnesium was introduced into a jar of nitrogen.
 - (i) State what was observed.
 - (ii) Write an equation for the reaction that took place.
- (b) Water was added to product of the reaction in (a) and the resultant mixture tested with litmus. State what was observed.
- (c) Name one other metal that reacts with nitrogen in a similar way to magnesium.

CHAPTER SIX

SULPHUR AND ITS COMPOUNDS

Sulphur, symbol S, is a yellow non-metallic solid. It belongs to group VI and period III in the Periodic Table.

Occurrence:

Sulphur occurs as a free element underground. It also occurs in combined form as sulphides.

E.g. Iron sulphide - FeS
 Copper pyrite - CuFeS_2
 Metallic sulphates and metallic sulphites.
 Petroleum oil and natural gas often contain sulphur.

Extraction of Sulphur

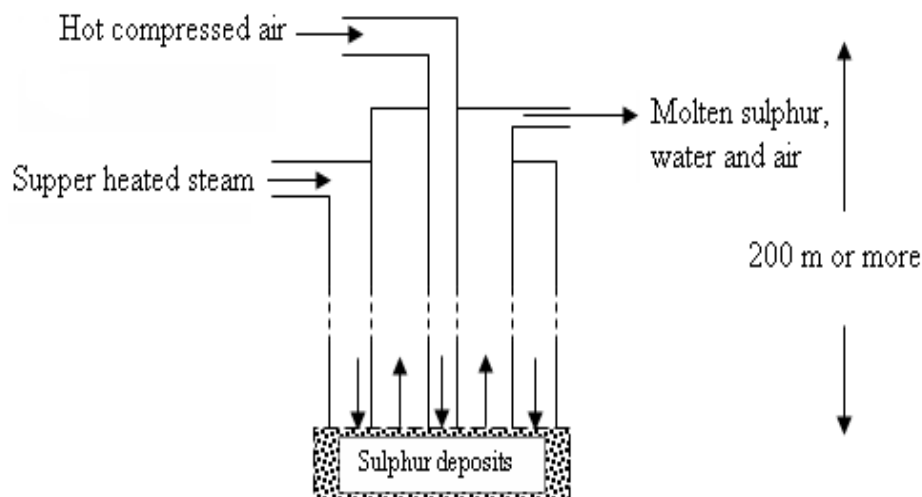
Several methods are used to extract free sulphur from the earth. But the most common method used is the Frasch method, invented in 1891 by the American chemist Herman Frasch.

The Frasch process for Extraction of Sulphur

In this method, a device called *sulphur pump* which consists of three concentric pipes (the largest being about 20 cm (8 in) in diameter) is sunk down to the sulphur deposit. Through the outermost pipe, steam of water that is super heated at 170°C and a pressure of 10-15 atmospheres is directed underground into the sulphur deposit to melt the sulphur.

When a sufficient quantity of sulphur has been melted, a blast of hot compressed air at about 15 atmospheres is forced down the inmost pipe i.e the narrowest pipe and the mixture of the molten sulphur, water and air is forced up to the surface through the middle pipe. The sulphur is run into wooden bins where it solidifies on cooling as roll sulphur, a product that is, about 99.5 percent pure.

A diagram showing Frasch process for Extraction of Sulphur



Allotropes of Sulphur

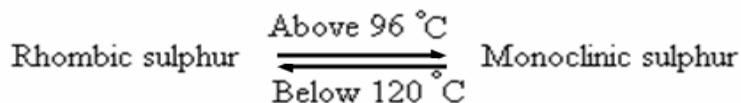
Sulphur as an allotropic element exists in variety of allotropic crystalline forms at room temperature and also amorphous (non-crystalline). The most common allotropes are:

- Rhombic sulphur and Monoclinic sulphur.

Rhombic Sulphur

Rhombic sulphur is the most stable allotrope of sulphur. It has the following properties:

- It is a yellow, crystalline solid.
- It has a density of 2.06 g/cm^3 at 20°C .
- It melts at 119°C and
- Boils at 444°C .
- It is insoluble in water and slightly soluble in alcohol and ether, moderately soluble in oils and extremely soluble in carbon disulphide.
- When kept at temperatures above 96°C , but below 120°C , the rhombic form changes to monoclinic sulphur.



The temperature at which rhombic and monoclinic sulphur are in equilibrium, 96°C , is known as the transition temperature.

Monoclinic sulphur

- It consists of elongated, transparent, needlelike structures
- It has a density of 1.96 g/cm^3 at 20°C .

The difference between rhombic sulphur and monoclinic sulphur

Rhombic Sulphur	Mono-clinic Sulphur
Stable below 96°C	Stable above 96°C
Octahedral crystal	Needle-shaped crystal
Bright yellow	Pale yellow
Melting point 114°C	Melting point 119°C
Density higher (2.08 g/cm ³) at 20°C	Density lower (1.98 g/cm ³) at 20°C

Two facts prove that the rhombic and monoclinic sulphur are allotropes of sulphur.

1. One gram of monoclinic sulphur slowly changes at room temperature into one gram of rhombic sulphur.
2. One gram of either form burns in oxygen to yield the same mass (2g) of sulphur dioxide and nothing else.



Uses of Sulphur

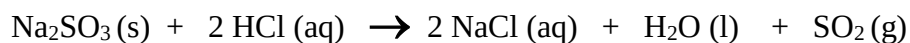
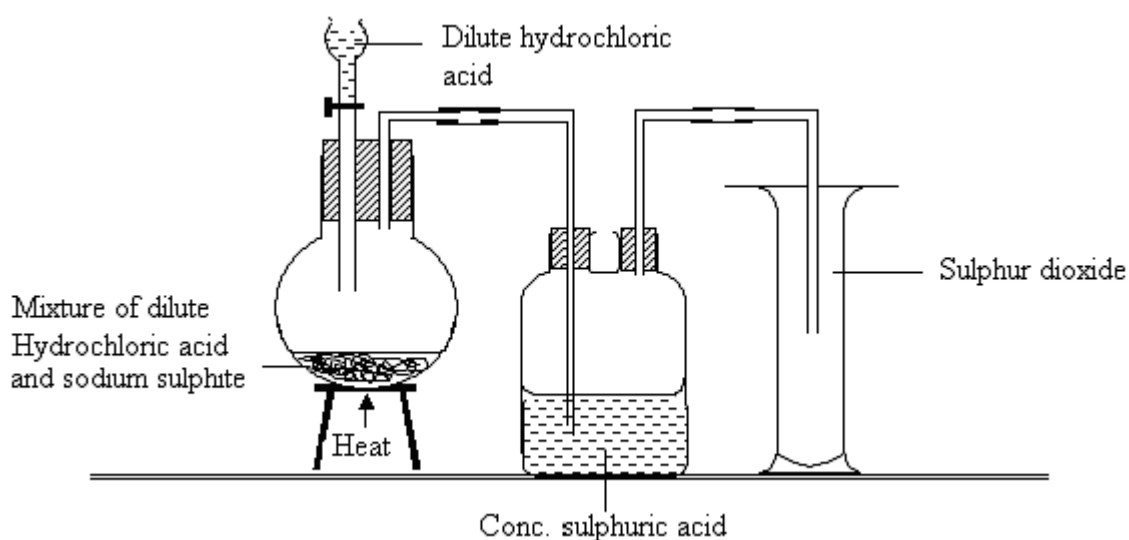
- (i) Medicinally, it is used to manufacture drugs and skin ointments.
- (ii) Industrially, Sulphur is employed in the production of:
 - Matches,
 - Vulcanized rubber,
 - Dyes, and
 - Explosives e.g Gunpowder
- (iii) In agriculture, finely divided sulphur, mixed with lime, is used as a fungicide on plants.
- (iv) In Photography, sulphur is used to make a salt called *hypo* (sodium thiosulphate, Na₂ S₂O₃·5H₂O) which is used for “fixing” negatives and prints.
- (v) In education, the most important use of sulphur is in the manufacture of sulphur compounds, such as:
 - Salts e.g - Metallic Sulphites, Sulphates and sulphides,
 - Acids e.g - Sulphurous acid and Sulphuric acid
 - Gases e.g - Sulphur dioxide and Sulphur trioxide.

Sulphur Dioxide

Laboratory preparation of Sulphur dioxide

Sulphur dioxide is prepared in the laboratory by heating a mixture of dilute hydrochloric acid and sodium sulphite in a round bottom flask and the gas is collected by downward delivery method. If the gas is required dry, it is passed through concentrated sulphuric acid.

The diagram showing Laboratory Preparation of dry Sulphur dioxide



Properties of Sulphur dioxide

(a) Physical Properties

Sulphur dioxide has the following properties:

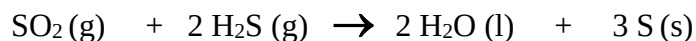
- (i) It is a colourless gas.
- (ii) It has an irritating smell.
- (iii) It is denser than air.
- (iv) It is very soluble in water forming sulphurous acid.
- (v) It turns moist (damp) blue litmus paper red.

(b) Chemical Properties

In its chemical nature sulphur dioxide is an oxidizing agent. It oxidizes hydrogen sulphide and magnesium and is itself reduced to sulphur, as a result, a yellow solid is observed.

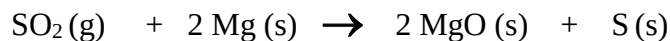
(i) Reaction with hydrogen sulphide

It reacts with moist hydrogen sulphide forming yellow particles of sulphur.



(ii) Reaction with magnesium metal

It reacts with burning magnesium forming yellow particles of sulphur.



(iii) Bleaching action

It bleaches flowers and brown paper.

Chemical Test for Sulphur dioxide

- (i) A filter paper soaked in acidified solution of potassium dichromate is placed in a stream of sulphur dioxide.

Observation: The solution turns from orange to green.

- (ii) Sulphur dioxide is bubbled through acidified potassium permanganate solution.

Observation:

The purple colour of the solution becomes colourless.

Uses of Sulphur dioxide

Sulphur dioxide is used in the following areas:

- (i) Bleaching of pulp in the paper factory during the manufacture of paper.
- (ii) Preservation of tinned food and drinks.
- (iii) Refrigeration as a refrigerant.
- (iv) Manufacture of sulphuric acid by contact process.

Sulphuric acid (H₂SO₄)

Manufacture of sulphuric acid by the Contact Process

The Contact Process

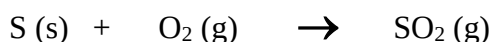
This process uses the following raw materials.

- Sulphur or
- Sulphur dioxide
- Air,
- Concentrated sulphuric acid and
- Water.

There are **four** main stages involved. They include:-

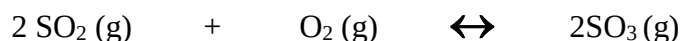
1. Oxidation of sulphur to sulphur dioxide

Sulphur is burned in air to form sulphur dioxide.



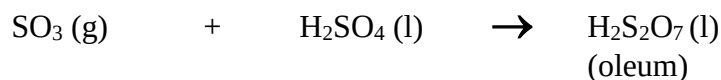
2. Oxidation of sulphur dioxide to sulphur trioxide

The sulphur dioxide is mixed with air. The mixture is dried and passed over vanadium (V) oxide catalyst heated at 450 °C – 500 °C and pressure of 1 atmosphere. The sulphur dioxide is oxidized to sulphur trioxide.



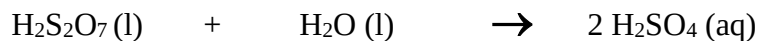
3. Absorption of Sulphur trioxide

The sulphur trioxide formed is absorbed into concentrated sulphuric acid forming a fuming liquid called '*oleum*'.

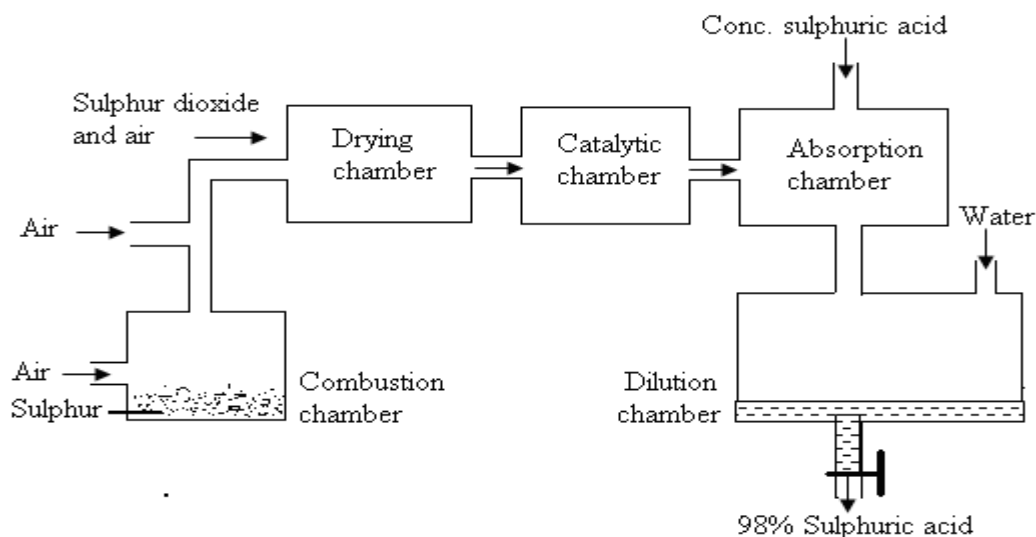


4. Absorption, cooling and Dilution

The oleum formed is absorbed, cooled and carefully diluted with the correct amount of water to form ordinary, sulphuric acid which is 98%.



The flow diagram for the Contact Process



NOTE:

- (i) Vanadium (V) oxide is preferred to platinum as a catalyst in the Contact process because of the following reasons:
 - It is inexpensive i.e. it is cheaper than platinum.
 - Unlike platinum it is not easily poisoned by impurities.
- (ii) Sulphur trioxide is not dissolved in water directly but in Concentrated sulphuric acid because:
 - There is a violent reaction with the water resulting to a mist of fine drops of sulphuric acid which fills the factory. This would damage equipment and harm the plant operators.

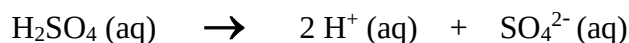
Properties of Sulphuric acid

The properties of sulphuric acid are divided into three categories. These are where the acid acts as:

- an acid when dilute,
- an oxidizing agent and
- a dehydrating agent, when concentrated.

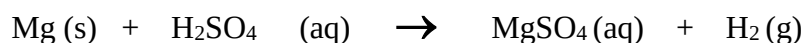
(a) Sulphuric acid as an acid when dilute

Dilute sulphuric acid dissociates completely into H^+ ions and SO_4^{2-} ions.

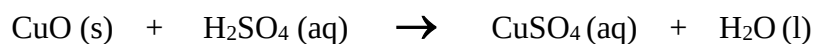


The dilute acid reacts with the reactive metals above hydrogen in the electrochemical series, bases, carbonates in the usual way.

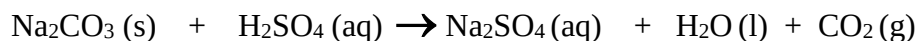
E.g. (i) With magnesium metal



(ii) With copper (II) oxide



(iii) *With sodium carbonate*



(b) Sulphuric acid as an oxidising agent when concentrated

Reaction with copper metal

When a mixture of concentrated sulphuric acid and copper is heated, the copper dissolves with effervescence, giving off sulphur dioxide and leaving a blue solution of copper (II) sulphate.



NB: *This reaction is one of the reactions for the laboratory preparation of sulphur dioxide.*

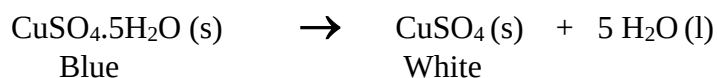
(c) Sulphuric acid as a dehydrating agent when concentrated

Concentrated sulphuric acid is hygroscopic. That is it has a great affinity for water. It removes the elements of water from substances.

The dehydrating properties of concentrated sulphuric acid are seen the following reactions:

(i) Dehydration of hydrated copper (II) sulphate

When concentrated sulphuric acid is added to blue crystals of copper (II) sulphate, it removes the water molecules leaving white crystals of anhydrous copper (II) sulphate crystals.

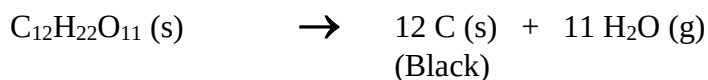


Observation:

The blue crystals of copper (II) sulphate turn to white powder.

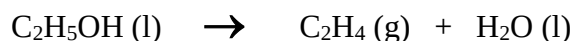
(ii) Dehydration of hydrated sugar crystals

When concentrated sulphuric acid is added to sugar crystals, the white crystals turn dark, then swell up forming a black porous mass of carbon. A lot of heat is given out. This shows that the reaction is exothermic.



(iii) Dehydration of ethanol to form ethene

When a mixture of excess concentrated sulphuric acid and ethanol is heated at a temperature of 180°C, ethene is formed according to the equation:



Uses of Sulphuric acid

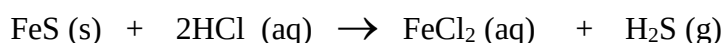
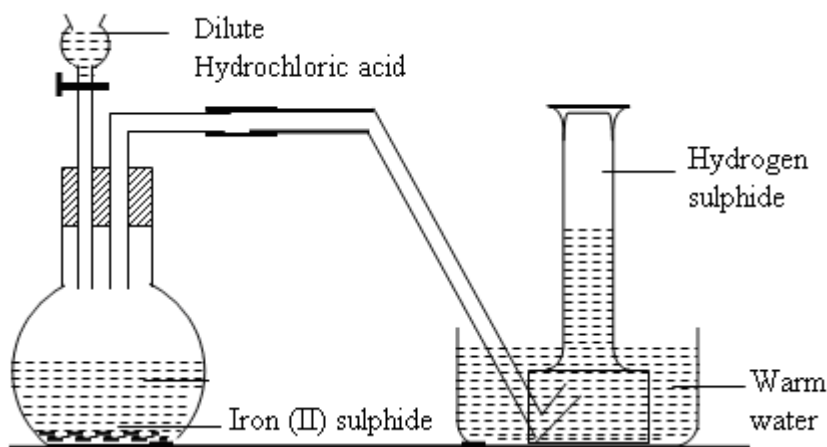
Sulphuric acid is used in the manufacture of:

- (i) fertilizers
- (ii) drugs
- (iii) dyes
- (iv) detergents
- (v) plastics

Hydrogen Sulphide

Laboratory Preparation of Hydrogen Sulphide

Hydrogen Sulphide is prepared in the laboratory by the action of dilute hydrochloric acid on iron (II) sulphide. And the gas is collected by downward delivery method or over warm water as shown in the diagram below.



If the gas is required dry, it is passed through calcium chloride packed in a u-tube and is collected by down ward delivery method.

Properties of Hydrogen Sulphide

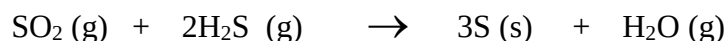
(a) Physical Properties

- (i) It is colourless gas.
- (ii) It is poisonous gas.
- (iii) It has a foul smell of rotten egg.
- (iv) It is denser than air.

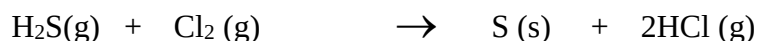
(b) Chemical Properties

It is a good reducing agent, being oxidized to sulphur, a yellow solid. This is seen in the following reactions.

(i) Reaction with Sulphur dioxide



(ii) Reaction with Chlorine



Pollution of Sulphur Compounds

Sulphur is a major impurity in petroleum and coal. The combustion of these fuels releases hydrogen sulphide, sulphur dioxide, carbon monoxide and carbon dioxide into the atmosphere. These gases are air pollutants. They lead to the degradation of the environment.

Explanation

Sulphur dioxide and hydrogen sulphide are poisonous gases.

(i) Sulphur dioxide

Sulphur dioxide readily dissolves in rain water forming sulphurous acid. This sulphurous acid is oxidized to sulphuric acid by aerial oxygen.

Effect of the acid rain

The acid rain increases soil acidity, which dissolves away minerals leading to low yield of agricultural produce. The acid rain also dissolves away minerals in stone works of buildings thus weakening the buildings.

(ii) Hydrogen sulphide

Hydrogen sulphide also leads to formation of acid rain, but more directly darkens buildings, which have lead paints. The darkening is due to formation of black lead (II) sulphide.

SELF-CHECK 6.1

1. (2005 Q.3).

Which one of the following compounds is formed when excess sulphur dioxide is passed through sodium hydroxide solution?

- | | |
|---------------------|------------------------------|
| A. Sodium sulphate. | B. Sodium hydrogen sulphite. |
| C. Sodium sulphide. | D. Sodium hydrogen sulphate. |

2. (2004Q. 26)

The **false** statement about Sulphur dioxide is that sulphur dioxide

- | | |
|------------------------------------|----------------------------------|
| A. is a colourless gas. | B. dissolves readily in water. |
| C. turns moist red litmus to blue. | D. behaves as a bleaching agent. |

3. (2002 Q.15) Sulphur dioxide is used in the following ways except in the

- A. bleaching of materials such as wool and silk.
- B. large-scale production of sulphuric acid.
- C. treating of wool pulp when manufacturing paper.
- D. manufacture of vulcanised rubber.

4. (2002 Q.29) Which one of the following processes adds sulphur dioxide into the atmosphere?

- | | |
|---------------------------------|---|
| A. Burning coal and oil. | B. Decaying organic matter. |
| C. Heating limestone in a kiln. | D. Fractional distillation of liquid air. |

5. (2002 Q.38) Which one of the following statements about the reaction between iron and sulphur is **false**? The reaction
- | | |
|-------------------------------------|---------------------------------|
| A. is an oxidation-reduction one. | B. produces iron (II) sulphate. |
| C. proceeds with evolution of heat. | D. is induced by heat. |
6. (2001 Q.12) Which one of the following compounds is used as catalyst in the manufacture of sulphur trioxide?
- | | |
|--------------------------|------------------------|
| A. Alumina. | B. Vanadium (V) oxide. |
| C. Manganese (IV) oxide. | D. Iron powder. |
7. (2000 Q.13) Which one of the following reagents is used to test for sulphur dioxide?
- | | |
|---------------------|--------------------------------------|
| A. Chlorine water. | B. Acidified potassium permanganate. |
| C. Cobalt chloride. | D. Anhydrous copper sulphate. |
8. (1997 Q.5) The products of the reaction between concentrated sulphuric acid and copper are
- | |
|--|
| A. water, sulphur dioxide, and copper (II) sulphate. |
| B. water and copper (II) sulphate. |
| C. water, copper (II) sulphate and hydrogen. |
| D. sulphur dioxide and copper (II) sulphate. |
9. (1997 Q.24) Sulphur dioxide reacts with oxygen to form sulphur trioxide according to the equation: $\text{SO}_{2(g)} + \text{O}_{2(g)} \rightleftharpoons 2\text{SO}_{3(g)} + \text{heat}$. Which one of the following conditions favours the formation of sulphur trioxide?
- | | |
|---------------------------------------|--|
| A. Low pressure and low temperature. | B. High pressure and high temperature. |
| C. Low pressure and high temperature. | D. High pressure and low temperature. |
10. (1996 Q.7) Which one of the following gases turns potassium dichromate solution green?
- | | | | |
|--------------------|--------------------|--------------------|--------------------|
| A. NH_3 . | B. NO_2 . | C. Cl_2 . | D. SO_2 . |
|--------------------|--------------------|--------------------|--------------------|

SELF-CHECK 6.2

SULPHUR AND ITS COMPOUNDS

1. (1991 Q.11).
- Describe how sulphur is extracted by the Frasch process.
 - Write equations to show how fuming sulphuric acid can be obtained from sulphur.
 - State what would be observed if concentrated sulphuric acid is added to sugar.
 - State any two uses of sulphur.
2. (2003 Q.5).
- Write an ionic equation to show how sulphur dioxide can be formed from sodium sulphite and hydrochloric acid.
 - Name **one** reagent that can be used to test for sulphur dioxide.
 - State what would be observed if sulphur dioxide was reacted with the reagent you have named in (b) (i).
 - Sulphur dioxide was passed into a beaker containing a red flower and water.
 - State what was observed.
 - Give a reason for your answer in (c) (i).

3. (2002 Q.11).

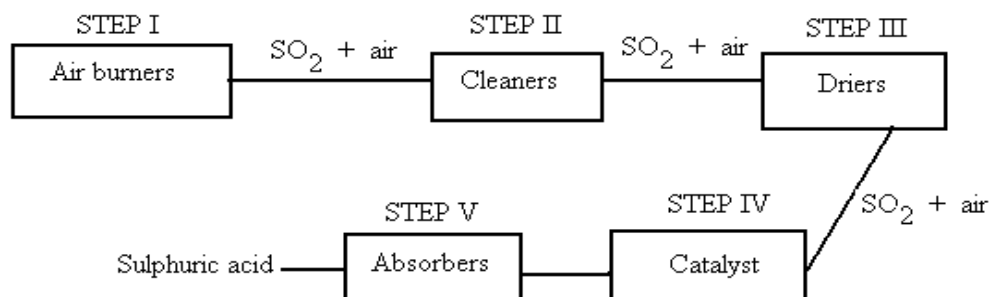
- (a) (i) Name **one** substance that is reacted with hydrochloric acid to produce sulphur dioxide in the laboratory.
(ii) State the conditions for the reaction.
(iii) Name a substance that can be used to dry the sulphur dioxide.
(iv) Write equation for the reaction leading to the formation of sulphur dioxide.
- (b) State what would be observed and explain what would happen if sulphur dioxide is passed through a solution containing
(i) Acidified potassium dichromate.
(ii) A dye.
- (c) Briefly describe how sulphur dioxide can be converted to sulphuric acid. Your answer should include equations and conditions for the reaction(s).

4. (1999 Q.7).

- (a) State what would be observed if concentrated sulphuric acid is added to sugar. Explain your answer.
- (b) Dilute sulphuric acid was added to zinc carbonate. State what was observed and write equation for the reaction.
- (c) State the conditions under which sulphuric acid reacts with copper and write equation for the reaction.
- (d) Describe a test that can be carried out to identify the sulphate ion in sulphuric acid.
- (e) State **one** use of sulphuric acid.

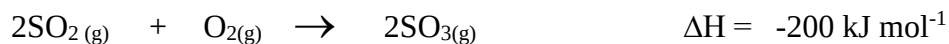
5. (1996 Q.13).

The following flow chart shows the steps in the manufacture of sulphuric acid by the Contact process.



- (a) Write an equation for the reaction that takes place in step I.
- (b) Why is step II necessary?
- (c) Name:
(i) The drying agent in step III.
(ii) The catalyst in step IV.

- (d) Describe the process that takes place in step V.
- (e) Sulphur dioxide combines with air to form sulphur trioxide according to the equation:



- (i) Give three conditions for maximum yield of sulphur trioxide.
- (ii) Calculate the volume of sulphur trioxide that would be formed when 20 cm^3 of sulphur dioxide was reacted with 100 cm^3 of oxygen.
- (f) Give **one** use of sulphuric acid.
- (g) Concentrated sulphuric acid is 18 M. Calculate the volume of concentrated sulphuric acid required to make 2 litre of 2 M sulphuric acid.

6. (1992 Q.11).

- (a) Draw a labelled diagram to show how a dry sample of sulphur dioxide can be prepared in the laboratory.
- (b) Write equation for the reaction that takes place in (a).
- (c) Describe a test that can be carried out to confirm the presence of sulphur dioxide.
- (d) Excess sulphur dioxide was bubbled through a solution of sodium hydroxide. Write equation for the reaction that took place.
- (e) 25.0 cm^3 of 0.1 M sodium hydrogen carbonate solution reacted completely with 27.8 cm^3 of sulphuric acid. Calculate the concentration of sulphuric acid in moles per litre.

CHAPTER SEVEN

CHLORINE AND ITS COMPOUNDS

Chlorine

Chlorine, symbol Cl, with atomic number 17 and mass number 35.5, is one of the halogens. It belongs to group VII and period III in the Periodic Table.

Occurrence:

Chlorine does not occur as a free element. It occurs in combined form as ionic and covalent chloride compounds.

Preparation of Chlorine gas

In the laboratory, chlorine is prepared from concentrated hydrochloric acid by oxidation with:

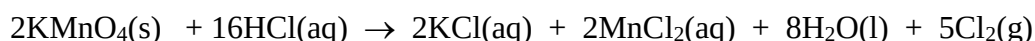
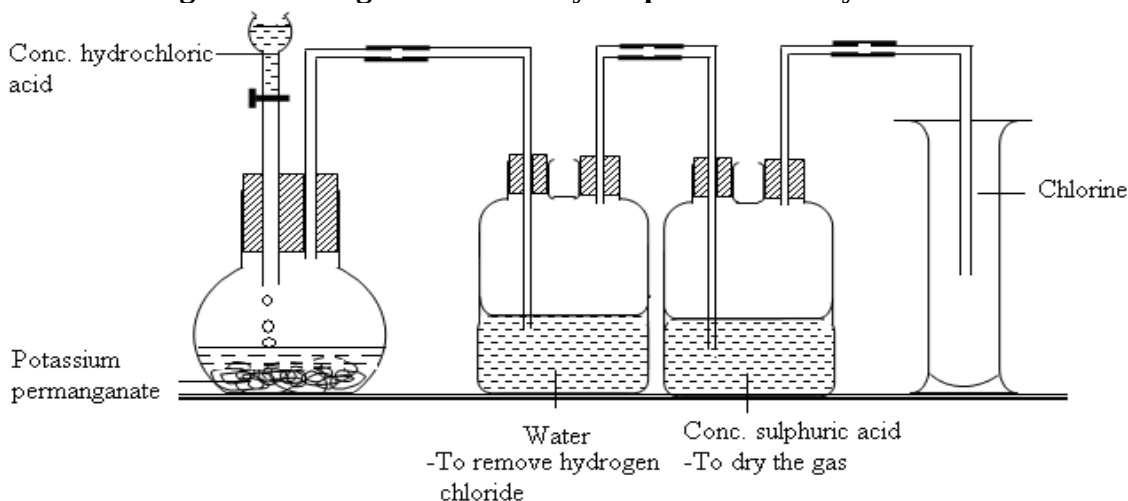
- Potassium permanganate and
- Manganese (IV) oxide.

With potassium permanganate, the reaction takes place in the cold and with manganese (IV) oxide, the mixture is heated. The gas produced in both reactions is passed through water to remove any hydrogen chloride gas and then through concentrated sulphuric acid to dry the gas.

Laboratory Preparation of dry Chlorine using Conc. Hydrochloric acid and potassium permanganate

Chlorine is prepared in the laboratory by the action of concentrated hydrochloric acid on potassium permanganate crystals in a flat-bottomed flask. Effervescence occurs in the cold evolving chlorine gas. And the gas is passed through water to absorb/remove any hydrogen chloride gas and then through concentrated sulphuric acid to dry the gas. It is then collected by downward delivery method since it is denser than air as shown in the diagram below.

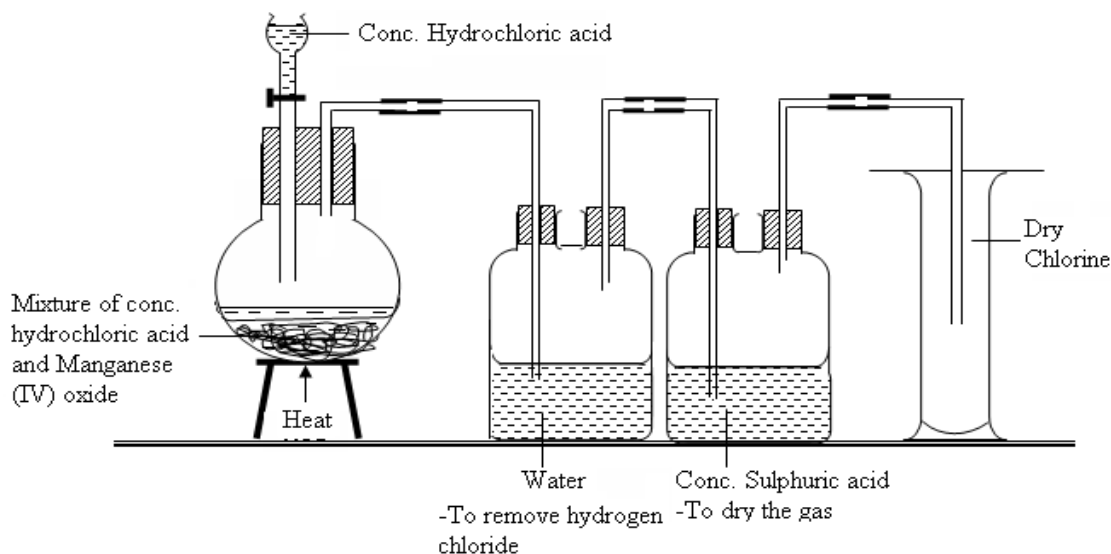
The diagram showing the Laboratory Preparation of Dry Chlorine



Laboratory Preparation of dry Chlorine using Conc. Hydrochloric acid and manganese (IV) oxide

Chlorine is prepared in the laboratory by heating a mixture of concentrated hydrochloric acid and manganese (IV) oxide crystals in a round-bottomed flask. Effervescence occurs, evolving chlorine gas. And the gas is passed through water to absorb/remove any hydrogen chloride gas and then through concentrated sulphuric acid to dry the gas. It is then collected by downward delivery method since it is denser than air as shown in the diagram below.

The diagram showing the Laboratory Preparation of Dry Chlorine



Properties of Chlorine

(a) Physical Properties

- (i) Chlorine is a yellowish green gas.
- (ii) It has a choking, unpleasant, irritating smell.
- (iii) It is denser than air (being about $2\frac{1}{2}$ times as dense as air).
- (iv) It is very poisonous gas.
- (v) It is fairly soluble in water (which dissolves about 3 times of its volume of the gas) forming yellowish water, contains hydrochloric acid and chloric (I) acid.



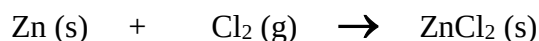
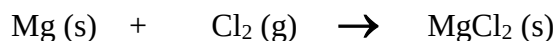
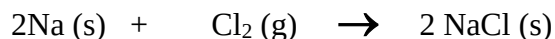
- (vi) It readily liquefies under pressure of 6.8 atmospheres, at 20°C .
- (vii) It melts at -101°C .
- (viii) It boils at -34.05°C , 1 atmosphere pressure, and
- (ix) Has a specific gravity of 1.41 at -35°C .

Chemical Properties

Chlorine is an active element. It reacts with many metals, non-metals, organic compounds (hydrocarbons), alkalis, some metallic halogen compounds and with water. These properties are seen in the following reactions.

(a) Reaction with metals (Sodium, Magnesium and Zinc)

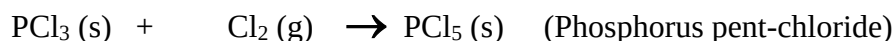
When burning pieces of the above metals are plunged into gas jar of chlorine, they continue to burn in chlorine forming white clouds, which settles as white solid.



(b) Reaction with non-metals (Phosphorus, Sulphur and Hydrogen)

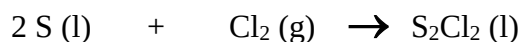
(i) Reaction with Phosphorus.

When a piece of white phosphorus is plunged into dry chlorine, it catches fire and burns forming white fumes of phosphorus tri-chlorides and pent-chloride.



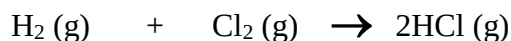
(ii) Reaction with Sulphur

When dry chlorine is bubbled into molten sulphur, a red liquid, disulphur dichloride is formed.



(iii) Reaction with Hydrogen

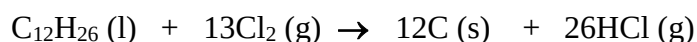
The mixture of hydrogen and chlorine does not react in darkness, reacts slowly in diffused sunlight and explodes if placed in bright sunlight. Burning hydrogen continues to burn quietly in chlorine, forming hydrogen chloride.



(c) Reaction with Hydrocarbons

Burning hydrocarbon continues to burn in chlorine forming clouds of black solid particles (soot – carbon) and hydrogen chloride.

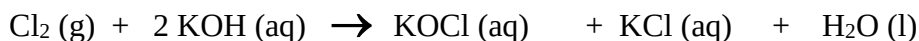
For example, when a filter paper or piece of cotton wool soaked in paraffin is plunged into a gas jar of dry chlorine, it burns with a reddish flame forming clouds of soot and hydrogen chloride.



(d) Reaction with alkalis

(i) Reaction with cold dilute alkaline solution

When chlorine is bubbled through cold dilute aqueous solution of sodium hydroxide or potassium hydroxide, it is absorbed forming a pale yellow solution of the hypochlorite and chloride of the metal.

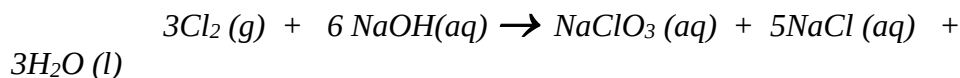


Ionicly the reaction equation is written as:

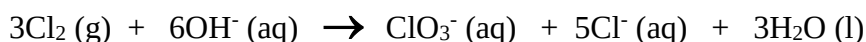


(ii) Reaction with hot dilute alkaline solution.

When chlorine is bubbled through hot concentrated aqueous solution of sodium hydroxide or potassium hydroxide, for some time, a mixture of the chloride and chlorate of the metal is formed and the later can be obtained by crystallizing the mixture where crystals of chlorate separate first. (These can be purified by crystallization).



Ionicly the reaction equation is written as:

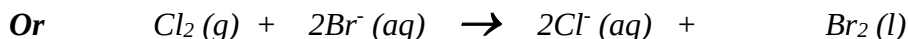
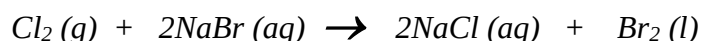


(e) Reaction of Chlorine with the compounds of Halogens, bromides and iodides.

Chlorine displaces the halogens below it in the Periodic Table from their compounds.

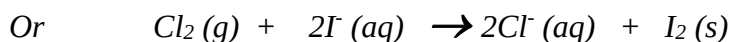
(i) Reaction with Sodium bromide

When chlorine is bubbled into a solution of sodium bromide, the solution turns from colourless to orange.



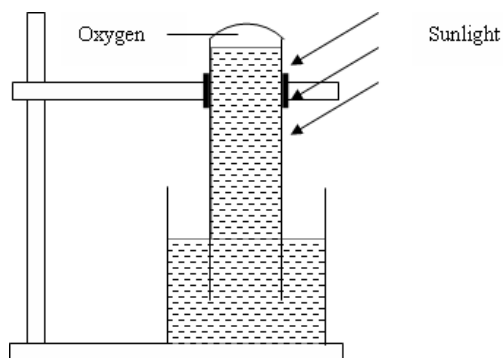
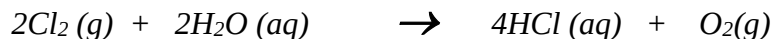
(ii) Reaction with potassium iodide

When chlorine is bubbled into a solution of potassium iodide, the solution turns from colourless to brown. And a black solid is also deposited. The brown-black colour is due to the formation of solid iodine discharged.



The action of sunlight on Chlorine water

When chlorine water filled in a long glass tube is inverted in a beaker containing some of the chlorine water exposed to bright sunlight, after some time, a colorless gas that rekindles a glowing splint collects above the water in the U-tube.



The above reaction probably occurs in two stages, as indicated in the equations below:

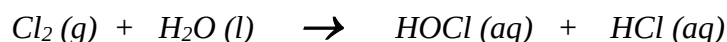


The unstable hypochlorous acid gives up its oxygen forming hydrochloric acid.



The Bleaching action of Chlorine

Chlorine dissolves in water and reacts with it forming hypochlorous acid and hydrochloric acid.

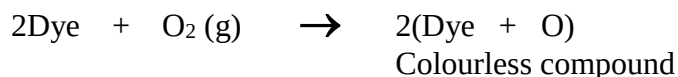


Chlorine dissolved in water is known as chlorine water. It is a bleaching agent.

The bleaching action of chlorine is due to the ion, OCl^- , in the hypochlorous acid. The hypochlorous acid is a very reactive compound and readily gives up its oxygen.



The oxygen given off reacts with the dye, to form a colourless compound.



Uses of Chlorine

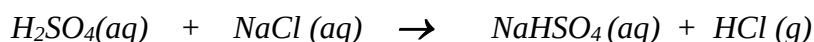
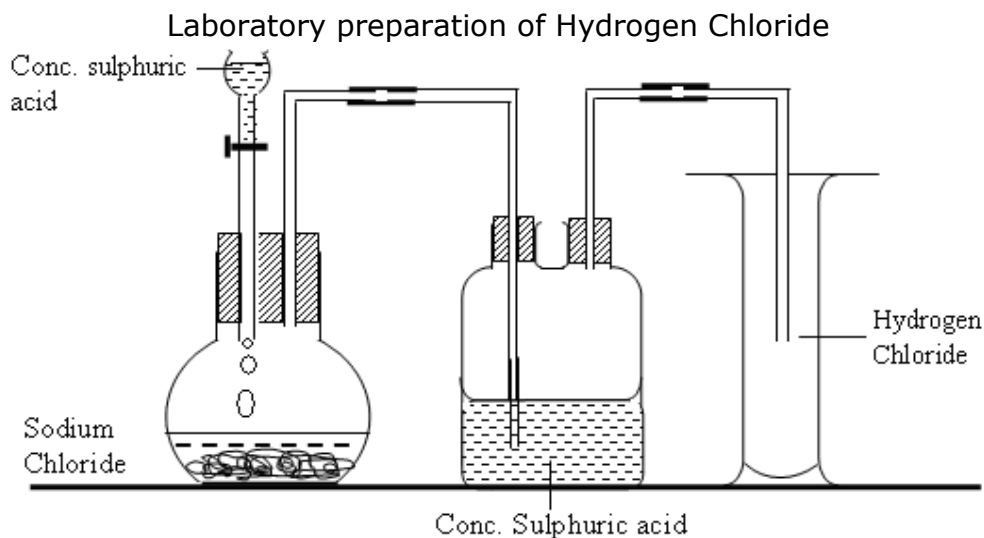
Chlorine is used to:

- (i) Sterilize (Treat) water for domestic and industrial and swimming baths.
- (ii) Manufacture plastics.
- (iii) Manufacture insecticides
- (iv) Manufacture antiseptics.
- (v) Manufacture hydrochloric acid.
- (vi) Bleach textiles and paper in industries.

Hydrogen Chloride

Laboratory preparation of Hydrogen Chloride

Hydrogen Chloride is prepared in the laboratory by adding concentrated sulphuric acid on sodium chloride in a flat bottom flask. Effervescence occurs and the gas is dried by passing it through concentrated sulphuric acid and is collected by downward delivery method, as it is denser than air.



Properties of Hydrogen Chloride

(a) Physical Properties

- (i) It is a colourless gas.
- (ii) It is denser than air.
- (iii) It is very soluble in water, being the second to ammonia (1 cm^3 of water dissolves about 540 cm^3 of the gas), as a result it can be used to perform fountain experiment.
- (iv) It is acidic and therefore turns a moist blue litmus paper red.

(b) Chemical Properties

- (i) It forms white dense fumes with ammonia gas.
 $\text{HCl}(\text{g}) + \text{NH}_3(\text{g}) \rightarrow \text{NH}_4\text{Cl}(\text{s})$
- (ii) It forms a white precipitate with silver nitrate solution.
 $\text{HCl}(\text{g}) + \text{AgNO}_3(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{HNO}_3(\text{aq})$
The precipitate dissolves in excess ammonia solution forming a colourless solution.
- (iii) Dry hydrogen chloride reacts with heated iron to form iron (II) chloride and hydrogen.
 $2\text{HCl}(\text{g}) + \text{Fe}(\text{s}) \rightarrow \text{FeCl}_2(\text{s}) + \text{H}_2(\text{g})$

Chemical test for Hydrogen Chloride

The gas forms misty fumes in moist air, turns blue litmus red, forms white precipitate with acidified silver nitrate solution.

SELF-CHECK 7.1

1. (2003 Q.39)

Chlorine dissolves in cold aqueous solution of sodium hydroxide to produce the following substance.

- A. Sodium chlorate.
B. Sodium chloride.
C. Sodium chlorite.
D. Sodium hydrogen chloride

2. (2002 Q.40).

The reaction of iron and chlorine

- A. occurs at ordinary temperature.
B. requires platinum catalyse.
C. produces iron (II) chloride.
D. produces iron (III) chloride

3. (2001Q.9).

The best reason for including water in the laboratory preparation of chlorine is to

- A. remove chlorine gas.
B. cool chlorine gas.
C. cool hydrogen chloride gas.
D. remove hydrogen chloride gas.

4. (2000 Q.26)

A compound reacts with concentrated sulphuric acid to give a colourless gas which fumes with ammonia. The compound contains a

- A. chloride. B. carbonate. C. nitrate. D. sulphate

5. (1998 Q.20)

What is the correct equation for the reaction between chlorine and heated steel wool?

- A. $\text{Cl}_{2(\text{g})} + 2\text{Fe}_{(\text{s})} \rightarrow 2\text{FeCl}_{(\text{s})}$
 B. $\text{Cl}_{2(\text{g})} + \text{Fe}_{(\text{s})} \rightarrow \text{FeCl}_{2(\text{s})}$
 C. $3\text{Cl}_{2(\text{g})} + 2\text{Fe}_{(\text{s})} \rightarrow 2\text{FeCl}_{3(\text{s})}$
 D. $3\text{Cl}_{2(\text{g})} + 4\text{Fe}_{(\text{s})} \rightarrow 2\text{Fe}_2\text{Cl}_{3(\text{s})}$

6. (1997 Q.4)

The gas formed when chlorine water is exposed to sunlight is

- A. oxygen.
B. hydrogen.
C. chlorine.
D. hydrogen chloride.

7. (1996 Q.13)

Which one of the following gases is an oxidising agent?

- A. CO. B. H₂S. C. Cl₂. D. NH₃.

8. (1996 Q.26) Chlorine can be prepared in the laboratory by reacting

- A. sodium chloride and concentrated ethanoic acid.
B. potassium chlorate and concentrated sulphuric acid.
C. sodium chloride and concentrated sulphuric acid.
D. potassium permanganate and concentrated hydrochloric acid.

9. (1995 Q.11)

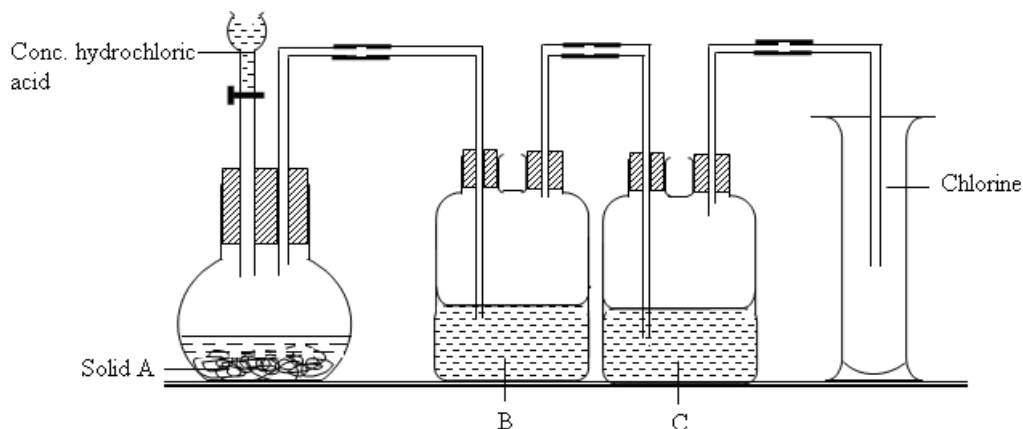
Which one of the methods below is used to prepare a sample of copper (II) chloride?

- A. Heating copper in stream of dry chlorine.
B. Action of copper on dilute hydrochloric acid.
C. Reacting copper (II) oxide with hydrochloric acid.
D. Heating copper in a stream of dry hydrogen chloride

SELF-CHECK

7.2

1. The diagram in Fig.1 shows a set up of the apparatus for the laboratory preparation of dry chlorine from hydrochloric acid.



- (a)
 - (i) Name substances A, B and C.
 - (ii) State the role of substance B.
 - (b) State the conditions for the reaction.
 - (c) Write equation for the reaction.
2. (2002 Q.9).
- (a) Chlorine can be prepared from concentrated hydrochloric acid.
 - (i) Name a substance that can react with hydrochloric acid to produce chlorine.
 - (ii) Write equation for the reaction.
 - (b) Chlorine gas was passed through cold dilute sodium hydroxide solution.
 - (i) State what was observed.
 - (ii) Write equation for the reaction that took place.
3. (1999 Q.14). Chlorine
- (a)
 - (i) Describe with the aid of a well-labelled diagram how a dry sample of chlorine can be prepared in the laboratory.
 - (ii) Write an equation for the reaction that took place.
 - (iii) State any **three** uses of chlorine.
 - (b) State with the aid of equations, what would be observed if chlorine was added to:
 - (i) Iron (II) chloride solution,
 - (ii) Potassium iodine solution.
 - (c) Burning sodium was plunged into a jar of chlorine.
 - (i) State and explain what was observed.
 - (ii) Write the equation for the reaction.

4. (1997 Q.2) Chlorine

During the preparation of chlorine in the laboratory, the gas may be passed through water and concentrated sulphuric acid before collection.

- (a) State the use of:
 - water.
 - concentrated sulphuric acid.
 - (b) Chlorine is a bleaching agent when in the presence of water.
 - (i) Write an equation for the reaction between chlorine and water.
 - (ii) Using equations explain the bleaching action of chlorine.
 - (c)
 - (i) State what would be observed if chlorine was bubbled through a solution of iron (II) sulphate solution.
 - (ii) Write an ionic equation for the reaction between chlorine and iron (II) ions.
5.
 - (a) When lead (IV) oxide was added to concentrated hydrochloric acid and heated, the gas evolved bleached litmus.
 - (i) Identify the gas.
 - (ii) Write the equation for the reaction.
 - (b) When chlorine was passed into a solution of potassium bromide, a red colouration developed owing to liberation of bromine
 - (i) Write the equation for the reaction.
 - (ii) Describe what will be observed when fluorine is passed into potassium bromide solution, given that fluorine is just above chlorine in the group
 - (iii) Write the ionic equation for the reaction in b (ii) above.
6.
 - (a) When concentrated hydrochloric acid was added to manganese(IV)oxide and heated the gas, thus evolved, bleached litmus paper. Write the equation to show what happened
 - (b) Chlorine dissolves in water to give chlorine water
 - (i) Write the equation for the reaction that occurs
 - (ii) When the chlorine water is exposed to sunlight a colourless gas which gives a positive test with a glowing splint is given off. Explain.
 - (c) Chlorine can react with a cold dilute solution of potassium hydroxide to form water and two other products.
 - (i) Write the equation for the reaction
 - (ii) Name the other two products
 - (iii) Give the use of the solution obtained above

SELF-CHECK 7.3

HYDROGEN CHLORIDE

7. (2003 Q.12).

- (a) Write equation to show how hydrogen chloride can be prepared from sodium chloride.
- (b) Draw a labelled diagram to show how aqueous hydrogen chloride can be prepared in the laboratory.
- (c) State what would be observed and write equation for the reaction that would take place when aqueous hydrogen chloride is reacted with:
 - (i) Solid calcium carbonate.
 - (ii) Silver nitrate solution.
 - (iii) Magnesium.
- (d) State why aqueous hydrogen chloride does **not** react with copper.

8. (1998 Q.11)

- (a) Describe how a dry sample of hydrogen chloride can be prepared from a named chloride. (Diagram **not** required). Your answer should include the following:
 - Conditions for the reaction.
 - Name of the drying agent.
 - Method of collection.
 - Equation for the reaction.
- (b) Name the substance that is formed when hydrogen chloride is passed through water.
- (c) (i) Name **one** reagent that can be used to test for the presence of chloride ions in solution.
 - (ii) State what would be observed if the reagent was added to chloride solution.
- (d) Write an ionic equation to show the reaction between aqueous hydrogen chloride and calcium hydrogen carbonate solution.
- (e) 25.0 cm³ of a 0.2 M lead (II) nitrate solution was shaken with excess aqueous hydrogen chloride. Lead (II) ions react with chloride ions according to the following equation: $\text{Pb}^{2+}(\text{aq}) + 2 \text{Cl}^{-}(\text{aq}) \rightarrow \text{PbCl}_2(\text{s})$.
Calculate the mass of lead (II) nitrate formed.

9. (1991 Q.14)

- (a) Draw a well-labelled diagram to show how a sample of dry hydrogen chloride can be prepared.
 - (b) Dry hydrogen chloride gas was passed over heated iron fillings. Write an equation for the reaction that took place.
 - (c) The solid product in (b) was dissolved in water and aqueous sodium chloride added to the resultant solution drop wise until in excess.
 - (i) State what was observed.
 - (ii) Write equation for the reaction.
 - (d) Chlorine gas was passed through a solution of the product in (b).
 - (i) State what was observed.
 - (ii) Write an ionic equation for the reaction.
 - (e) Name one reagent that can be used to test for
 - (i) The cation formed in (d).
 - (ii) The anion formed in (d).
- In each cases state what is observed when the reagent you have named is used.

CHAPTER EIGHT

ION CHEMISTRY

INTRODUCTION

Chemistry Practical paper tests the knowledge of Candidates in two major areas, namely:

- Qualitative Analysis and
- Volumetric Analysis.

QUALITATIVE ANALYSIS

Qualitative analysis deals with the identification of **cations** (positively charged ions) and **anions** (negatively charged ions) from unknown compound(s) usually a mixture of salts.

However, reaching the ions involves several steps before arriving at the right ions. The steps act as a pointer thus enabling the candidate to predict the unknown ions present in the unknown compound.

The main steps involved are the following: -

- Preliminary Examination.
- Action of heat.
- Action of water (Solubility).
- Action of acids.
- Action of Soluble salts.
- Action of Sodium hydroxide.
- Action of Ammonia Solution.

All the deductions/ predictions made as a result of the above steps are always confirmed by carrying out specific test known as **confirmatory test**.

1. Preliminary Examination

The unknown compound is examined by applying some of the sense organs (excluding the tongue).

- Note the: -
- Appearance
 - Colour
 - Smell
 - Feel

Summary of the Preliminary Test

OBSERVATION	DEDUCTION
Coloured substance	Salt of transition metal. E. g. Copper salt, iron (II) or iron (III) salt
White powder or crystal	Salt of group I, II, III, or Zinc
Green crystal	Iron (II) or copper (II) salt.
Brown solid	Iron (III) salt.
Smell of ammonia	Ammonium salt

2. Action of Heat

A little (usually a spatula end- full) of the unknown compound is heated strongly in an ignition tube or in a small dry test tube until no further change occurs.

Action of heat enables the candidates to search for: -

- Evolution of gases.
- Evolution of water vapour.
- Change in colour of the oxide formed.

Summary of the Possible Observations and Deductions

OBSERVATION	DEDUCTION
Water vapour condenses on the cooler part of the test tube	Water of crystallization from a hydrated salt of :- HCO_3^- , HSO_4^- , SO_4^{2-}
A gas that turns lime water milky given off.	Carbonate (CO_3^{2-}), Hydrogen Carbonate (HCO_3^-)
A colourless gas that turns red litmus paper blue given off.	Ammonium salt
Brown fumes of a gas given off.	Nitrate (NO_3^-)
Cracking sound and brown fumes of a gas given off.	Lead (II) nitrate
A gas that turns moist blue litmus paper red given off.	Sulphite (SO_3^{2-}) or certain sulphates (SO_4^{2-})
Colour change of the oxide formed: - Yellow when hot and - White when cold. - Reddish brown when hot. - Yellow when cold.	Zinc oxide Lead (II) oxide

3. Action of Acids

Acids have two major roles to play when added to the unknown compound in a test tube.

- These are:
- to dissolve an insoluble salt. E. g A carbonate or a sulphite of group II or a transition metal.
 - to acidify the solution for the next test.

(a) Addition of dilute Hydrochloric acid or dilute nitric acid

OBSERVATION	DEDUCTION
A colourless gas that turns lime water milky given off.	CO_3^{2-} , HCO_3^-
A colourless gas that turns moist blue litmus paper red given off.	Sulphite (SO_3^{2-})
A colourless gas with a smell of rotten egg.	Sulphide (S^{2-})

(b) Addition of Concentrated Sulphuric acid

OBSERVATION	DEDUCTION
A gas that turns lime water milky given off.	CO_3^{2-} , HCO_3^- ions.
A gas that turns moist blue litmus paper red given off.	Sulphite (SO_3^{2-}) ions
A gas which smells like a rotten egg.	Sulphide (S^{2-}) ions

NB: Both acids evolve the same gases i. e. CO_2 , SO_2 , & H_2S , except NO_2 , which is only evolved by concentrated Sulphuric acid.

4. IDENTIFICATION OF GASES

Once a gas is evolved under the action of heat or an acid, candidates are advised to carryout the specific tests to identify the gas so as to draw a concrete deductions of the anion present.

Gases are identified according to their: -

- colour.
- smell.
- action on moist litmus paper.
- action with glowing splint.
- action on lime water.

The table below shows the common gases that may be encountered during examinations

GAS	COLOUR	SMELL	ACTION ON LITMUS PAPER	ACTION ON GLOWING SPLINT
Oxygen (O ₂)	None	None	None	Relights a glowing splint
Carbon dioxide (CO ₂)	None	None	Blue – faint red.	Does not burn
Hydrogen sulphide (H ₂ S)	None	Rotten egg	Blue – faint red.	Burns with blue flame.
Sulphur dioxide (SO ₂)	None	Burning sulphur	Blue – Red	Does not burn.
Ammonia (NH ₃)	None	Choking characteristic smell	Red – blue	Does not burn.

NB: *Ammonia is the ONLY gas that turns red moist litmus blue.*

5. Solubility (Action of water)

Candidates are often required to add a small amount of the unknown compound in a given amount of water. The main aim of this step is to free the ions of the soluble salt. Action of water therefore tests for the solubility of the salt.

Usually a mixture of soluble and insoluble salts is provided. Therefore candidates are always instructed to add a spatula end- full of the unknown compound to some amount of water, shake well, filter and keep both filtrate and the residue.

Candidates should therefore know that the filtrate contains ions of a soluble salt and the residue contains the insoluble salt.

The table below shows the summary of the solubility of the common salts.

SOLUBLE SALTS	INSOLUBLE SALTS
All group I/ Ammonium salts.	
All nitrates.	
All Chlorides except:	Lead (II) & Silver chlorides. NB: <i>Lead(II) chloride is soluble in hot water.</i>
All Sulphates except:	Lead(II) & Barium sulphates NB: <i>Calcium sulphate is only slightly soluble.</i>
All group I/ Ammonium carbonates.	All other carbonates.

6. IDENTIFICATION OF ANIONS

Anions are negatively charged ions. The common anions include the following:-

- HCO_3^- , CO_3^{2-}
- HSO_4^- , SO_4^{2-} and SO_3^{2-}
- NO_2^- , NO_3^-
- Cl^-

All anions are tested by using the knowledge of solubility of salts. Suitable cations with which they show characteristic precipitates are added in to their aqueous states.

The common cations used are:

- Ag^+ ,
- Ba^{2+} and
- Pb^{2+}

They are obtained from their soluble salts.

Ag^+ and Pb^{2+} ions are obtained from their nitrates, AgNO_3 , and $\text{Pb}(\text{NO}_3)_2$,

While Ba^{2+} ions are obtained from $\text{Ba}(\text{NO}_3)_2$ or BaCl_2

Stages involved in Testing for Anions

(a) Dry Testing

The substance in the solid form is strongly heated in a dry test tube or ignition tube until no further change.

Summary of Effects, observations and Deductions

EFFECT	OBSERVATION	DEDUCTION
Decompose	A colourless gas that relights a glowing splint (oxygen) given off.	Nitrate of heavy metal or Nitrate of group I
Decompose	A gas that turns lime water milky (CO_2) given off.	Carbonate of heavy metal.
Decompose	A colourless gas that turns limewater milky (CO_2) and water vapour given off.	Hydrogen carbonate of heavy metal
Does not decompose	No observable change.	Carbonate of group I metal.
Decompose	White choking fumes of sulphur trioxide (SO_3) and water vapour.	Hydrogen Sulphate (HSO_4^-)

(b) Wet Testing

This test involves addition of a solution to solution or solution to solid.

It makes the use of precipitation or redox reaction involving colour changes for the test purpose.

This test is used:

- To test for a carbonate.
- To test for a metal oxide or hydroxide and
- To dissolve insoluble substance.

(i) Solution to solid reaction

A solution of a known reagent is added to the solid. The possible observations and deductions are then made.

SUMMARY OF RESULTS

REAGENT ADDED	OBSERVATION	DEDUCTION
Dilute nitric acid (HNO_3)	Effervescence released. A gas that turns lime water milky given off.	Carbonate (CO_3^{2-})
	Solid dissolves with effervescence and releases out heat.	Metal oxide or hydroxide.

NB: *Dilute nitric acid is usually used because the nitrates formed are soluble in water, therefore dissolution continues to the end.
When the salt formed is insoluble in water, it tends to precipitate and coat the solid, rendering it unapproachable and the reaction ceases (stops) at an earlier stage.
This is why dilute hydrochloric acid and sulphuric acid may not be used to dissolve lead carbonate. Lead chloride and lead sulphate are insoluble and would coat the carbonate or chloride to be dissolved thus cutting it off from the acid.*

(ii) Solution to solution reaction

Solutions containing suitable metal ions are used to precipitate out the anions.

The observations and deductions are made.

The above deductions are confirmed by carrying out specific tests called

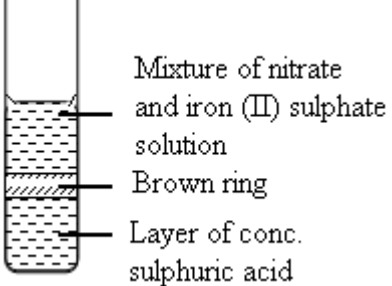
Confirmatory Test

Summary of results Confirmatory Test

REAGENT ADDED	OBSERVATION	DEDUCTION
Cations of heavy metals e. g. Pb^{2+} , Zn^{2+} , Cu^{2+} etc. in their soluble salts and then followed by dilute nitric acid.	Precipitates dissolve with effervescence releasing a colourless gas that turns limewater milky given off.	CO_3^{2-} , or HCO_3^- ions.
(i) Barium nitrate solution followed by nitric acid.	- White precipitate insoluble in the acid.	SO_4^{2-} ions.
(ii) Lead (II) nitrate solution followed by barium nitrate and heat.	-White precipitate which persists on heating.	
-Add equal volume of freshly prepared Iron (II) Sulphate solution to the solution of the unknown compound. Tilt the test tube and then pour conc. Sulphuric acid slowly down the side of the test tube.	Brown ring formed in between the two layers.	NO_3^-
(i) Silver nitrate solution followed by dilute nitric acid.	- White precipitate, insoluble in the acid.	Cl^-
(ii) Silver nitrate solution followed by ammonia solution.	- White precipitate that turns to grey on standing in light forms.	
(i) Lead (II) nitrate solution and heat.	- White precipitate, soluble on heating.	

Confirmation Tests for Anions

A confirmatory test is a specific test to identify specific ions

ANION	TEST	OBSERVATION
SO ₄ ²⁻	(i) Add barium nitrate solution followed by dilute nitric acid or add barium chloride followed by dilute hydrochloric acid.	White precipitate insoluble in the acid. $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$
	(ii) add lead (II) nitrate solution and heat the mixture.	White precipitate formed. Precipitate insoluble on heating $\text{Pb}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{PbSO}_4(\text{s})$
Cl ⁻	(i) Add silver nitrate solution followed by dilute nitric acid.	White precipitate insoluble in the acid. $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$
	(ii) add silver nitrate solution followed by ammonia solution.	White precipitate formed. White precipitate soluble in ammonia solution forming a colourless solution. $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$ $\text{AgCl}(\text{s}) + 2\text{NH}_{3(\text{aq})} \rightarrow \text{Ag}(\text{NH}_3)_2^+(\text{aq}) + \text{Cl}^-(\text{aq})$
	(iii) Add lead (II) nitrate solution and heat the mixture.	White precipitate formed. Precipitate dissolves on heating, and reappears on cooling. $\text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) \rightarrow \text{PbCl}_2(\text{s})$
NO ₃ ⁻	Add equal volume of freshly prepared iron (II) sulphate solution to the solution of unknown compound. Tilt the test tube and then pour conc. sulphuric acid slowly on the side of the test tube.	Two layers form. A brown ring forms between the two layers. 
CO ₃ ²⁻	(i) Add dilute nitric acid	Effervescence of a colourless gas that turns lime water milky given off. $\text{CO}_3^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
	(ii) Add barium nitrate solution followed by dilute nitric acid.	White precipitate formed. Precipitate dissolves with effervescence in the acid giving off a colourless gas that turns lime water milky. $\text{Ba}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{BaCO}_3(\text{s})$ $\text{CO}_3^{2-}(\text{s}) + 2\text{H}^+(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

	(iii) Add barium nitrate or barium chloride solution followed by dilute hydrochloric acid.	White precipitate formed. Precipitate dissolves with effervescence in the acid giving off a colourless gas that turns lime water milky. $\text{Ba}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{BaCO}_3(\text{s})$ $\text{CO}_3^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
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7. IDENTIFICATION OF CATIONS

The common cations encountered by candidates during examinations include the following:-

GROUP	CATION
I	Na^+ , K^+ ,
II	Ca^{2+} and Mg^{2+}
III	Al^{3+}
IV	Pb^{2+}
Transition cations: - Colourless - Coloured	Zn^{2+} Fe^{2+} , Fe^{3+} and Cu^{2+}
Others	Ag^+ and NH_4^+

The reagents used to test for the cations are:-

- Sodium hydroxide (NaOH) solution.
- Ammonium hydroxide (NH₄OH) solution.

(a) Solution to Solution Reaction

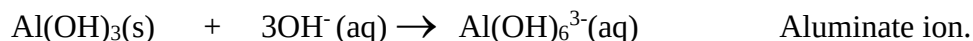
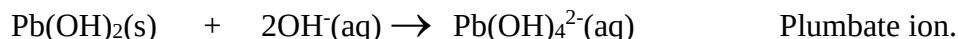
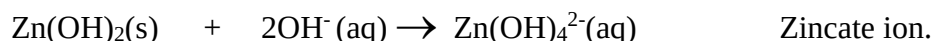
(i) Action of sodium hydroxide solution.

To the solution of the unknown compound, add sodium hydroxide solution drop wise until in excess. If no observable change, warm gently for ammonia.

Summary of Observations and Deductions

OBSERVATION	DEDUCTION
No observable change and no ammonia gas given off.	Na^+ , K^+ , Ag^+ ions probably present.
The solution remains colourless. A gas that turns red moist litmus paper blue is given off on warming.	NH_4^+ ions probably present.
White precipitate soluble in excess sodium hydroxide forming a colourless solution.	Zn^{2+} , Pb^{2+} and Al^{3+} ions probably present
White precipitate insoluble in excess sodium hydroxide solution.	Ca^{2+} , Mg^{2+} ions. probably present.
Blue precipitate insoluble in excess sodium hydroxide solution.	Cu^{2+} ions probably present.
Green precipitate insoluble in excess sodium hydroxide solution.	Fe^{2+} ions probably present.
Reddish brown precipitate insoluble in excess sodium hydroxide solution.	Fe^{3+} ions probably present

NB: The hydroxides of Pb^{2+} , Zn^{2+} and Al^{3+} are amphoteric and dissolve in excess sodium hydroxide solution to give their respective complexes.



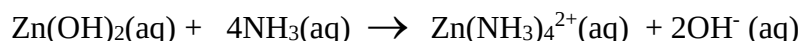
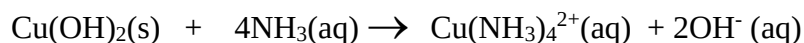
(ii) Action of Ammonia solution

Add ammonia solution to the solution of the unknown compound dropwise until in excess.

Summary of Observations and Deductions

OBSERVATION	DEDUCTION
No observable change.	Na^+ , K^+ , Ag^+ ions probably present.
White precipitate soluble in excess ammonia solution.	Zn^{2+} ions probably present.
White precipitate insoluble in excess ammonia solution.	Pb^{2+} , Ca^{2+} , Mg^{2+} , & Al^{3+} ions probably present.
Pale blue precipitate soluble in excess ammonia solution forming a deep blue solution.	Cu^{2+} ions confirmed.
Green precipitate insoluble in excess ammonia solution	Fe^{2+} ions probably present.
Reddish brown precipitate insoluble in excess ammonia solution	Fe^{3+} ions probably present.

NB: Copper (II) hydroxide and zinc hydroxide dissolve in excess ammonia solution to form their respective complexes, tetra amino copper and tetra amino zinc.



CONFIRMATORY TESTS FOR COMMON CATIONS

CATION	TEST	OBSERVATION
Na^+ & K^+	(i) Add sodium hydroxide solution and warm gently.	No observable change
	(ii) Add ammonia solution.	No observable change
NH_4^+	Add sodium hydroxide solution and warm.	A gas that turns red moist litmus paper blue is given off or smell of ammonia.
Zn^{2+}	(i) Add sodium hydroxide solution dropwise until in excess.	White precipitate forms. Precipitate soluble in excess sodium hydroxide.
	(ii) Add ammonia solution dropwise until in excess.	White precipitate forms. Precipitate soluble in excess ammonia solution.
Pb^{2+}	(i) Add dilute sulphuric acid.	White precipitate forms. $\text{Pb}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{PbSO}_4(\text{s})$
	(ii) Add conc. Hydrochloric acid. Heat the mixture and allow to cool.	White precipitate forms. Precipitate dissolves on warming and reappears on cooling.
	(iii) Add potassium iodide solution.	A yellow precipitate forms. $\text{Pb}^{2+}(\text{aq}) + 2\text{I}^-(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$
	(iv) Add Potassium Chromate (VI) solution.	A yellow precipitate forms..
Cu^{2+}	(i) Add ammonium hydroxide drop wise until in excess.	Deep blue solution formed.
Fe^{2+}	(i) Add Potassium hexacy-noferrate (II) solution.	Light blue or nearly white precipitate forms.
	(ii) Add Potassiumhexacy-noferrate (III) solution.	Dark blue precipitate forms.
	(iii) Add Potassium thiocyanate	Traces of iron (II) salt give colour.
Fe^{3+}	(i) Add Potassium hexacy-noferrate (II) solution.	Dark blue precipitate forms.
	(ii) Add Potassium hexacy-noferrate (III).	Dark solution forms.
	(iii) Add Potassium thiocynate.	Deep red solution forms.

(a) Flame Test

Place a little of the substance on watch glass and moisten it with pure conc. Hydrochloric acid. Heat a little of the mixture on clean platinum or nichrome wire and note the colour of the flame.

OBSERVATION	DEDUCTION
Brilliant yellow	Sodium
Lilac	Potassium
Red	Calcium
Blue	Lead
Blue- green	Copper

METHOD OF WRITING RESULTS

A sample of a substance **W** containing two cations and two anions was analyzed and the observations were recorded in the table below.

Complete the column for the deductions.

TESTS	OBSERVATIONS	DEDUCTIONS
1. Add 10 cm ³ of distilled water to two spatula endfuls of W . Shake well and filter. Keep both the filtrate and residue.	- Salt partial soluble. - Forming a Colourless Filtrate - Leaving white residue.	
2. Divide the filtrate into 3 parts and treat each part as follows:- (i) Add sodium hydroxide solution dropwise until in excess.	- White ppt formed. - Precipitate soluble in excess sodium hydroxide forming - a colourless solution.	
(ii) Add ammonia solution dropwise until in excess.	- White precipitate formed. -Precipitate soluble in excess ammonia solution forming - a colourless solution.	
(iii) Add barium chloride solution followed by dilute hydrochloric acid.	- White precipitate formed. - Precipitate insoluble in the acid.	

3. Wash the residue with distilled water then add 4 cm ³ of dilute nitric acid. Divide the solution into 2 parts.	- Residue dissolved with effervescence giving off a - a colourless gas that turns limewater milky . - A colourless solution was formed.	
(i) To the first part add dilute hydrochloric acid and heat gently to the boiling then allow to cool.	- White precipitate formed. - Precipitate soluble on heating. - The precipitate reformed on cooling.	
(ii) To the second part add potassium iodide solution.	- Bright yellow precipitate was formed	

The cations in **W** are: (i)

(ii)

The anions in **W** are: (i)

(ii)

SELF-CHECK 8.1

1. (2005 Q16)

Which one of the following reagents can be used to distinguish between Zn²⁺_(aq) and Al³⁺_(aq)?

- A. Lead (II) nitrate
C. Potassium iodine.

- B. Sodium hydroxide.
D. Ammonia solution.

2. (2005 Q.29)

The hydroxide that dissolves in excess aqueous ammonia but does **not** in sodium hydroxide solution is

- A. Lead (II) hydroxide.
C. Aluminium hydroxide.

- B. Zinc hydroxide.
D. Copper (II) hydroxide.

3. (2005 Q.35)

The cation that forms a green precipitate with sodium hydroxide is

- A. Cu²⁺_(aq).
C. Al³⁺_(aq).

- B. Fe²⁺_(aq).
D. Fe³⁺_(aq).

4. (2005 Q.39)

The hydroxide which turns brown when exposed to air from the list below is

- A. copper (II) hydroxide.
B. iron (II) hydroxide.
C. lead (II) hydroxide.
D. iron (III) hydroxide.

5. (2004 Q.15)

Which one of the following ions reacts with $\text{NH}_4(\text{aq})$ to form a precipitate that dissolves in excess ammonia solution?

- A. $\text{Pb}^{2+}_{(\text{aq})}$
C. $\text{Ca}^{2+}_{(\text{aq})}$

6. (2004 Q.19)

A white salt, X, reacts with dilute hydrochloric acid producing brown fumes. When dilute Sulphuric acid is added to the resulting solution, a white precipitate is observed. Salt X is

- A. Barium nitrate.
C. Potassium nitrate.
- B. Sodium nitrate.
D. Sodium bromide.

7. (2004 Q.33)

The solution that could be containing zinc ions is one that forms a

- A. reddish – brown precipitate with magnesium.
B. green precipitate with aqueous ammonia.
C. white precipitate that is soluble in excess sodium hydroxide solution.
D. white precipitate with dilute Sulphuric acid.

8. (2004 Q.37)

When sodium hydroxide solution was added to an aqueous solution of salt, X, a white precipitate insoluble in excess alkali was formed. X contained

- A. lead (II) ions.
B. magnesium ions.
C. zinc ions.
D. aluminium ions.

9. (2003Q.7)

Which one of the following solutions is used to test for sulphate ions in acid condition?

- A. Barium nitrate.
C. Silver nitrate.
- B. Lead (II) chloride.
D. Silver chloride.

10. (2003Q Q.24)

Which one of the following ions reacts with ammonia to form a precipitate which dissolves in excess ammonia to form a colourless solution?

- A. Zn^{2+} . B. Mg^{2+} . C. Cu^{2+} . D. Fe^{2+} .

11. (2002 Q.7)

Which one of the following anions does **not** form a precipitate with $\text{Pb}^{2+}_{(\text{aq})}$?

- A. $\text{CO}_3^{2-}(\text{aq})$. B. $\text{OH}^-(\text{aq})$. C. $\text{NO}_3^-(\text{aq})$. D. $\text{SO}_4^{2-}(\text{aq})$.

12. (2002 Q.8)

Which one of the following ions reacts with $\text{Cl}^-_{(\text{aq})}$ to form a precipitate which dissolves on heating?

- A. $\text{Cu}^{2+}_{(\text{aq})}$. B. $\text{Fe}^{2+}_{(\text{aq})}$. C. $\text{Pb}^{2+}_{(\text{aq})}$. D. $\text{Ca}^{2+}_{(\text{aq})}$.

13. (2002 Q.9)

A colourless solution reacts with dilute hydrochloric acid to give a white precipitate. The white precipitate dissolves in warm water. The colourless solution is

- A. zinc chloride.
C. magnesium sulphate.
- B. lead nitrate.
D. sodium carbonate.

14. (2002 Q.20)

Which one of the following will occur when aluminium is added to an aqueous solution of silver nitrate?

- A. Heat will be absorbed. B. Aluminium hydroxide will be formed.
C. Silver will precipitate out. D. Aluminium will form an alloy with silver.

15. (2001 Q.27)

A white precipitate was formed when an aqueous solution of a salt was reacted with aqueous barium nitrate. The white precipitate dissolved in nitric acid. The anion in the salt is

- A. SO_3^{2-} .
C. SO_4^{2-} .
- B. NO_3^- .
D. Cl^- .

16. (2001 Q.40)

Lead and silver nitrates are frequently used as reagents because

- A. they are the most suitable.
B. most nitrates of metal never decompose on heating.
C. other salts of lead and silver are insoluble.
D. other metal nitrates are insoluble.

17. (1999 Q.12)

Which one of the following salts when in solution will form a white precipitate with acidified barium nitrate solution?

- A. ZnSO_4 .
C. Na_2CO_3 .
- B. Na_2SO_3 .
D. ZnCl_2 .

18. (1999 Q.22)

Which one of the following ions can be confirmed by the brown ring test?

- A. Cl^- .
C. CO_3^{2-} .

19. (1999 Q.28)

Which one of the following oxides below is soluble in both excess sodium hydroxide solution and aqueous ammonia?

- A. Al_2O_3 .
C. PbO .
- B. ZnO .
D. Fe_2O_3 .

20. (1999 Q.36)

Which one of the following ions forms a green precipitate with excess sodium hydroxide?

- | | |
|-----------------------|-----------------------|
| A. Fe^{3+} . | B. Fe^{2+} . |
| C. Cu^{2+} . | D. Zn^{2+} . |

21. (1998 Q.12)

When dilute nitric acid followed by silver nitrate solution were added to a certain solution, a white precipitate which darkens when exposed to sunlight was formed. The solution contained a

- | | |
|--------------|---------------|
| A. sulphate. | B. chloride. |
| C. nitrate. | D. carbonate. |

22. (1998 Q.32)

Which one of the following ions when reacted with ammonia will form a blue precipitate that dissolves to give a deep blue solution?

- | | |
|---------------------|---------------------|
| A. Fe^{2+} | B. Cu^{2+} |
| C. Fe^{3+} | D. Zn^{2+} |

23. 1997 (Q.23)

Which one of the following mixtures would not form a precipitate?

- | | |
|--|--|
| A. Barium nitrate and sodium sulphate. | B. Lead nitrate and potassium iodine. |
| B. Copper nitrate and sodium sulphate. | D. Silver nitrate and potassium bromine. |

24. (1997 Q.26)

Which one of the following hydroxides will dissolve in excess ammonia?

- | | |
|-------------------------------|-------------------------------|
| A. $\text{Pb}(\text{OH})_2$. | B. $\text{Cu}(\text{OH})_2$. |
| C. $\text{Fe}(\text{OH})_2$. | D. $\text{Al}(\text{OH})_3$. |

25. (1995 Q.19)

When silver nitrate was added to a solution followed by dilute nitric acid, a white precipitate was formed. The solution contained

- | | |
|-------------------|--------------------|
| A. sulphate ions. | B. carbonate ions. |
| C. chloride ions. | D. sulphite ions. |

SELF-CHECK 8.2

QUALITATIVE ANALYSIS

1. (2004 Q.3)

A mixture containing copper (II) sulphate and copper (II) carbonate was shaken with excess water and filtered.

- (a) Identify the residue.
- (b) The dry residue was heated strongly.
 - (i) State what was observed.
 - (ii) Write an equation for the reaction.
- (c)
 - (i) Name a reagent that can be used to identify the anion in the filtrate.
 - (ii) Write an ionic equation for the anion and the reagent you have named in (c)(i).

2. (2004 Q.8)

The table below shows some tests that were carried out on a certain substance, Z, and the observations made.

<i>Test</i>	<i>Observation</i>
1. Solid Z was heated.	A colourless gas evolved which turned limewater milky.
2. Aqueous sodium hydroxide was added to aqueous solution of Z.	No apparent change.
3. Dilute hydrochloric acid was added to a solution of Z.	Effervescence and a gas that turned limewater milky evolved.
4. (i) Aqueous magnesium chloride was added to a solution of Z.	No apparent change.
(ii) Resultant solution from 4(i) was heated.	White precipitate formed.

- (a) What deduction can you make concerning the solubility of the hydroxide of the metal ion in Z?
- (b) State the: (i) likely anions present in Z.
(ii) anion actually present in Z.
- (c) (i) Identify the white precipitate in test 4(ii).
(ii) Write an equation to show the reaction resulting in the formation of the substance you have identified in (c)(i).

3. (2002 Q.10)

State **one** reagent that can be used to distinguish between each of the following pairs of ions and in each case, state what would be observed if each ion is treated with the reagent.

- (a) $\text{Pb}^{2+}_{(\text{aq})}$ and $\text{Al}^{3+}_{(\text{aq})}$.
- (b) $\text{SO}_4^{2-}_{(\text{aq})}$ and $\text{CO}_3^{2-}_{(\text{aq})}$.

4. (2000 Q.14)

Iron forms compounds in which it shows a valency of two and three.

- (a) State the general colour of iron compounds in which iron is:
(i) divalent.
(ii) trivalent.
- (b) Write a formula and the name of the sulphates of iron in which iron is:
(i) divalent.
(ii) trivalent.
- (c) (i) Name a reagent that can be used to distinguish between the sulphates in (b)(i) and (ii).
(ii) State what would be observed if each of the iron sulphate is reacted with the reagent you have named in (c)(i).
(iii) Write equations for the reaction in (c)(ii).

- (d) Starting from iron wool, state how the anhydrous chloride of iron (II) can be prepared and write equation to illustrate your answer. (Diagrams not required).

5. (1999 Q. 9)

The table below shows the results of tests carried out on an aqueous solution of a salt. Study the table and answer the questions that follow.

Tests	Observation
(i) With dilute sodium hydroxide solution.	White precipitate in excess alkali.
(ii) With aqueous ammonia.	A white precipitate insoluble in excess ammonia.
(iii) With aqueous potassium iodine.	A bright yellow precipitate.

- (i) Identify the cation.
(ii) Write an ionic equation for the reaction in test (ii) in the table.
(iii) Explain the observation in test (ii).

6. (1992 Q.12)

- (a) Name one reagent that can be used to distinguish between each of the following pairs of species.

In each case state what would be observed if each member of the pair is treated with the reagent and write equation for the reaction that takes place.

- (i) Lead (II) ions and zinc ions.
(ii) Carbonate ions and chloride ions.

CHAPTER NINE

OXIDATION AND REDUCTION

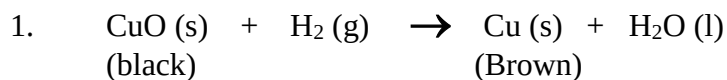
(a) Oxidation as the addition of oxygen or removal of hydrogen

(i) Oxidation is the addition of oxygen to a substance or the removal of hydrogen from a substance.

An oxidizing agent is a substance which transfers oxygen to another substance, or removes hydrogen from that substance.

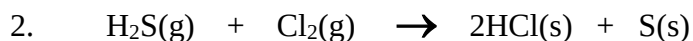
Examples

Consider the following reactions:



In this reaction, the hydrogen has been oxidized to water (since oxygen has been added to it).

The copper (II) oxide is described as an oxidizing agent.



In this reaction although oxygen does not appear in any of the substances involved in the reaction. The reaction is still called **REDOX** reaction. This is because the chlorine molecule removes hydrogen from the hydrogen sulphide (H_2S). The oxidizing agent in this case is chlorine.

(ii) Reduction

Reduction is the addition of hydrogen to a substance or the removal of oxygen from a substance.

A reducing agent is a substance which transfers hydrogen to another substance, or removes oxygen from another substance.

Examples

Consider examples 1 and 2 above.

In example (1), the copper (II) oxide has been reduced to copper (since oxygen has been removed from it). The hydrogen that caused the reduction is called a reducing agent.

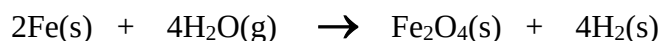
In example (2), the hydrogen sulphide (H_2S) has added hydrogen to chlorine, therefore, reduction has occurred. Thus, the hydrogen sulphide is called a reducing agent.

Note that:

*Reduction and oxidation occur simultaneously in a reaction; therefore, in short, such a reaction is called **REDOX** reaction.*

Exercise

Identify the oxidizing agent and the reducing agent in the following reaction.



Answer

Oxidising agent: Water (H_2O)

Reducing agent: Iron (Fe)

(b) Oxidation as the addition of electronegative element

(i) Oxidation

Oxidation is *the addition of an electronegative element to a substance*

Or *the removal of an electropositive element from a substance.*

(ii) Reduction

Reduction is *the addition of electropositive element to a substance.*
or *the removal of an electronegative element from a substance.*

NB: *Electronegative elements are elements that tend to attract electrons to themselves (to become ultimately, negative ions).*

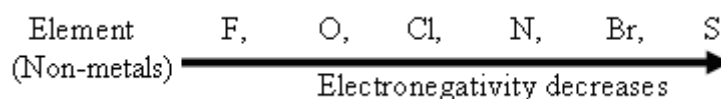
E.g. Non-metallic elements (F, O, C, N, Br, S).

Electropositive elements are elements that tend to donate electrons to other substances (to become ultimately, positive ions).

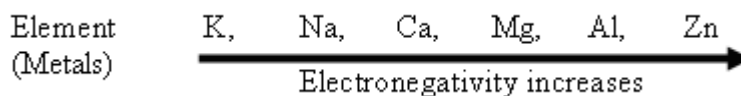
E.g. Metallic elements (K, Na, Ca, Mg, Al, Zn)

The electro-negativity of some elements

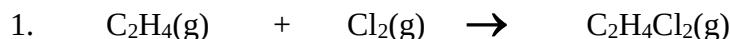
(i) Non-metals



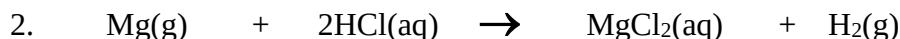
(ii) Metals



Examples



The ethane (C_2H_4) has been oxidized because chlorine (electronegative) has been added to it.



Magnesium is oxidized by addition of chlorine and hydrochloric acid (HCl) is reduced by removing chlorine from it. Therefore, magnesium is a reducing agent, while the hydrogen chloride is an oxidizing agent.



Both sulphur and chlorine are electronegative elements, but because chlorine is much more electronegative than sulphur, then it is regarded as the oxidizing agent, while the sulphur is reducing agent.

(c) Oxidation and Reduction in terms of a change in oxidation number

(i) Oxidation

Oxidation is an increase in oxidation number.

(ii) Reduction

Reduction is the decrease in oxidation number.

Simple rules concerning the idea of oxidation number

- (i) All elements in the free state (i.e. uncombined with any other elements) have oxidation number of zero.
- (ii) In the case of a simple ion the element has oxidation number with the same size and sign of the charge on the ion.

Ion	Oxidation number/sign
Cu^{2+}	+2
Fe^{2+}	+2
Fe^{3+}	+3
S^{2-}	-2
O^{2-}	-2

2. The sum of all the oxidation numbers of the elements in a compound is zero.

The rules enable us to work out the oxidation number of an element of which we are uncertain.

Example

Work out the oxidation number of sulphur in the compound FeSO_4 .

$$\begin{array}{ll}\text{Let the oxidation number of S} & = x \\ \text{Oxidation number of oxygen} & = -2 \\ \text{Oxidation number of iron (II)} & = +2\end{array}$$

$$+2 + x + (4 \times -2) = 0$$

$$+2 + x - 8 = 0$$

$$x - 6 = 0$$

$$x = +6$$

$$\therefore \text{The oxidation number of S} = +6$$

With these simple rules in mind the concept of oxidation and reduction may be redefined.

Oxidation: when oxidation occurs the oxidation number of the element increases.

Reduction: when reduction occurs the oxidation number of the element decreases.

Exercise

Find the oxidation number of Fe and Ca in the following compounds.

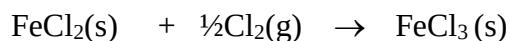
- (i) FeCl_3 (ii) CaCl_2 (Given that oxidation number of Cl = -1)

Answers

- (i) +3 (ii) +2

Example

Now consider the reaction below.



The oxidation number of Fe changes from +2 in FeCl_2 to +3 in FeCl_3 and therefore, FeCl_2 has been oxidized to FeCl_3 by chlorine, which is therefore the oxidizing agent.

In the similar way the oxidation number of chlorine changes from 0 in the Free State to -1 in FeCl_3 the chlorine has therefore been reduced by the FeCl_2 .

(d) Oxidation and Reduction in terms of electron transfer

From the electronic point of view the following are the most recent definitions of *oxidation* and *reduction*.

- (i) *Oxidation is the loss of electrons*
- (ii) *Reduction is the gain of electrons*

An oxidizing agent is an electron acceptor.

A reducing agent is an electron donor.

Memory aid:

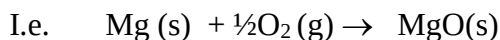
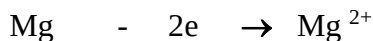
O	<i>Oxidation</i>
I	<i>Is</i>
L	<i>Loss of electrons</i>
R	<i>Reduction</i>
I	<i>Is</i>
G	<i>Gain of electrons</i>

- NB:**
- (i) *Oxidation and reduction always occur simultaneously. They are complementary processes of electron loss and electron gain respectively.*
 - (ii) *The electrons lost by the reducing agent (electron donor) must be accepted by the oxidizing agent (electron acceptor) present.*

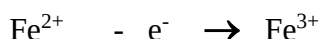
Examples

1. *When magnesium is oxidized by combination with oxygen, the metal is oxidized by losing two electrons per atom.*

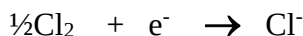
These electrons are accepted by oxygen atoms, which are reduced as a result: Magnesium (giving out electrons) is the reducing agent; while oxygen (accepting electrons) is the oxidizing agent.



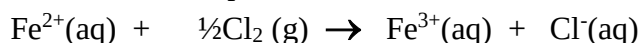
2. *If a metallic ion, e.g. the iron (II) ion, Fe^{2+} , is so treated that it loses a further electron, it is oxidized and is a reducing agent. The process is:*



An iron (III) ion is formed. An agent must be present, e.g., chlorine, to accept the electrons made available by the ferrous ions. It acts as the oxidizing agent (electron acceptor) and is reduced.



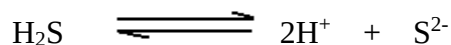
The complete reaction can be represented:



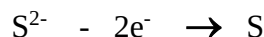
It will be observed that ‘valency’ of the metal increases from 2 to 3 during the oxidation.

3. *The ‘removal of hydrogen’ aspect of oxidation is interpreted in the following way.*

Consider the oxidation of hydrogen sulphide by chlorine. Hydrogen sulphide is slightly ionized as:



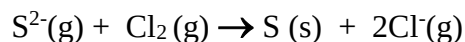
The sulphide ion parts with its two electrons and is, therefore, oxidized,, acting as a reducing agent.



The electrons are accepted by chlorine atoms, so that chlorine acts as an oxidizing agent and is reduced.

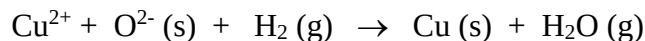


Adding the two equations, we have:

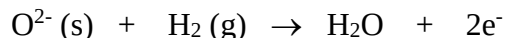


The hydrogen ion of the hydrogen sulphide is unchanged.

4. *The reduction of hot copper (II) oxide by hydrogen is given as:*



It is clear that the copper (II) ion is reduced by the reaction between the oxide ion O^{2-} and hydrogen:



By combining with oxygen in this way and supplying electrons to the metallic ion, hydrogen exercises reducing properties. The oxide ion is oxidized by electron loss and the oxygen atom remains in combination with hydrogen as water.

SELF-CHECK 9.1

1. Which one of the following processes is not an example of oxidation?

A. The burning of methane in air.

C. The melting of a candle wax.

B. The rusting of iron.

D. The smouldering of phosphorus.

2. Which one of the following equations does **not** represent oxidation of the first substance?

A. $\text{Fe}_2\text{O}_{3(s)} + 3\text{CO}_{(g)} \rightarrow 2\text{Fe}_{(s)} + 3\text{CO}_{2(g)}$.

B. $2\text{H}_2\text{S}_{(aq)} + \text{O}_{2(g)} \rightarrow 2\text{S}_{(s)} + 2\text{H}_2\text{O}_{(l)}$.

C. $\text{S}_{(s)} + \text{O}_{2(g)} \rightarrow \text{SO}_{2(g)}$.

D. $\text{Zn}_{(s)} + \text{Cu}^{2+}_{(aq)} \rightarrow \text{Zn}^{2+}_{(aq)} + \text{Cu}_{(s)}$.

3. Which one of the following equations does not represent reduction reaction?

A. $\text{N}_{2(g)} + 3\text{H}_{2(g)} \rightarrow 2\text{NH}_{3(g)}$.

C. $\text{Cl}_{2(g)} + 2\text{e}^- \rightarrow 2\text{Cl}^{-}_{(aq)}$.

B. $\text{Fe}_{(s)} \rightarrow \text{Fe}^{3+}_{(aq)} + 3\text{e}^-$.

D. $\text{Cu}^{2+}_{(aq)} + 2\text{e}^- \rightarrow \text{Cu}_{(s)}$.

4. Which **one** of the following equations represents an oxidation – reduction reaction?

A. $\text{H}^{+}_{(aq)} + \text{OH}^{-}_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)}$.

B. $\text{Pb}^{2+}_{(aq)} + \text{SO}_4^{2-}_{(aq)} \rightarrow \text{PbSO}_{4(s)}$.

C. $\text{Cu}^{2+}_{(aq)} + \text{Fe}_{(s)} \rightarrow \text{Cu}_{(s)} + \text{Fe}^{2+}_{(aq)}$.

D. $\text{Ca}^{2+}_{(aq)} + \text{CO}_3^{2-}_{(aq)} \rightarrow \text{CaCO}_{3(s)}$.

5. Which one of the following represents a reduction-oxidation reaction?

A. $2\text{NaOH}_{(aq)} + \text{CuCl}_{2(aq)} \rightarrow \text{Cu}(\text{OH})_{2(s)} + 2\text{NaCl}_{(aq)}$.

B. $2\text{FeCl}_{2(aq)} + \text{Cl}_{2(g)} \rightarrow 2\text{FeCl}_{3(aq)}$.

C. $2\text{NaOH}_{(aq)} + \text{H}_2\text{SO}_{4(aq)} \rightarrow \text{Na}_2\text{SO}_{4(aq)} + 2\text{H}_2\text{O}_{(l)}$.

D. $\text{ZnCO}_{3(s)} + 2\text{HNO}_{3(aq)} \rightarrow \text{Zn}(\text{NO}_3)_{2(aq)} + \text{H}_2\text{O}_{(l)} + \text{CO}_{2(g)}$.

6. Which one of the **underlined** substances in the equation below is being reduced?

A. CuO_(s) + H₂(g) → Cu(s) + H₂O(l)

B. 2Fe²⁺_(aq) + Cl₂(g) → 2Fe³⁺_(aq) + 2Cl⁻_(aq)

C. MnO₂(s) + 4HCl_(aq) → MnCl₂(aq) + 2H₂O(l) + Cl₂(g)

D. H₂SO₄_(l) + C_(s) → CO₂(g) + 2SO₂(g) + 2H₂O(l)

In each questions 7 to 8 one or more of the answers given may be correct. Read each question carefully and then indicate the correct answer A, B, C or D according to the following:

A. if 1, 2, 3 only are correct.

B. if 1, 3 only are correct

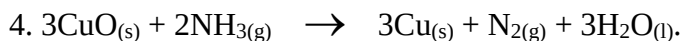
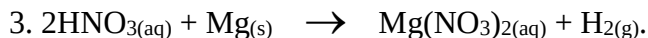
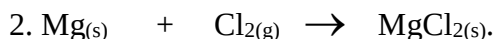
C. if 2, 4 only are correct.

D. if 4 only is correct.

INSTRUCTIONS SUMMARISED:

Instruction Summarised			
A	B	C	D
1, 2, 3	1, 3	2, 4	4
only correct	only correct	only correct	only correct

7. Which of the following equation(s) show(s) a redox reaction?



8. Oxidation is a reaction in which

1. oxygen is removed from a substance.
2. hydrogen is removed from a substance.
3. hydrogen is added to a substance.
4. electron is lost from a substance.

SELF-CHECK 9.2

1. (a) Define **oxidation** in terms of electrons.

(b) Write the equations of the following half reactions and in each case state whether the reaction is a reduction or an oxidation reaction.

(i) The conversion of hydrogen ions (H^+) to hydrogen molecules (H_2).

(ii) The conversion of iron (II) ions (Fe^{2+}) to iron (III) ions (Fe^{3+}).

(iii) The conversion of chlorine molecules (Cl_2) to chloride ions (2Cl^-).

(c) Which elements in the second and third periods of the periodic table are strong

(i) Oxidizing agents?

(ii) Reducing agents?

(iii) What properties do strong oxidizing agents have?

(iv) What properties do strong reducing agents have?

CHAPTER TEN

ENERGY CHANGES IN CHEMICAL REACTIONS

Introduction:

Energy is either given out to or absorbed from the surrounding whenever a chemical reaction takes place.

The change in energy is the driving force for a chemical reaction.

There are several forms of energy. The most common form of energy in chemical reaction is the heat energy also called *enthalpy*.

Enthalpy and enthalpy changes

(i) Enthalpy

Enthalpy is the energy (heat) content of a substance which is stored in its bonds.

Its symbol is H.

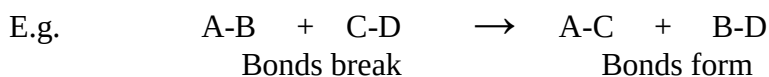
(ii) Enthalpy change

When a chemical reaction takes place, the enthalpy of the reactants change.

The change in enthalpy is denoted by ΔH .

Bond breaking and Bond formation

When a chemical reaction occurs, bonds in the reactants break and new bonds are formed in the products.



Bond between **AB** and **CD** break and new bonds form between **AC** and **BD**.

During bond breaking, energy is absorbed from the surrounding while during bond formation, energy is released to the surrounding.

It is the energy difference in bond breaking and bond formation that we measure. This can be shown as a temperature rise or fall in the thermometer.

The Joule (J)

The unit of energy is the joule, J.

Definition:

The Joule is the energy that raises the temperature of 1g of water through 0.24 K.

The larger unit is Kilo Joule, KJ. 1KJ = 1000 J

Enthalpy changes in chemical reactions refer to one mole of a particular reactant or product in a balanced equation.

The enthalpy changes involved are normally large. Therefore, KJ is preferred to joule which is too small.

Types of Chemical Reactions

There are two types of reactions, namely:

- (i) **Exothermic reaction**
- (ii) **Endothermic reaction**

(a) Exothermic Reaction

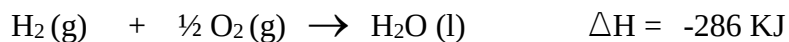
Exothermic reaction is the reaction in which heat energy is given out to the surroundings.

In exothermic reaction there is temperature rise in the thermometer and the temperature rise is a measure of this energy.

The energy of the products is less than the energy of the reactants.

∴ The enthalpy change, ΔH has a negative sign.

Examples of exothermic reaction include:



(b) Endothermic Reaction

Endothermic Reaction is a reaction in which heat energy is absorbed from the surroundings.

There is temperature drop in the thermometer and the temperature drop is a measure of this energy.

The energy of the products is more than the energy of the reactants.

Therefore, ΔH is positive.

Examples of endothermic reaction



$$\text{Enthalpy change} = \left(\text{Enthalpy of products} \right) - \left(\text{Enthalpy of reactants} \right)$$

$$\Delta H = H (\text{Products}) - H (\text{reactants})$$

Types of Enthalpy change

The heat change/enthalpy change which occurs in a reaction is named after the type of reaction in which it occurs. They include the following:-

- (i) Enthalpy of combustion (Heat of combustion).
- (ii) Enthalpy of formation (Heat of formation).
- (iii) Enthalpy of neutralization (Heat of Neutralization).
- (iv) Enthalpy of solution (Heat of solution).

(a) Enthalpy of Combustion

Enthalpy of Combustion is the enthalpy change that occurs when one mole of a substance is completely burnt in (pure) oxygen.

Or *is the heat liberated when one mole of a substance is completely burnt in oxygen.*

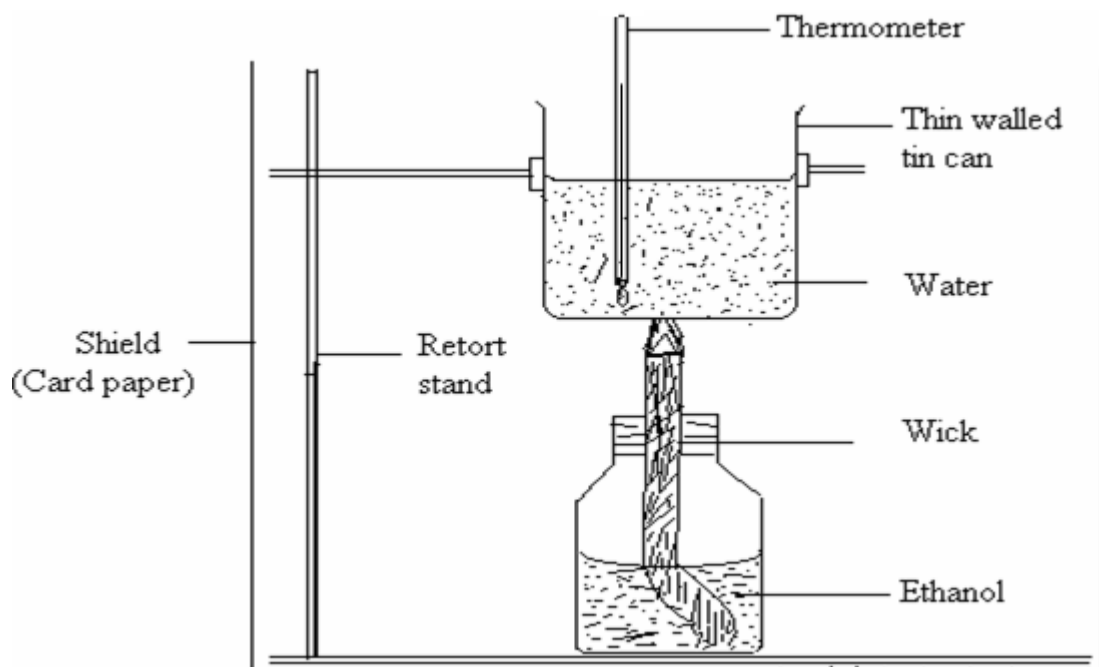
Determination of Enthalpy of Combustion of Liquid eg ethanol

Apparatus:

Retort stand/clamp, a thin walled tin can, thermometer, water, ethanol, and a spirit lamp.

Procedure:

- Weigh a spirit burner containing pure ethanol.
- Fill a thin-walled tin can with a known volume of water.
- Clamp the tin can with its content above the lamp as shown in the diagram below.
- Insert the thermometer into the water.
- Read and record the initial temperature of the water.



- Heat the water while carefully stirring with the thermometer.
- When the thermometer shows a convenient temperature rise by 25 °C, put out the flame and read and record the final temperature.
- Reweigh the can and its content.

Results:

Volume of cold water	=	V_1	cm^3
Initial temperature of cold water	=	t_1	$^{\circ}\text{C}$
Final temperature of water	=	t_2	$^{\circ}\text{C}$
Mass of spirit burner before burning	=	m_1	g
Mass of spirit burner after burning	=	m_2	g
Specific heat capacity of water	=	4.2	$\text{J/g } ^{\circ}\text{C}$

Calculation

$$\begin{aligned}
 \text{Heat gained by water} &= mc\Delta t \\
 &= \left(\text{Mass of water} \right) \times \left(\text{Sp. ht. cap of water} \right) \times \left(\text{Change in Temperature} \right) \\
 &= (\text{Vol.} \times \text{Density}) \times 4.2 \times (t_2 - t_1) \\
 &= (V_1 \times 1 \times 4.2 \times (t_2 - t_1)) \\
 &= (V_1 \times 4.2 \times (t_2 - t_1)) \text{ Joules}
 \end{aligned}$$

Heat produced/lost by ethanol is calculated as follows:

The relative molecular mass of ethanol, C_2H_5OH ,

$$= (2 \times 12) + (1 \times 6) + 16$$

$$= 24 + 6 + 16$$

$$= \mathbf{46 \text{ g}}$$

$$\begin{aligned} \text{Mass of ethanol used up (burnt)} &= \left(\text{Mass of spirit lamp before burning} \right) - \left(\text{Mass of spirit lamp after burning} \right) \\ &= m_1 - m_2 \end{aligned}$$

Assuming heat gained by water = Heat given out by ethanol
 $\therefore (m_1 - m_2)$ g of ethanol produce, $V_1 \times 4.2 \times (t_2 - t_1)$ joules of energy

$$46 \text{ g of ethanol produce} = \frac{46}{m_1 - m_2} \times (t_2 - t_1) \times (V_1 \times 4.2) \quad \text{KJ mol}^{-1}$$

Or in terms of moles

$$\begin{aligned} \text{Moles of ethanol used up} &= \frac{\text{Mass of ethanol used}}{\text{Relative molecular mass}} \\ &= \frac{m_1 - m_2}{46} \end{aligned}$$

If $\frac{m_1 - m_2}{46}$ moles of ethanol produce $V_1 \times 4.2 \times (t_2 - t_1)$ joules of KJ energy

Then 1 mole of ethanol produces $\left(\frac{46}{m_1 - m_2} \right) \times (V_1 \times 4.2) \times (t_2 - t_1) \quad \text{KJ mol}^{-1}$

This is the heat of combustion of ethanol.

Example:

An S.4 student in TLA carried out an experiment to determine the enthalpy of combustion of ethanol in the laboratory and obtained the following results.

$$\text{Volume of water} = 120 \text{ cm}^3$$

$$\text{Initial temperature of water} = 20^\circ \text{C}$$

$$\text{Final temperature of water} = 46^\circ \text{C}$$

$$\text{Mass of the spirit burner before} = 33.0 \text{ g}$$

$$\text{Mass of the spirit burner after} = 32.5 \text{ g}$$

- (a) Using the above results calculate the heat of combustion of ethanol.
 (b) State any precautions that could be taken to obtain accurate result.
 (specific heat capacity of water = $4.2 \text{ g}^{-1}\text{K}^{-1}$, and $= 1 \text{ g/cm}^3$)

Solution

$$\begin{aligned}
 \text{(a) Heat energy gained by water} &= mc\Delta t \\
 &= \left(\text{Mass of water} \right) \times \left(\text{Sp. ht cap of water} \right) \times \left(\text{Change in Temperature} \right) \\
 &= \left(\text{Vol.} \times \text{Density} \right) \times \left(\text{Sp. ht cap of water} \right) \times \left(\text{Change in Temperature} \right) \\
 &= 120 \times 1 \times 4.2 \times (t_2 - t_1) \\
 &= 120 \times 4.2 \times (46 - 20) \\
 &= 120 \times 4.2 \times 26 \\
 &= \mathbf{13104 \text{ J}}
 \end{aligned}$$

$$\begin{aligned}
 \text{Mass of ethanol burned} &= \left(\text{Mass of spirit lamp before burning} \right) - \left(\text{Mass of spirit lamp after burning} \right) \\
 &= 33 - 32.5 \\
 &= \mathbf{0.5 \text{ g}}
 \end{aligned}$$

$$\begin{aligned}
 \text{Formula mass of } \text{C}_2\text{H}_5\text{OH}, &= 2 \times 12 + 1 \times 6 + 16 \\
 &= \mathbf{46 \text{ g}}
 \end{aligned}$$

0.5 g of ethanol liberates 13104 J of heat energy.

$$1 \text{ g of ethanol liberates } \frac{13104}{0.5} \text{ J of heat energy.}$$

$$\begin{aligned}
 46 \text{ g of ethanol will liberate } \frac{13104}{0.5} \times 46 &= 1205568 \text{ J} \\
 &= \mathbf{1205.568 \text{ KJ mol}^{-1}}
 \end{aligned}$$

$$\therefore \text{The enthalpy of combustion of ethanol} = \mathbf{-1205.568 \text{ KJ mol}^{-1}}$$

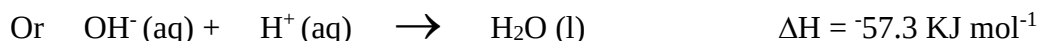
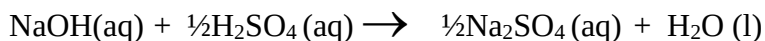
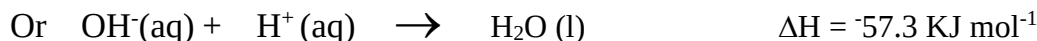
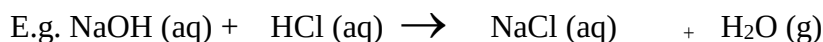
- (b) This experiment gives less value because some heat was lost to the surroundings and in heating the can itself.

The flame should be shielded from wind by fixing a hard cardboard around the setup.

Enthalpy of neutralization

Enthalpy of neutralization is the enthalpy change that occurs when an acid reacts with an alkali to produce a salt and a mole of water.

Or *is the enthalpy change that occurs when one mole of aqueous hydrogen ions reacts with one mole of aqueous hydroxide ions to form 1 mole of water.*



NB:

- ❖ The basicity of H_2SO_4 is two i.e. one mole of H_2SO_4 supplies two moles of H^+ ions.
- ❖ Therefore to get one mole of H^+ ions half a mole of H_2SO_4 is required.
- ❖ The enthalpy of neutralization of any strong acid and alkali is -57.3 KJmol^{-1} .
- ❖ Strong acids and alkalis dissociate completely into ions when dissolved in water.
- ❖ For weak acid and alkali, the enthalpy of neutralization is less than $-57.3 \text{ KJ mol}^{-1}$.
- ❖ This is because the acid or alkali is slightly dissolved into ions when in aqueous solution. Some heat is absorbed in order for complete dissociation to take place.

Determination of the enthalpy of neutralization

Experiment

To determine the enthalpy of neutralization of sodium hydroxide by hydrochloric acid

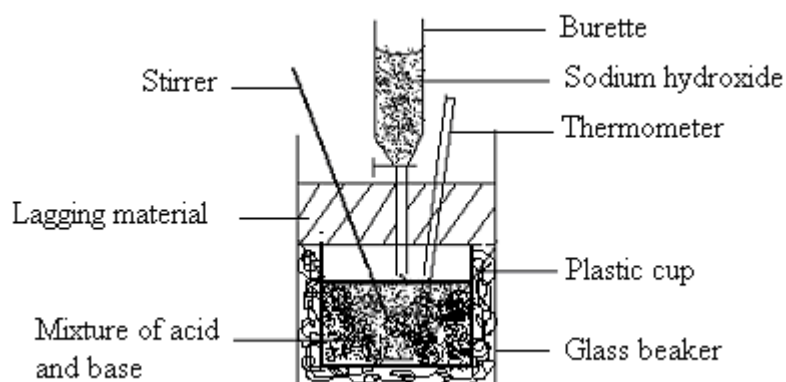
Apparatus:

Thermometer, a plastic cup or beaker, stirrer, lagging material.

Procedure:

- Place hydrochloric acid of known volume and molarity into the plastic cup.
- Read and record the initial temperature t_1 , of the acid.
- Fill sodium hydroxide of the same volume and molarity as for the acid in the burette.
- Read and record the initial temperature t_2 , of the alkali.
- Determine the initial temperature = average temperature, $\frac{t_1 + t_2}{2}$, of the reactants.

- As rapidly as you can transfer the base solution to the plastic cup using burette and stir the mixture using the thermometer



- Read and record the maximum temperature attained.

Specimen results

Volume of 2M of hydrochloric acid solution	= 20.0 cm ³
Volume of 2M of sodium hydroxide solution	= 20.0 cm ³
Initial temperature hydrochloric acid acid	= 15.0 °C
Initial temperature of sodium hydroxide	= 15.4 °C
Final maximum temperature of mixture	= 28.2 °C
Density of water	= 1g/cm ³

Calculations:

- Hint:** (i) *In the calculations, it is assumed that the total volume of the reactants is equal to the volume of water formed when the acid reacts with the base.*
- (ii) *All the heat energy liberated is absorbed by the water formed.*
- (iii) *The mass of water formed obtained by multiplying the density of water and the volume.*

$$\begin{aligned}
 \text{Final volume of solution} &= \text{Vol. of acid} + \text{Vol. of base} \\
 &= (20 + 20) \text{ cm}^3 \\
 &= 40 \text{ cm}^3
 \end{aligned}$$

$$\begin{aligned}
 \text{Mass of solution} &= \text{Volume} \times \text{density} \\
 &= 40 \times 1 \\
 &= 40 \text{ g}
 \end{aligned}$$

$$\text{Initial temperature} = \text{average of initial temp. of acid base}$$

$$= \frac{15.0 + 15.4}{2}$$

$$= 15.2^{\circ}\text{C}$$

$$\text{Final temperature of mixture} = 28.2^{\circ}\text{C}$$

$$\begin{aligned}\text{Change in temperature, } \Delta t &= (\text{Final} - \text{Initial}) \text{ temperature} \\ &= 28.2 - 15.2 \\ &= \mathbf{13^{\circ}\text{C}}\end{aligned}$$

Given the specific heat capacity of water is 4.2 Jg^{-1} ,

$$\begin{aligned}\text{Heat evolved} &= \text{Gained by the water formed} \\ &= \text{Mass} \times \text{heat cap} \times \text{temperature rise} \\ &= 40 \times 4.2 \times 13.0 \text{ J}\end{aligned}$$

$$\therefore \text{Heat evolved} = \mathbf{2184 \text{ J}}$$

$$\begin{aligned}\text{Moles of hydrochloric acid} &= \text{Molarity} \times \text{volume in litres} \\ &= \frac{2 \times 20}{1000} \\ &= \mathbf{0.04 \text{ moles of acid}}\end{aligned}$$



And since one mole of hydrochloric acid produces one mole of water, 0.04 moles will produce 0.04 moles of water.

Hence 0.04 moles of water is produced with the evolution of 2184 J heat energy.

$$\begin{aligned}\therefore 1 \text{ mole of water will be produced with the evolution of } &\frac{2184 \times 1}{0.04} \\ &= 54600 \text{ J} \\ &= \mathbf{54.6 \text{ KJ mol}^{-1}}.\end{aligned}$$

Therefore the heat of neutralization is $\sim 54.6 \text{ KJ mol}^{-1}$ water formed.

NB: The value obtained is less than the actual value $\sim 57.3 \text{ KJ mol}^{-1}$ due to experimental errors.

Enthalpy change of solution

Enthalpy change of solution is the enthalpy change that occurs when one mole of a solute/compound completely dissolves in a stated amount of solvent/water.

NB: Heat of solution depends on the quantity of water used.

Determination of heat of solution

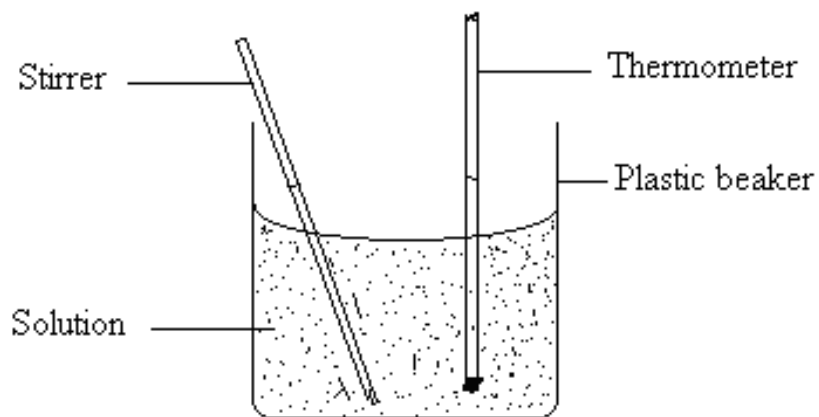
Experiment: To determine heat of solution of:

1. Sulphuric acid

Add 53.5 cm³ of concentrated sulphuric acid (i.e 98 g or 1 mole) carefully to 900 cm³ of water in a large, thin glass or polythene beaker.

Stir well and record the rise in temperature (17°C)

(The final solution is about M concentration, its mass is about 1000 g and its specific heat capacity is about 4.2 Jg⁰C).



$$\begin{aligned}\text{Heat of solution} &= \text{Mass} \times \text{specific heat capacity} \times \text{temperature rise} \\ &= 1000 \times 4.2 \times 17 \\ &= 71,400 \\ &= 71.4 \text{ KJ mol}^{-1}\end{aligned}$$



2. **Heat of solution of Sodium hydroxide**

Add 40 g sodium hydroxide (1 mol) pellets to 960 cm³ of water in a beaker as in 1 above.

Stir well and record the temperature rise, θ .

Calculate the heat of solution of sodium hydroxide.

$$\begin{aligned}\text{Heat of solution} &= \text{Mass} \times \text{sp. heat capacity} \times \text{temperature rise} \\ &= 1000 \times 4.2 \times \theta \\ &= 4200 \times \theta \text{ KJ mol}^{-1}\end{aligned}$$

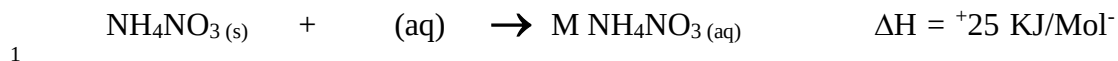
3. **Heat of solution Ammonium nitrate**

Add 80 g (1 mol) of the finely powdered Ammonium nitrate to 920 cm³ of water.

Record the temperature fall (θ °C).

The concentration of the final solution is M, its mass is 1000 g and its specific heat capacity is 4.2 J/g°C.

Use the above equation to calculate the heat of solution of the salt.

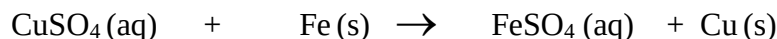


$$\begin{aligned}\text{Heat of solution} &= \text{Mass} \times \text{sp. heat capacity} \times \text{temperature rise} \\ &= 1000 \times 4.2 \times \theta \\ &= 4200 \times \theta \text{ KJ mol}^{-1}\end{aligned}$$

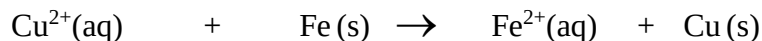
4. **Enthalpy of displacement of a metal ion in solution by another metal**

When iron filings are added to copper (II) sulphate solution, the blue solution turns pale green and a brown solid is deposited.

The pale green colour is due to iron (II) sulphate solution and the brown solid is copper.



Or ionically



In this reaction, iron displaces copper (II) ions.

The reaction is exothermic and a thermometer would show a temperature rise.

Example

Candidates of **TLA** carried out an experiment to determine the enthalpy change of displacement. They added 1.0 g of zinc powder to 50 cm³ of 0.2M copper (II) sulphate solution into a plastic beaker and stirred the mixture well. The temperature at the beginning of the experiment was 20 °C and it rose to a maximum of 27 °C.

- (a) Why was a plastic beaker used in this experiment?
- (b)
 - (i) State what was observed.
 - (ii) Write an ionic equation for the reaction which took place.
- (c) Calculate:
 - (i) The heat produced during the reaction.
 - (ii) The enthalpy change of the reaction.
 - (iii) Would the value of ΔH have been bigger or smaller if iron fillings were used instead of zinc?

Solution:

- (a) *Plastic beaker is a good heat insulator and absorbs hardly any of the heat produced in the reaction.*
- (b)
 - (i) *The blue colour of copper (II) sulphate solution turned to Colourless and A brown precipitate was formed.*
 - (ii)
$$\text{Cu}^{2+}(\text{aq}) + \text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$$
- (c)
 - (i) *Volume of solution = 50 cm³, sp. ht. cap. = 4.2 J/g/K, Density of water = 1 g/cm³.*

$$\begin{aligned}
 \text{Heat produced} &= mc\Delta\theta \\
 &= \text{Vol.} \times \rho \times c \times (t_2 - t_1) \\
 &= 50 \times 1 \times 4.2 \times (27 - 20) \\
 &= 50 \times 4.2 \times 7 \\
 &= 1470 \text{ J}
 \end{aligned}$$

$$\therefore \text{Heat produced} = 1.47 \text{ KJ}$$

- (ii) *1000 cm³ of copper (II) sulphate solution contains 0.2 moles.*

$$\begin{aligned}
 50 \text{ cm}^3 \text{ of copper (II) sulphate solution contains } &\frac{0.2 \times 50}{1000} \\
 &= 0.01 \text{ moles.}
 \end{aligned}$$

0.01 moles of copper liberates 1470 KJ of heat energy.

$$1 \text{ mole of copper will liberate } \frac{1 \times 1470}{0.01} = 147 \text{ KJ mol}^{-1}$$

- (iii) **Smaller:** *this is because iron is less reactive than zinc*

(i.e. iron is below zinc in the reactivity series) and would require more energy to break the bonds in copper (II) sulphate and in the formation of the new bonds.

Enthalpy of formation

Enthalpy of formation is the heat liberated or absorbed when 1 mole of substance is formed from its constituent elements under standard conditions.

Example

1. (a) Define the heat of formation
- (b) Iodine and chlorine react as follows



Calculate the heat of formation of iodine monochloride. From the equation; Formation of 2 moles of ICl liberates 68 KJ.

Solution

- (a) *Enthalpy of formation is the heat liberated or absorbed when 1 mole of substance is formed from its constituent elements under standard conditions.*

$$\begin{aligned} \text{(b)} \quad \text{The heat of formation of 1 mole of ICl} &= \frac{1 \times 68}{2} \\ &= 34 \text{ KJ mol}^{-1} \end{aligned}$$

SELF-CHECK 9.1

1. (2005 Q.36)

10 g of methanol, CH_3OH , burns in air to liberate 226 kJ of heat. The amount of heat liberated when 1 mole of methanol is burnt in air is [H = 1; C = 12]

A. $\frac{32 \times 226}{10}$

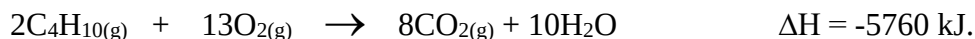
B. $\frac{10 \times 32}{226}$

C. $\frac{10}{32 \times 226}$

D. $\frac{10 \times 226}{32}$

2. (2005 Q.12)

Butane burns in excess air according to the following equation:



The quantity of heat evolved when 1.6 dm^3 of butane is burnt at room temperature [1 mole of gas occupies 24 dm^3 at room temperature]

$$\text{A. } \frac{5760 \times 116}{2 \times 24}$$

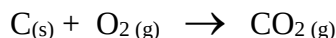
$$\text{B. } \frac{5760 \times 1.6}{2 \times 24}$$

$$\text{C. } \frac{5760 \times 1.6}{24}$$

$$\text{D. } \frac{5760 \times 116}{24}$$

3. **(2004 Q.5)**

Carbon burns in oxygen according to the following equation:



The amount of the heat evolved when 480g of carbon is burnt completely in oxygen is [Molar heat of combustion of carbon is $2.2 \times 10^{-7} \text{ kJ mol}^{-1}$, C = 12]

$$\text{A. } 480 \times 12 \times 2.2 \times 10^{-7}$$

$$\text{B. } \frac{480 \times 12}{2.2 \times 10^{-7}}$$

$$\text{C. } \frac{2.2 \times 10^{-7} \times 12}{480}$$

$$\text{D. } \frac{2.2 \times 10^{-7} \times 480}{12}$$

4. **(2003 Q.12)**

When 1 g of methanol was burnt in excess air, 22.6 kJ of heat were liberated. What was the quantity of heat in kJ liberated when 1 mole of methanol was burnt under similar conditions?

$$\text{A. } 22.6$$

$$\text{B. } 32.0$$

$$\text{C. } 723.2$$

$$\text{D. } 777.8$$

5. **(2003 Q.27)**

5.3 kJ of heat energy are required to vaporise 13 g of a liquid X (X = 78).

The molar heat of vaporisation of X in kJ/mole is

$$\text{A. } \frac{5.3 \times 78}{13}$$

$$\text{B. } \frac{80}{1.88 \times 188}$$

$$\text{C. } 13 \times 5.3 \times 78$$

$$\text{D. } \frac{188}{1.88 \times 80}$$

6. **(2003 Q.37)**

When 0.4 g ethanol was burnt, it raised the temperature of 0.1 kg of water by 20 °C. The heat of combustion of ethanol is

[Specific heat capacity of water = 4.2 kJ/kg/°C, C₂H₅OH = 46]

$$\text{A. } \frac{4.2 \times 20 \times 46}{0.4 \times 0.1} \text{ kJ mol}^{-1}$$

$$\text{B. } \frac{0.4 \times 4.2 \times 20}{46 \times 0.1} \text{ kJ mol}^{-1}$$

$$\text{C. } \frac{0.1 \times 4.2 \times 20 \times 46}{0.4} \text{ kJ mol}^{-1}$$

$$\text{D. } \frac{0.1 \times 4.2 \times 20}{46 \times 0.4} \text{ kJ mol}^{-1}$$

7. (2002 Q.13)

The amount of heat evolved when 6.0 g of metal, M was displaced from a solution was 28.8 kJ. The amount of heat produced when 0.5 mole of M was displaced is

A. $\frac{28.8 \times 54.9 \times 0.5}{6.0}$

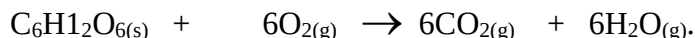
B. $\frac{54.9 \times 6.0}{28.8 \times 0.5}$

C. $\frac{28.8 \times 54.9}{6.0 \times 0.5}$

D. $28.8 \times 54.9 \times 6.0 \times 0.5$

8. (2002 Q.30)

Glucose burns in oxygen at 25 °C according to the equation below, giving out 2802 kJ mol⁻¹ of heat energy.



The amount of heat produced when 18.0 g of glucose is burnt in oxygen at the same temperature is (H = 1, C = 12, O = 16)

A. $\frac{2802 \times 18.0}{180 \times 25}$

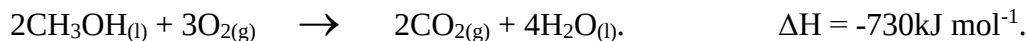
B. $\frac{180}{2802 \times 18.0}$

C. $\frac{180 \times 25 \times 18.0}{2802}$

D. $\frac{2802 \times 18.0}{180}$

9. (2001 Q.3)

Methanol burns in excess air according to the equation



The amount of heat liberated when 3.2 g of methanol, (Mr = 32.0) is completely burnt is

A. 73 kJ.

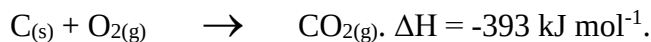
B. 730 kJ.

C. 1416 kJ.

D. 2929 kJ.

10. (2001 Q.38)

Carbon burns in excess oxygen according to the equation:



What mass of carbon in grams would produce 750 kJ of energy?

A. $\frac{393 \times 12}{750}$

B. $\frac{750 \times 12}{393000}$

C. $\frac{750 \times 12}{393}$

D. $\frac{750 \times 393}{12}$

11. (2000 Q.32)

13.70 kJ of heat was evolved when 4.0 g of copper was displaced from copper(II) sulphate solution by zinc. The amount of heat evolved when one mole of copper was displaced is

A. $\frac{63.5 \times 4}{13.7}$

B. $\frac{13.7 \times 63.5}{4}$

C. $\frac{13.7 \times 4}{63.5}$

D. $\frac{63.5 \times 4}{13.7}$

12. (1998 Q.14)

When 8 grams of a salt was dissolved in 100 g of water the temperature decreased by 10°C. The drop in temperature when 2 grams of the salt is dissolved in 100 g of water would be

A. 10 °C.

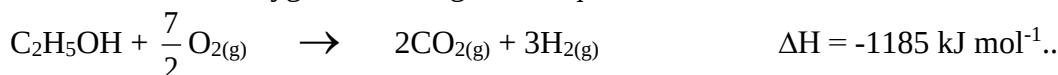
B. 8.5 °C.

C. 5.0 °C.

D. 2.5 °C.

13. (1998 Q.36)

Ethanol burns in oxygen according to the equation



Calculate the amount of heat given out when 0.2 moles of ethanol is burned completely.

A. -237 kJ.

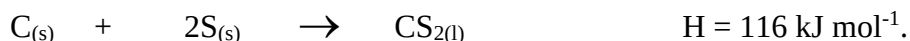
B. -592.5 kJ.

C. -1185 kJ.

D. -2370 kJ.

14. (1997 Q.17)

Carbon reacts with sulphur according to the following equation:



The amount of heat absorbed when 16 g of sulphur reacts with excess carbon is (C = 12, S = 32)

A. 7 kJ.

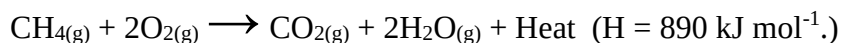
B. 29 kJ.

C. 58 kJ.

D. 116 kJ.

15. (1996 Q.8)

Methane burns in air according to the equation:



The energy liberated when 4 g of methane is burnt in air: (C = 12, H = 1)

A. 222.5 kJ.

B. 445.0 kJ.

C. 1780.0 kJ.

D. 3560.0 kJ.

SELF-CHECK 10.2

1.
 - (a) Define enthalpy of combustion.
 - (b) Describe an experiment that can be carried out to determine the heat of combustion of propanol. Draw a diagram to illustrate your answer.
 - (c) When 0.54 g of propanol was burnt, the heat evolved raised the temperature of 150 cm³ of water from by 21.5 °C. Calculate the heat of combustion of propanol. (C = 12, H = 1, O = 16, Specific heat capacity of water is 4.2 kJ g⁻¹, density of water = 1 gcm³,).
2. (2001 Q.8)
 - (a) Define the term **enthalpy (heat) of neutralisation**.
 - (b) When 50 cm³ of a 0.5M hydrochloric acid was added to 50 cm³ of a 0.5 M potassium hydroxide in a calorimeter, there was a temperature rise from 27.5 °C to 30.8 °C. Calculate the enthalpy of the reaction.
(Density of water is 1 g dm⁻³ and its heat capacity is 4.2 Jg⁻¹)

EXTRA SELF-CHECK QUESTIONS

1. (1994 Q.14)
 - (a) Describe an experiment that can be carried out to determine the heat of combustion of ethanol. Draw a diagram to illustrate your answer.
 - (b) Would you expect the heat of combustion as determined in the experiment in (a) to be greater than, lower than or equal to the theoretical value? Give a reason.
 - (c) When 0.382 g of ethanol was burnt, the heat evolved raised the temperature of 100 g of water from 16.5°C to 43.5°C.
Calculate the heat of combustion of ethanol.
(Heat capacity of water is 4.2 kJ g⁻¹)
 - (d) Name two products, other than water, of incomplete combustion of ethanol.
2. (2003 Q. 8)
 - (a) What is meant by the term **enthalpy of neutralisation**?
 - (b) When 50.0 cm³ of a 1M sulphuric acid was added to 50.0 cm³ of a 2M sodium hydroxide, the temperature of the resultant mixture rose by 13.6 °C.
 - (i) Write an ionic equation for the reaction that took place.
 - (ii) Calculate the enthalpy of neutralisation of sodium hydroxide. [Specific heat capacity of water = 4.2 J g⁻¹°C⁻¹, density of water = 1 gcm⁻³]

CHAPTER ELEVEN

EXTRACTION OF METALS

(Mineral Resources)

Introduction

The reactivity or electrochemical series is a valuable guide in the study of metals and their compounds. Only the metals low in the activity series such as gold, silver, mercury and copper which are unreactive metals, occur naturally. All the other metals occur as compounds in mineral ore.

The reactive metals are extracted by electrolysis, while the metals in the middle of the series (zinc and iron) are obtained by reducing their oxides by coke or carbon monoxide.

- NB:**
- The method of extraction of metals depends on the stability of the metal.
 - the extraction of metals is a reduction process.
 - The metals are in their ores as positive ions and extracted by supplying the ions with electrons.
 - The cathode supplies electrons to the metallic ions during electrolysis and
 - The coke combines with oxide ions of the oxides to release electrons which reduce the metallic ions.

Metal	Chief ores	Method of Extraction
Na	Rock salt, NaCl Chile saltpetre, NaNO ₃	Electrolysis
Al	Bauxite (Al ₂ O ₃ .2H ₂ O) Cryolite, Na ₃ AlF ₆ Corundum (Al ₂ O ₃) Mica and China clays	Electrolysis of oxide
Zn	Zinc blende, Zinc sulphide (ZnS) Calamine, Zinc carbonate (ZnCO ₃)	Reduction of oxide by carbon or carbon monoxide
Fe	Hematite, Iron (III) oxide (Fe ₂ O ₃) Magnetite, Fe ₃ O ₄ Siderite, Iron (II) carbonate, FeCO ₃ Pyrite, iron disulphide, FeS ₂	
Cu	Copper pyrite, iron disulphide (CuFeS ₂) Cuprite, Cu ₂ O Chalcocite, Cu ₂ S Malachite, CuCO ₃ .Cu(OH) ₂	Roast in air

EXTRACTION OF IRON

Occurrence:

Iron is the second most abundant metal after aluminium. It forms 4% of the earth's crust.

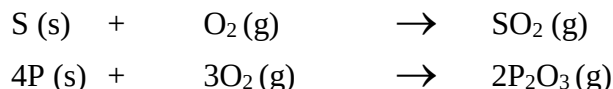
It exists in form of ores. The chief ores being:

- Hematite - (Fe_2O_3) - found in United States, Australia & Russia.
- Magnetite - (Fe_3O_4) - found in Sweden & North America.
- Siderite - (FeCO_3). - found in Britain.

These iron ores are mined from their deposits and transported to the plants for extraction.

Concentration of the ore

The richer ores are picked from the earthly impurities, crushed and then roasted in air to remove water vapour and non metallic impurities such as sulphur and phosphorous into their respective oxides according to the equations:

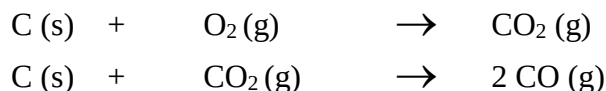


The removal of these non metals leaves Silicon (IV) oxide (SiO_2) as the main impurity.

Extraction of Iron from Hematite

The mixture of the *iron ore*, *coke* and *limestone* is fed into a blast furnace from the top.

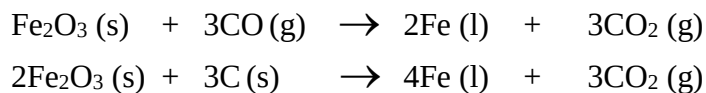
Preheated air is injected into the furnace through the pipes called tuyeres at the lower part of the furnace. The air oxidizes the hot coke to carbon dioxide.



NB: *The reaction is exothermic hence large quantity of heat is liberated.*

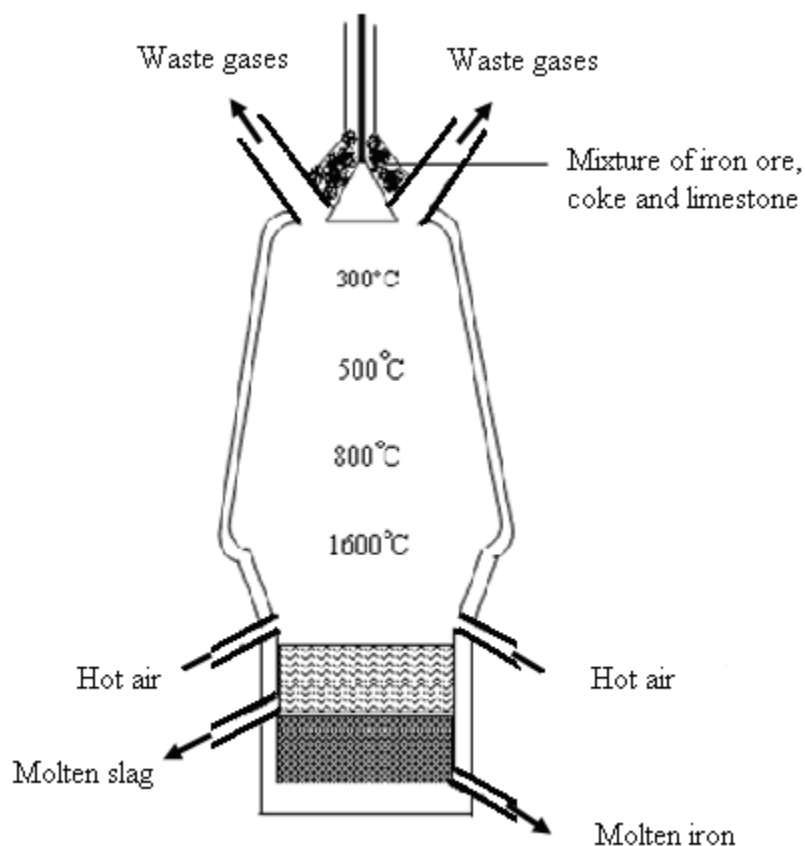
Reduction of the ore Fe_2O_3 by Carbon monoxide (CO) and Carbon (C)

The carbon monoxide and the unchanged carbon reduce the iron (II) oxide to iron at temperature ranges of (500 – 800) °C according to the equations:



The hot iron formed absorbs carbon from the coke and lowers its melting point. It then sinks and collects at the bottom of the furnace in molten form, where it is tapped off and solidified into blocks called “*pig iron*” or *cast iron*.

The diagram showing the blast furnace

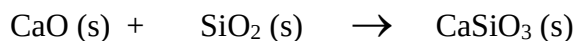


The Role of the Limestone

- The purpose of the limestone is to remove Silicon (IV) oxide which is the main impurity.
- It first decomposes to calcium oxide (quick lime CaO) according to the equation:



The calcium oxide then combines with Silicon (IV) oxide to form Calcium silicate (CaSiO₃), the slag, according to the equation:



The slag floats on the molten iron and is tapped off separately.

Purification of Cast Iron (Puddling)

- The cast iron obtained from the blast furnace is 90 - 95% pure. The main impurities are carbon, silicon, manganese, sulphur and phosphorus depending on the original ore.
- Pure iron (wrought iron) is obtained by heating a mixture of pig iron with iron (II) oxide (FeO) as an oxidizing agent and lime stone in the blast furnace.
- The oxygen of the iron (II) oxide oxidizes the impurities as follows:
 - Carbon and sulphur oxidized to gaseous oxide (CO₂ and SO₂) which escape as waste gases at the top of the furnace,
 - Phosphorus to phosphate and
 - Silicon to silicate.

Use of Iron

(a) Cast Iron

Contains about 3-5% carbon, 1% silicon and 2% phosphorus.

It is brittle but extremely hard.

(i) It is used in making:

- Furnace, grates, Bunsen burner bases, drainage pipes, iron boxes etc.

(ii) It is used in the manufacture of wrought iron and steel.

(b) Wrought Iron

It contains about 0.1% carbon. It is malleable and can be welded. As a result it is used to make:

- Iron nails,
- iron sheets,
- horse shoes and
- agricultural implements.

(c) Steel

Steel is manufactured by blowing oxygen into molten pig iron in presence of lime. The impurities carbon, sulphur and phosphorus present are oxidized by the oxygen as gaseous oxides. The lime serves to form slag and is removed. Calculated amount of carbon is added to the iron formed. One or more metals such as chromium, nickel, manganese and tungsten are also added in small calculated amounts. This gives steel. It is hard, tough and strong.

Uses of steel

Steel is used for making the following:

- Cutlery i.e. stainless forks, spoons and knives.
- Drilling and high speed cutting tools.
- Electromagnets.
- Rock drills and railway lines
- Spanners.
- Tin is used in food preservation for canning fruits, meat and fish.

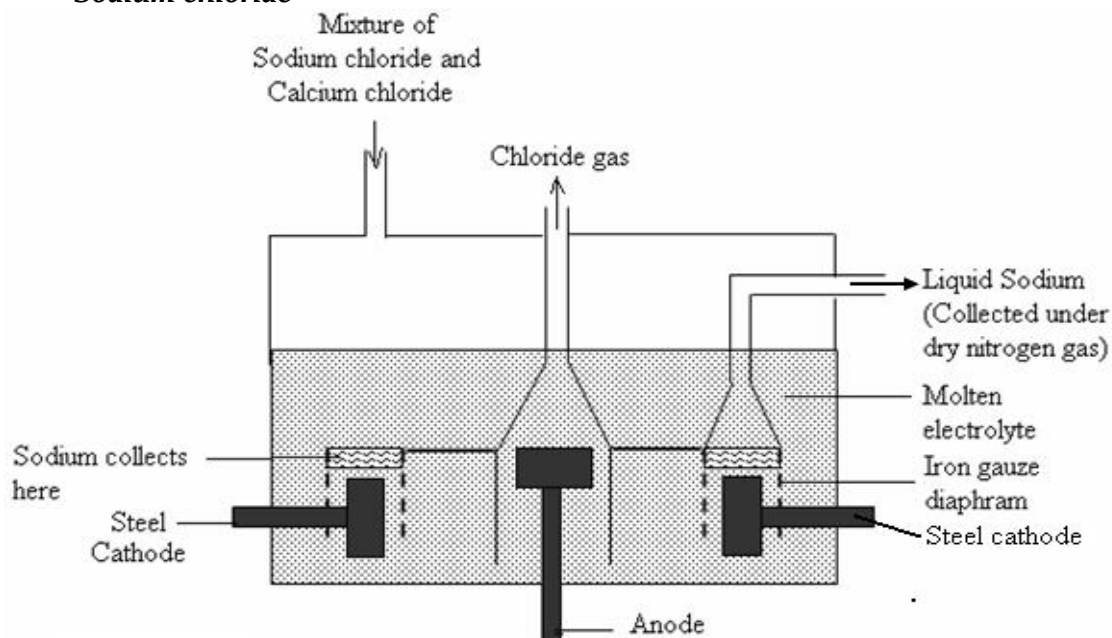
3. **Extraction of reactive metals** e.g. Potassium, Sodium, Calcium, Magnesium and Aluminium.

Extraction of reactive metals Sodium from Sodium Chloride.

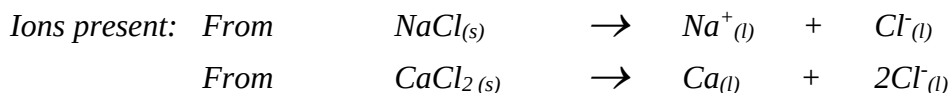
Sodium is extracted by the electrolysis of molten sodium chloride in the Downs Cell. The cell consists of a cylindrical steel cathode and graphite anode.

Sodium chloride is heated and melts at 801°C. Calcium chloride is added as an impurity to lower the melting point to about 600 °C.

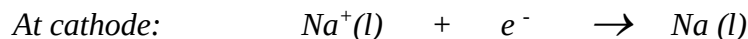
The Diagram of Downs Cell showing the Extraction of Sodium from Molten Sodium chloride



Explanation:



Reaction equation



The products Sodium and Chlorine produced are kept apart to prevent them from reacting and reforming sodium chloride.

The liquid Sodium is collected under dry nitrogen gas. This prevents the metal from reacting with the atmosphere.

SELF-CHECK 11.1

- (1996 Q.37) Which one of the following methods can be used to extract magnesium from its ore?
- A. Decomposition by heat. B. Electrolysis.
C. Reduction with carbon monoxide. D. Crystallisation.
2. (2004 Q.9) Which one of the following metals is extracted by electrolysis?
- A. Zinc. B. Lead. C. Sodium. D. Copper.
3. (2002 Q.1) Which one of the following ores is a valuable ore of copper?
- A. Malachite. B. Siderite. C. Haematite. D. Dolomite.
4. (2002 Q.33) Which one of the following gases is used to extract iron from its ore?
- A. Chloride. B. Nitrogen monoxide.
C. Carbon monoxide. D. Sulphur trioxide.
5. (2001 Q.4) The role of coke in the extraction of iron in the blast furnace is to
- A. produce carbon monoxide which reduces the oxides.
B. produce quick lime which combines with silica.
C. combine with iron to form steel.
D. reduce excessive heat produced in the furnace.
6. (1998Q.18) Which one of the following substances is not used in the extraction of iron?
- A. Coke. B. Air. C. Silica. D. Limestone.
7. (1996 Q.32) Which of the following metals can be extracted by reduction of the oxide with carbon?
- A. Potassium. B. Aluminium.
C. Zinc. D. Magnesium.
8. (1996 Q.35) Which one of the following reactions does **not** take place in the extraction of iron in the Blast furnace?
- A. Carbon monoxide, reduces iron (II) oxide to iron.
C. Coke burns in air forming carbon dioxide.
B. Limestone decomposes to form calcium oxide.
D. Limestone reduces iron (II) oxide to iron

SELF-CHECK 11.2

1. 1988 Q.12

- (a)
 - (i) Name one ore of each of the following metals; sodium and iron.
 - (ii) Briefly describe how sodium and iron are extracted from their ores. Explain why the method you have described can be used to extract the metal from the ore.
- (b) State the conditions under which sodium and iron react with water. Write equations for the reaction in each case.

2. 1995 Q.14

- (a) Name one ore from which sodium can be extracted.
- (b) Describe how sodium is extracted from the ore you have named in (a). Your answer should include the following:
 - (i) Names of the materials used as the electrodes.
 - (ii) *Equations for the reactions that took place at the electrodes.*
 - (iii) *Method of collecting the sodium produced.*

[A diagram is not required]
- (c) State two factors that you would consider if you were to choose a place in Uganda where a plant for extracting sodium could be built.
- (d) A piece of sodium was heated and plunged into a jar of chlorine.
 - (i) State what was observed.
 - (ii) Write the equation for the reaction which took place.
- (e) When aqueous silver nitrate was added to a solution of the product in (d), a white precipitate was formed.

Write equation for the reaction that took place.

3. (2005 Q.12)

In the extraction of cast iron using a blast furnace, spathic iron ore, which contains some impurities, is first roasted in air. It is then mixed with some other substances and finally introduced into the blast furnace. Cast iron can be obtained from iron (II) carbonate ore.

- (a) Name the major impurity in the iron ore.
- (b)
 - (i) Give the chemical name of the spathic iron ore.
 - (ii) Write an equation for the reaction which takes place when iron (II) carbonate is roasted in air.
- (c) Name the substances that are fed into the blast furnace:
 - (i) from the top.
 - (ii) from the bottom.
- (d) Outline the reactions leading to:
 - (i) the formation of cast iron.
 - (ii) the removal of the major impurity you have named in (a).
- (e) State the major components of steel.

4. (1993 Q.11)

In the extraction of iron ore, coke and limestone are fed into a blast furnace and hot air is blown into the mixture.

- (a)** Name and give the formula of one ore of iron.
- (b)** Why is limestone added to the mixture?
- (c)** Write equations for the reactions that lead to the formation of iron.
- (d)** Describe what happens when
 - (i)** dilute hydrochloric acid is added to iron fillings.
 - (ii)** steam is passed over heated iron fillings.

Write equation for the reaction that takes place in each case.

ANSWERS TO SELF CHECK QUESTIONS

CHAPTER ONE

THE MOLE CONCEPT SELF-CHECK 1.1

1. B.	2. A.	3. B.	4. D.	5. A.
6. B.	7. C.	8. D.	9. B.	10. A.
11. B.	12. B.	13. C.	14. B.	15. B.
16. C.	17. A.	18. D.	19. B.	20. C.

SELF-CHECK 1.2

1. (a) (i) Percentage of hydrogen = 20%, Percentage of carbon = $(100 - 20)\%$
= 80%

Elements	C	H	
Percentage mass	80	20	
Moles		$\frac{80}{12}$	$\frac{20}{1}$
	6.667	20	
Mole ratios	$\frac{6.667}{6.667}$	$\frac{20}{6.667}$	
	1	3	

The empirical formula of X is CH_3

- (ii) Mass of X = 7.5 g, Volume occupied by 7.5 g of X at stp = 5.6 dm^3
Molar mass = ?, Molar gas volume = 22.4 dm^3

7.5 g of X occupies 5.6 dm^3 at stp.

1 g of X occupies $\frac{5.6}{7.5} \text{ dm}^3$.

The mass of X (molar mass) that will occupy 22.4 dm^3 at stp

$$= \frac{7.5 \times 22.4}{5.6}$$

$$= 30.4$$

$\therefore$ The molar mass of X = 30 g

- (iii) (Empirical formula) n = Molecular mass
 $(\text{CH}_3) n$ = 30
 $[12 + (1 \times 3)] n$ = 30
 $15 n$ = 30

$$n = \frac{30}{12}$$

$$= 2$$



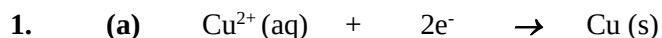
$\therefore$ The molecular formula is C_2H_6

- (b) (i) X is ethane.
(ii) CH_3CH_3

CHAPTER TWO

ELECTROLYSIS

SELF-CHECK 2.1



- (b) (i) $I = 0.20 \text{ A}$, $t = 20 \text{ min} \times 60 = 1200 \text{ s}$, $1F = 96500 \text{ C}$, $Cu = 64$

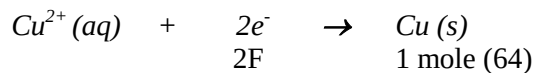
$$\begin{aligned} Q &= I t \\ &= 0.20 \times 1200 \\ &= 240 \text{ C} \end{aligned}$$

- (ii) 1 mole of coulombs = 96500 C.

Then the number of moles of coulombs in 240 C

$$\begin{aligned} &= \frac{1 \times 240}{96500} \\ &= 0.0025 \text{ moles} \end{aligned}$$

- (iii) Equation at cathode:



If $(2 \times 96500) \text{ C}$ liberates 64 g of copper,

$$1 \text{ coulomb liberates } \frac{64}{2 \times 96500} \text{ g of copper.}$$

$$240 \text{ C will liberates } \frac{64}{2 \times 96500} \times 240 = 0.08 \text{ g}$$

Or Alternatively:

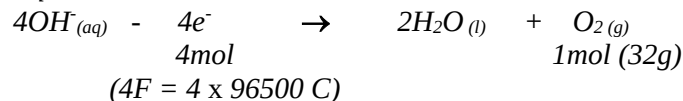
2 moles of coulombs of electricity liberates 1 mole of copper (64) g,

$$1 \text{ mole of coulomb liberates } \frac{64}{2} \text{ g of copper.}$$

$$0.0025 \text{ moles of coulombs liberates } \frac{64}{2} \times 0.0025 = \mathbf{0.08 \text{ g}}$$

2. (a) Anode.

(b) Equation at anode:



(c) $t = 35 \text{ min} = 35 \times 60 = 2100 \text{ s}$, $I = 0.65 \text{ A}$, $F = 96500 \text{ C}$, $Q = ?$

$$\begin{aligned} Q &= I t \\ &= 0.65 \times 2100 \\ &= \mathbf{1365 \text{ C}} \end{aligned}$$

(4 x 96500) C of electricity liberates 32 g of oxygen,

1 Coulomb of electricity liberates $\frac{32}{4 \times 96500}$ g of oxygen gas.

1365 C of electricity will liberate $\frac{32}{4 \times 96500} \times 1365 = \mathbf{0.113 \text{ g}}$

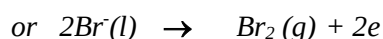
3. (a) (i) No observable change (i.e the bulb did not glow).

(ii) The lamp glowed brightly.

Reddish brown colouration was seen around the anode.

(b) Before the lead (II) bromide was melted, the Pb^{2+} ions and Br^- were not free to move. But after melting the ions were free to move in the electrolyte, hence conducting current, leading to the glowing of the lamp.

(c) (i) $P = \text{Anode: } 2\text{Br}^-(\text{l}) - 2e^- \rightarrow \text{Br}_2(\text{g})$



(ii) $Q = \text{Cathode: } \text{Pb}^{2+}(\text{l}) + 2e^- \rightarrow \text{Pb}(\text{l})$

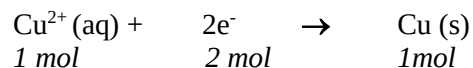
4. (a) (i) P cathode Q. anode

R cathode S. anode

(ii) Bubbles of a colourless gas that burns with a pop sound are evolved.

(b) P. $\text{Cu}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Cu}(\text{s})$

(c) S. $4\text{OH}^-(\text{aq}) - 4e^- \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$



If 2F produce 1 mol of copper, then the number of Faraday that produce 0.02 moles of copper
 $= 2 \times 0.02$

$$= 0.04 \text{ F}$$

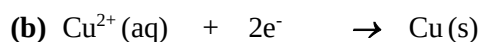
Equation for reaction at S.



If 4F liberate 1 mol of oxygen, then the number of moles of oxygen liberated by 0.04

$$\text{F} = \frac{1 \times 0.04}{4} = 0.01 \text{ moles}$$

5. (a) (i) At Anode: - Bubbles of a colourless gas that rekindles a burning splint were formed.
(ii) At Cathode: - A brown solid deposit.



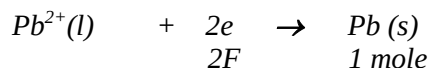
The positively charged ions Cu^{2+} and H^+ ions move to the cathode. The Cu^{2+} ions which are lower in the electrochemical series are preferentially discharged by gain of electrons from the cathode according to the equation above:

6. (a) No observable change.
a. The lamp glowed brightly.
Bubbles of reddish brown gas were seen around the anode.
(b) Before the lead (II) bromide was melted, the Pb^{2+} ions and Br^- were not free to move. But after melting the ions were free to move in the electrolyte, hence conducting current, leading to the glowing of the lamp.

7. (a) (i) At Anode: - Reddish brown colouration of bromine gas.
At Cathode : A grey melt solid formed melts and sinks to the bottom of the crucible.
(ii) At anode: $2\text{Br}^-(\text{l}) - 2e^- \rightarrow \text{Br}_2(\text{g})$
or $2\text{Br}^-(\text{l}) \rightarrow \text{Br}_2(\text{g}) + 2e^-$
At cathode: $\text{Pb}^{2+}(\text{l}) + 2e^- \rightarrow \text{Pb}(\text{l})$

$$\text{(b) } t = 1\frac{1}{2} \text{ hours} = 90 \times 60 = 5400 \text{ s, } I = 2 \text{ A, } Q = ?$$

$$\begin{aligned} Q &= I t \\ &= 2 \times 5400 \\ &= 10,800 \text{ C} \end{aligned}$$



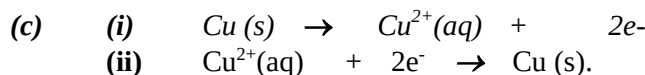
(2 x 96500) C of electricity liberates 207g of lead,

1 coulomb of electricity liberates $\frac{207}{2 \times 96500}$ g of lead.

10800 C of electricity will liberate $\frac{207}{2 \times 96500} \times 10800 = \mathbf{11.58\text{ g}}$

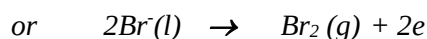
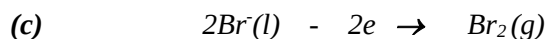
8. (a) (i) Impure copper.
(ii) Pure copper.

(b) Copper (II) sulphate solution. or any soluble salt of copper.

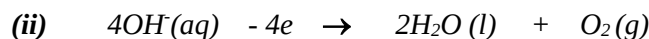


9. (a) In solid state, lead (II) bromide the ions Pb^{2+} ions and Br^- are not free to move. But in molten state the ions are free to move in the electrolyte, hence are able to conduct electricity between the electrodes.

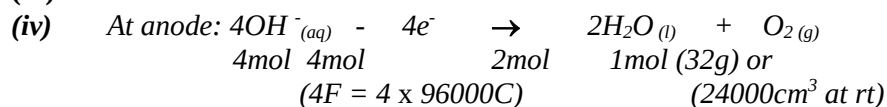
- (b) (i) Reddish brown colouration of bromine gas.
(ii) A grey solid formed melts and sinks to the bottom of the crucible.



10. (a) (i) X - Oxygen
Y - Hydrogen



(iii) Electron flow is from anode to cathode.



- (b) (i) $t = 10 \text{ min} = 10 \times 60 = 600 \text{ s}, I = 1 \text{ A}, F = 96000 \text{ C}, Q = ?$

$$\begin{aligned} Q &= I t \\ &= 1 \times 600 \\ &= \mathbf{600 \text{ C}} \end{aligned}$$

If $(4 \times 96000) \text{ C}$ liberates 24000 cm^3 of oxygen gas at room temperature,

1 Coulomb of electricity liberates $\frac{24000}{4 \times 96000}$ of oxygen gas.

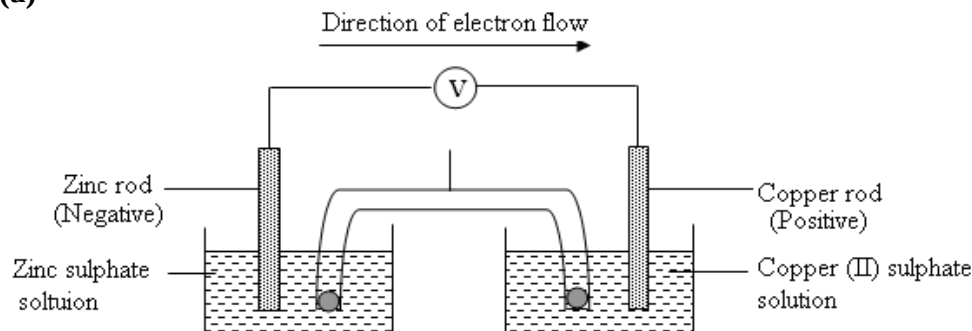
$$600 \text{ C liberates } \frac{24000}{4 \times 96000} \times 600 = \mathbf{375 \text{ cm}^3}$$

- (ii) Electroplating, Extraction of reactive metals and Purification of copper
(iii) Chlorine is a poisonous gas. Its leakage to the environment would poison the workers.

ELECTROCHEMICAL CELL

SELF-CHECK 2.2

1. (a)



(b) To complete the circuit between the zinc sulphate and copper (II) sulphate solutions.

- (c) (i) - At anode: $\text{Zn (s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ (Oxidation)
 (ii) - At cathode: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$ (Reduction)

(d) $E = 0.9 \text{ V}$, $n = 2$ (Valence of Zn and Cu = 2)

$$\begin{aligned} \text{Energy} &= EnF \\ &= 0.90 \times 2 \times 96500 \\ &= \mathbf{173.7 \text{ KJ}} \end{aligned}$$

(e) The voltmeter reading would be greater than 0.90 V

(f) Because magnesium is higher than zinc in the reactivity series.

2.

(a) (i) $\text{Zn (s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$

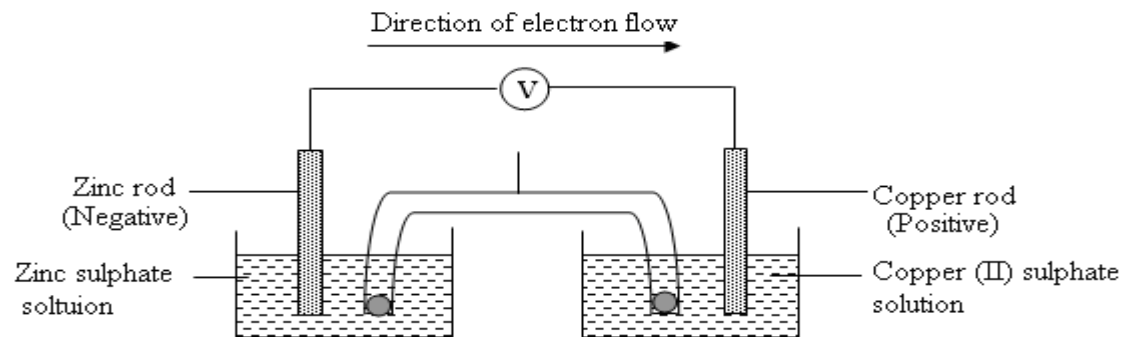
(ii) There will be a complete discharge of Cu^{2+} ions and dissolution of zinc rod leading to electrolysis in which the zinc rod is dissolved and Cu^{2+} ions are completely discharged and so the reaction ceases.

(b) $E = 1.10 \text{ V}$, $n = 2$ (Valence of Zn and Cu = 2)

$$\begin{aligned} \text{Energy} &= EnF \\ &= 1.10 \times 2 \times 96500 \\ &= \mathbf{212.3 \text{ KJ}} \end{aligned}$$

3.

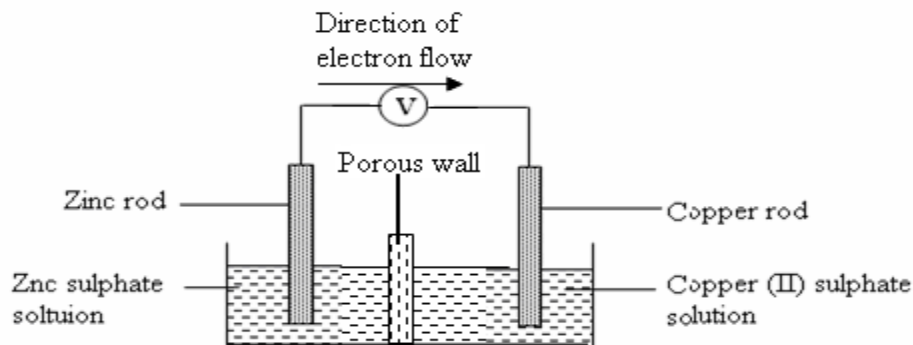
(a) and (b) The diagram showing Daniel cell



- (c) (i) - At anode: $\text{Zn (s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ (Oxidation)
 - At cathode: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$ (Reduction)

(ii) $\text{Zn (s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$

4. (a) (i) - At cathode: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$ (Reduction)
 - At anode: $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ (Oxidation)
 (ii) $\text{Zn}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$
 (b)



5. (a) (i) Zinc sulphate solution (ii) Lead (II) nitrate.
 (b) Zinc rod or plate.
 (c) (i) **At the anode** i.e the metal higher in the E.C.S, Zinc
 $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ (Oxidation)
 (ii) **At the cathode:** i.e the metal lower in the E.C.S, Lead
 $\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$ (Reduction)
 (d) $\text{Zn}(\text{s}) + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Pb}(\text{s})$

CHAPTER THREE

REACTION RATES

SELF-CHECK 3.1

1. C.	2. C.	3. B.	4. A.	5. A.
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SELF-CHECK 3.2

1. (i) Graph B.
Reason - Because the particles are larger, the total surface area is less for B compared to A.
NB: - Reaction rate decreases with decrease in surface area.
 (ii) The hydrogen peroxide is becoming dilute as more water molecules are produced in the reaction.
NB: - Reaction rate decreases with decrease in concentration.
 (iii) See the broken curve on the graph (Page 91)

Reason: Reaction rate decreases with decrease with drop in temperature.

- (iv) In each case the same quantity of hydrogen peroxide is used, the same volumes of oxygen gas is produced at the end of the reaction. So all the graphs begin to level at the same value. And the difference in the slopes is due to the different factors of rate of reaction affecting the reaction.

CHAPTER FOUR

ORGANIC CHEMISTRY

SELF-CHECK 4.1

1. D.	2. C.	3. C.	4. B.	5. D.
6. C.	7. A.	8. D.	9. D.	10. C.
11. B.	12. B.	13. D.	14. C.	15. B.
16. B.	17. B.	18. B.	19. D.	20. B.
21. B	22. A	23. C	24. D	25. D
26. A				

SELF-CHECK 4.2

- (a) $\text{CH}_2=\text{CH}_2$ or $\text{H}_2\text{C}=\text{CH}_2$

```

      H       H
       \     /
        C = C
       /     \
      H       H
    
```

or

(b) (i) The reddish brown colour of bromine becomes colourless
The brown colour of bromine water becomes colourless.

$$\text{C}_2\text{H}_4(\text{g}) + 2\text{Br}_2(\text{aq}) \rightarrow \text{C}_2\text{H}_4\text{Br}_2(\text{aq})$$

Or $\text{C}_2\text{H}_4(\text{g}) + 2\text{Br}_2(\text{aq}) \rightarrow \text{CH}_2\text{Br}-\text{CH}_2\text{Br}(\text{aq})$

(ii) Acidified potassium manganate (VII).

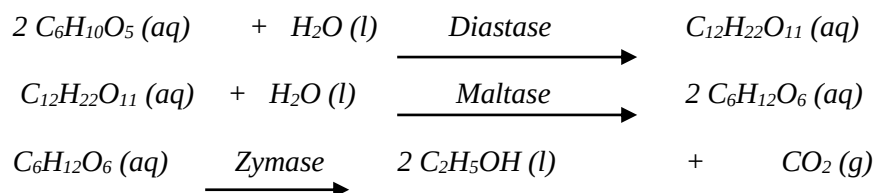
(c) Ethanol.
- (a) X - Vegetable oil

(c) Saponification

(d) Brine (Concentrated sodium chloride).

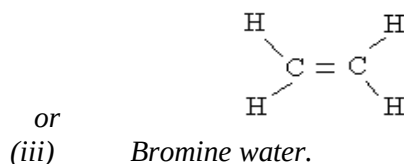
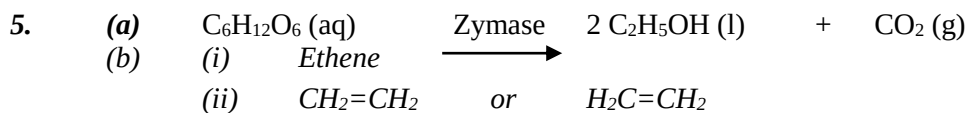
(e) White precipitate.

3. (a) *Millet flour, cassava flour, maize flour etc.*
- (b) *Maize, cassava and millet flour are mixed with water. The product is soaked in a tin or pot or buried in the ground for about a week to allow starch to be converted into sugar.*
The mixture is then removed, roasted and dried. The product is then allowed to ferment for two to three days in suitable container. During this process, the sugar formed is converted to crude ethanol.
The formation of the crude ethanol from the carbohydrates (starch) involves the participation of enzymes.
The relevant equations for the main reactions involved are given below.



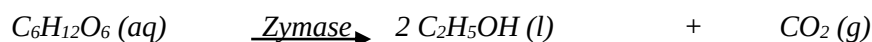
- (c) *The ethanol is fractionally distilled at 78 – 82°C to a liquid containing about 95% of ethanol and 5% water. The water can be absorbed by absorption using quicklime (CaO).*
- (d) (i) $\text{C}_2\text{H}_5\text{OH} (\text{l}) \rightarrow \text{C}_2\text{H}_4 (\text{g}) + \text{H}_2\text{O} (\text{l})$
- The sulphuric acid must be concentrated.
- The mixture of ethanol and concentrated sulphuric acid must be heated to a temperature of 170 °C.
- (ii) *- beverage.*
- fuel.
- thermometric liquid in thermometers.
- A solvent for organic compounds
4. (a) (i) *Oils occur naturally in plants while fats occur naturally in animals only.*
Fats are solid at room temperature while oils are liquid at room temperature.
 (ii) *Examples of oil include:*
- Sunflower oils, Groundnut oils, simsim oils, corn oil etc.
Examples of fats include:
- Lard in pigs, butter from milk.
- (b) *Soap is manufactured by a process called saponification.*
A mixture of vegetable oil and concentrated sodium hydroxide solution is boiled.
 $\text{Oil or Fat} + \text{Sodium hydroxide} \rightarrow \text{Soap} + \text{Glycerol}$
When the process is complete, brine (concentrated sodium chloride) is added to the mixture to precipitate (to solidify) the soap. The soap, which floats above the liquid, is removed and pressed into bars.
- (c) *White precipitate.*

- (d) The Mg^{2+} ions in the magnesium hydrogen carbonate reacts with soap to form white solid called scum.
- (e) There would be no observable change.
- (f) They cause water pollution in rivers and lakes.



Or Acidified potassium manganate (VII).

6. (a) (i) Decomposition (ii) Zymase.



- (b) The ethanol is fractionally distilled at $78 - 82^\circ\text{C}$ to a liquid containing about 95% of ethanol and 5% water. The water can be absorbed by absorption using quicklime (CaO).

7. (a) (i) Polymerization is the joining together of simple molecules called monomers to form a complex molecule called polymer.

- (ii) Natural polymers:

Starch, proteins, Fats, Cellulose, Glycogen, Rubber, Wool, Silk.

Synthetic polymer:

Polythene, Polypropene, Perspex, Nylon, Bakelite,

Synthetic rubber, Poly Vinyl Chloride (PVC), Nylon

- (b) $\text{CH}_3\text{CH}=\text{CH}_2$ (Propene).

- (c) - **Thermoplastics** are plastics which soften on heating and harden on cooling.
 - Softening and hardening are reversible.
 - As a result they can be molded or remolded under hot conditions.
 - **Thermo sets** (Thermosetting plastics) are plastics, which soften or melt on heating during manufacture and take the shape of the mold in which they are processed on cooling.
 - They cannot be remolded by heating after manufacture.
 - At high temperatures, they decompose.

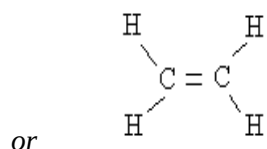
- (d) Cracking is the process by which long chain alkanes are broken down to produce shorter chain hydrocarbons. E.g. cracking of paraffin in the lab.

(e) See diagram on page 108

8. (a) A - ethanol

(b) (i) Ethene

(ii) $\text{CH}_2=\text{CH}_2$ or $\text{H}_2\text{C}=\text{CH}_2$



(iii) Bromine water.

Or Acidified potassium manganate (VII).

(iv) With bromine water:

The reddish brown colour of bromine becomes colourless.

With acidified potassium manganate (VII):

The pink colour of the solution becomes colourless.

CHAPTER FIVE

NITROGEN AND ITS COMPOUNDS

SELF-CHECK 5.1

1. C.	2. D.	3. D.	4. A.	5. B.
6. D.	7. A.	8. A.	9. B.	10. C.
11. B.	12. C.	13. B.	14. D.	15. C.
16. A.	17. B.	18. A.	19. B.	20. D.
21. C.	22. A	23. C.	24. B.	25. B.

SELF-CHECK 5.2

1. (a) An alkali is a base that is very soluble in water to give hydroxide ions.

Or An alkali is a base which gives hydroxide ions (OH^-) in solution.

(b) (i) Ammonia is a basic gas that is very soluble in water thus it would dissolve in water instead of being collected.

(ii) Ammonia is less dense than air or lighter than air and would not therefore displace air from the gas jar.

(c) Hydroxide ion (OH^-) ions.

(d) Volume of ammonia at room temperature = 240 cm^3
Molar Gas Volume any gas at rt = 24000 cm^3

$$\begin{aligned}\text{Number of moles of a gas} &= \frac{\text{Volume of at room temperature}}{\text{Molar Gas Volume}} \\ &= \frac{240}{24000} \\ &= \mathbf{0.01 \text{ moles}}\end{aligned}$$

(e) (i) The ammonium sulphate solution should be heated to evaporate the excess water or to concentrate the solution slowly to crystallization point, allowed to cool until crystals form, then filtered, dried using filter paper.

(ii) From the equation, 2 moles of ammonia produce 1 mole of ammonium sulphate. But from (d) moles of ammonia = 0.01.
Therefore, moles of ammonium sulphate produced by 0.01 moles

$$= \frac{0.01}{2}$$

$$= \mathbf{0.005}$$

Formula mass (Molar mass) of ammonium sulphate, $(\text{NH}_4)_2\text{SO}_4$

$$= 3(14 + 4 \times 1) + 32 + 4 \times 16$$

$$= 3 \times 18 + 32 + 64$$

$$= 36 + 32 + 64$$

$$= \mathbf{132 \text{ g}}$$

$$1 \text{ mole of ammonium sulphate} = 132 \text{ g}$$

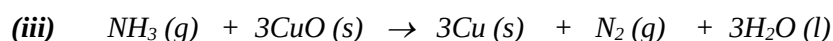
$$0.005 \text{ moles of ammonium} = \frac{0.005}{1} \times 132 \text{ g}$$

$$= \mathbf{0.66 \text{ g}}$$

2. (a) (i) Gas A is ammonia.

Liquid L is water.

(ii) Black solid, Copper (II) oxide changes to reddish brown solid because it is reduced to copper by ammonia.



(iv) I. From the equation, 2 moles of ammonia produce 1 mole of nitrogen gas.

Therefore, the volume of nitrogen produced by 320 cm³ of ammonia

$$= \frac{1}{2} \times 320$$

$$= 160 \text{ cm}^3$$

II. Number of moles of ammonia = $\frac{\text{Volume of at room temperature}}{\text{Molar Gas Volume}}$

$$= \frac{320}{24000}$$

$$= 0.0133$$

From the equation, 2 moles of ammonia reacts with 3 moles of Copper (II) oxide, then the number of moles of Copper (II) oxide that reacts with 0.0133 moles of ammonia

$$= \frac{3}{2} \times 0.0133$$

$$= 0.01995$$

$$= 0.02 \text{ moles}$$

Relative Formula Mass of Copper (II) oxide, CuO,

$$= 63.5 + 16.0$$

$$= 79.5 \text{ g}$$

From Number of moles = $\frac{\text{Mass of substance}}{\text{Relative Formula Mass}}$

Mass of Copper (II) oxide = Number of moles x RFM
 = 0.02 x 79.5

$$= 1.59 \text{ g}$$

III The excess ammonia from the reaction dissolves in the water in the beaker to form ammonium hydroxide which is a weak alkali or base of pH about 10.

(b) The burning splint would be extinguished.

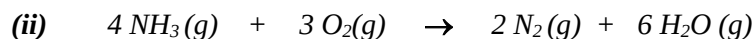
(c) Because it is cheaper and ammonia is made from nitrogen.

3. (a) A dry sample of ammonia is prepared in the laboratory by heating a mixture of ammonium chloride and calcium hydroxide. The gas is dried by passing it through calcium oxide (quick lime) and then is collected

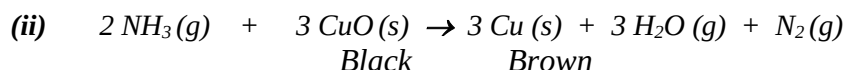
by upward delivery method since it is lighter (less dense) than air.

(b) Ammonia is tested by placing a moist red litmus paper in a stream of ammonia. The red litmus paper turns blue.

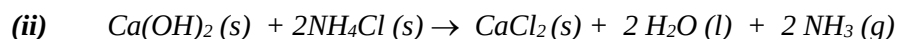
(c) (i) See diagram on page 124



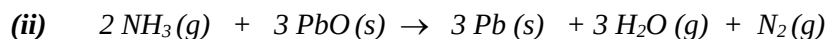
(d) (i) The black solid copper (II) oxide turned brown and
A colourless liquid that turned anhydrous copper (II) sulphate blue was formed.



4. (a) (i) See diagram on page 123

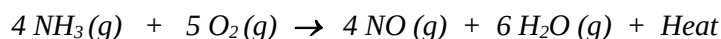


(b) (i) The black solid turned to grey and
A colourless liquid that turned anhydrous copper (II) sulphate blue was formed.



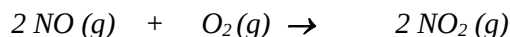
(c) **Oxidation of Ammonia**

Ammonia from Haber process is mixed with excess air and the mixture is passed over platinum catalyst heated at about 800 °C and 8 atmospheres pressure in the catalytic chamber. It is oxidized to nitrogen monoxide according to the equation:



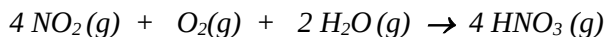
Oxidation of nitrogen monoxide to nitrogen dioxide

The nitrogen monoxide formed is rapidly cooled and combines with the oxygen from excess air to form nitrogen dioxide.

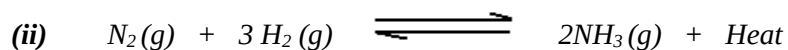


Oxidation of nitrogen dioxide and absorption

The nitrogen dioxide is oxidized by more air and then absorbed in hot water to form nitric acid.



5. (a) (i) Hydrogen and nitrogen



(b) (i) High pressure increases the percentage yield of ammonia.

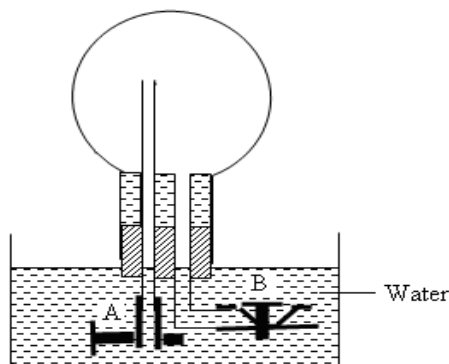
(ii) The percentage yield increases with increase in temperature.

(c) Presence of a catalyst

- (d) The black solid in the combustion tube turned brown and
A colourless liquid that turns anhydrous copper (II) sulphate blue was formed.
The ammonia reduces the copper (II) oxide to solid copper which is brown in colour.

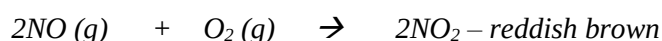
6. (a) See diagram on page 123.

(b) The fountain experiment

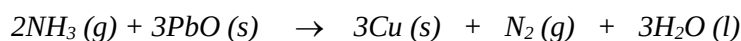


- Fill a large, dry, round – bottomed flask with dry ammonia.
- Place the mouth of the flask under water, coloured with litmus in a trough.
- Clamp it firmly, open the clip attached at the end of the long glass tube.
- So much ammonia is dissolved in the jet such that a partial vacuum is created in the flask.
- Air pressure forces water rapidly up the tube and enters the flask as a fountain. The litmus turns blue.

- (c) (i) The Copper coil glows red hot, reddish brown fumes of NO_2 are seen.
- (ii) This is because of the catalytic oxidation of ammonia by oxygen. The catalyst is the copper. The initial products are steam and nitrogen oxide, but the latter is immediately oxidized to NO_2 – the brown fumes.



(d) R.f.m of $\text{PbO} = 207 + 16 = 223$.



From the equation,

3 mols of PbO require 2 mols of NH_3

669g (2×223) of PbO require 2 mols of NH_3

1g of PbO requires $\frac{48}{669} \text{ dm}^3$ of NH_3

2.5g of PbO will require $\frac{48}{669} \times 2.5 = 0.1794 \text{ dm}^3$ of NH_3

7. (a) A Ammonia
- (b) (i) The black copper oxide turns brown (copper), a colourless vapour is seen.
(ii) $2\text{NH}_3(g) + 3\text{CuO}(s) \rightarrow 3\text{Cu}(s) + \text{N}_2(g) + 3\text{H}_2\text{O}(g)$
- (c) (i) Water
(ii) White anhydrous Copper (II) sulphate
8. (a) (i) A = Water
(ii) B = Nitrogen gas, N_2
(iii) Water
- (b) $2\text{NH}_3(g) + 3\text{CuO}(s) \rightarrow 3\text{Cu}(s) + \text{N}_2(g) + 3\text{H}_2\text{O}(g)$
- (c) Ammonia is a highly soluble gas and dissolve, in the water.
- (d) Lead (II) Oxide, PbO .
9. (a) The industrial manufacture of nitric acid is by the Ostwald process.
- Ammonia (NH_3) is mixed with air; the two are cleaned and heated in a heat exchanger.
 - The pure gas mixture is conveyed to a hot Platinum alloy catalyst in a catalytic chamber, where ammonia and Oxygen combine to form Nitrogen oxide and water.

$$4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g)$$
 - In the heat exchanger, more air oxidizes Nitrogen oxide to Nitrogen dioxide

$$2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$$
 - The Nitrogen dioxide NO_2 is mixed with excess air and conveyed to towers, down which hot water flows. The combination of nitrogen dioxide and air and hot water gives nitric acid.

$$4\text{NO}_2(g) + \text{O}_2(g) + 2\text{H}_2\text{O}(l) \rightarrow 4\text{HNO}_3(aq)$$
- (b) When conc HNO_3 is added to copper, a reddish brown nitrogen dioxide and a green solution of copper (II) nitrate are formed.
The reaction between copper and conc HNO_3 is vigorous. Copper is oxidized by losing two electrons to conc HNO_3 . the latter is reduced to NO_2 , NO_3^- and H_2O .

$$\text{Cu}(s) \rightarrow \text{Cu}^{2+}(aq) + 2e$$

$$2\text{HNO}_3(aq) + e \rightarrow \text{NO}_2(g) + \text{NO}_3^-(aq) + \text{H}_2\text{O}(l)$$
- (c) **Confirmatory test for NO_3^- ion**
To a solution of a nitrate, add an equal volume of freshly prepared Iron (II) sulphate solution.

Slope the test tube and carefully pour conc H_2SO_4 down the side. The dense acid sinks to the bottom. A brown ring forms where the two layers meet.

This confirms the presence of a NO_3^- ion.

OR

To a solution of a nitrate, add excess NaOH solution followed by Aluminium or Zinc powder or Devarda's alloy (Alloy of zinc and aluminium). A vigorous reaction takes place, effervescence of a colourless alkaline gas (NH_3) confirms the presence of NO_3^- ion.

CHAPTER SIX

SULPHUR AND ITS COMPOUNDS

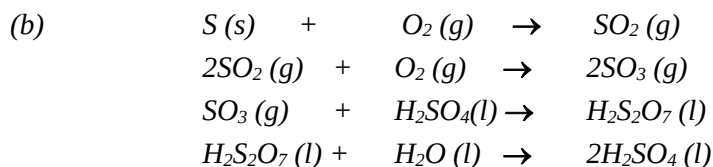
SELF-CHECK 6.1

1. B	2. C.	3. D.	4. A.	5. B.
6. B.	7. B.	8. A.	9. C.	10. D.

SELF-CHECK 6.2

1. (a) See diagram on page 142

- In the Frasch process for the extraction of sulphur, 3 concentric pipes are sunk to the sulphur deposit.
- Super heated water at $170^\circ C$ is pumped down through the outer pipe while hot compressed air is pumped down through the central pipe.
- As a result of froth of air, water and molten sulphur is forced up through the middle pipe and is led off to tanks where sulphur separates out.



(c) When concentrated sulphuric acid is added to sugar, heat is evolved; the acid turns sugar into black solid mass of carbon.

(d) Any two of the following:
Manufacture drugs and skin ointments.
Production of:

- Matches,
- Vulcanized rubber,
- Dyes, and
- Explosives e.g Gunpowder

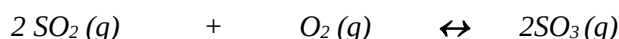
2. (a) $\text{SO}_3^{2-}(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{SO}_2(\text{g})$
- (b) (i) Potassium dichromate or acidified potassium permanganate solution.
- (ii) - **With potassium dichromate**
The solution turns from orange to green.
- (ii) - **With acidified potassium permanganate solution.**
The purple colour of the solution becomes colourless.
- (c) (i) The red colour of the flower becomes colourless.
(ii) The bleaching property of the sulphur dioxide bleaches the colour.

3. (a) (i) Sodium sulphite.
(ii) The hydrochloric acid must be dilute.
The mixture of the dilute hydrochloric acid and sodium sulphite must be heated.
(iii) Concentrated sulphuric acid.
(iv) $\text{Na}_2\text{SO}_3(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow 2 \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{SO}_2(\text{g})$

- (b) (i) The solution turns from orange to green.
(ii) The colour of the dye turns to colourless.

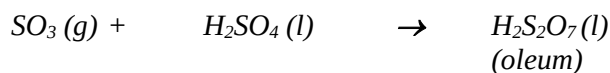
(c) **Oxidation of sulphur to sulphur dioxide by air**

The sulphur dioxide is mixed with air. The mixture is dried and passed over vanadium (V) oxide catalyst heated at $450^\circ\text{C} - 500^\circ\text{C}$ and pressure of 1 atmosphere. The sulphur dioxide is oxidized to sulphur trioxide.



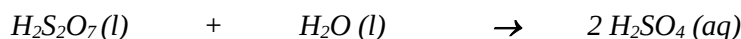
Absorption of Sulphur trioxide

The sulphur trioxide formed is absorbed into concentrated sulphuric acid forming a fuming liquid called 'oleum'.



Absorption, cooling and Dilution

The oleum formed is absorbed, cooled and carefully diluted with the correct amount of water to form ordinary, sulphuric acid which is 98%.



4. (a) Heat is evolved, the sugar is changed into a black porous mass of carbon.
This is an exothermic reaction in which concentrated sulphuric acid dehydrates the sugar i.e. removes the elements of water from the sugar.
- $$\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s}) \rightarrow 12 \text{C}(\text{s}) + 11 \text{H}_2\text{O}(\text{g})$$
- (b) The two react with effervescence of a colourless and weakly acid gas that turns lime water milky; the resultant solution is colourless
- $$\text{ZnCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$$

- (c) – The sulphuric acid must be concentrated
 - The mixture must be heated.
- $$\text{Cu (s)} + 2\text{H}_2\text{SO}_4 \text{ (l)} \rightarrow \text{CuSO}_4 \text{ (s)} + 2\text{H}_2\text{O (l)} + \text{SO}_2 \text{ (g)}$$
- (d) By use of acidified Barium nitrate or Barium chloride solution. When a little Barium nitrate or Barium chloride is added to sulphuric acid, a white precipitate is formed which is insoluble in the acid.
- (e) - In the manufacture of fertilizers e.g. $(\text{NH}_4)_2\text{SO}_4$
 - In the production of synthetic fibres and plastics
 - In car batteries
 - To make detergents



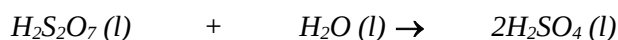
- (b) Step II is necessary to remove any dust which would “poison” the catalyst i.e. make it inactive and also to purify the two gases SO_2 and O_2

- (c) (i) Drying agent in step III – Concentrated sulphuric acid.

- (ii) Catalyst in step IV: – Vanadium pentaoxide (V_2O_5)

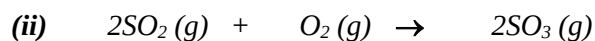
- (d) The sulphur trioxide formed is absorbed or passed into Concentrated sulphuric acid to form fuming sulphuric acid ($\text{H}_2\text{S}_2\text{O}_7$) – oleum.

Oleum is then diluted with water to form the ordinary conc. sulphuric acid.



- (e) (i) Three conditions for the maximum yield of sulphur trioxide.

- Low temperature (450°C)
- High pressure
- Catalyst of Vanadium pentaoxide (V_2O_5)



From the equation, 2 moles of sulphur dioxide produces 2 moles of sulphur trioxide, the volume of sulphur trioxide produced by 20 cm^3 of sulphur dioxide

$$= \frac{2 \times 20}{2}$$

$$= 20 \text{ cm}^3$$

- (f) Sulphuric acid is used in the manufacture of:

- Fertilizers
- Explosives
- drugs
- dyes
- detergents
- plastics

$$(g) \quad \text{Mols of dilute } H_2SO_4 = \frac{2 \times 2000}{1000} = 4 \text{ moles}$$

18 moles are in 1000cm^3

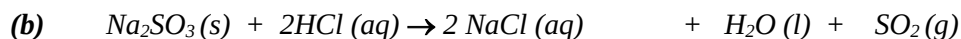
$$1 \text{ mole is in } \frac{1000}{18} \text{cm}^3$$

$$4 \text{ moles are in } \frac{1000 \times 4}{18} = 222.2\text{cm}^3$$

$\therefore$ Required volume of sulphuric acid = 222.2 cm^3

6. (a) Sulphur dioxide is prepared in the laboratory by heating a mixture of dilute hydrochloric acid and sodium sulphite in a round bottom flask and the gas is collected by downward delivery method. If the gas is required dry, it is passed through concentrated sulphuric acid.

See diagram on page 144



(c) The presence of sulphur dioxide can be tested by as follows:

- (i) A filter paper soaked in acidified solution of potassium dichromate is placed in a stream of sulphur dioxide.

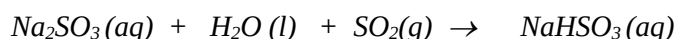
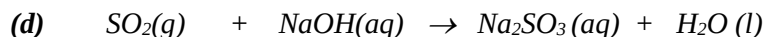
Observation:

The solution turns from orange to green.

- (ii) Sulphur dioxide is bubbled through acidified potassium permanganate solution.

Observation:

The purple colour of the solution becomes colourless.



(e) The reaction equation



1000 cm^3 of sodium hydrogen carbonate solution contains 0.1 moles.

1.00 cm^3 of sodium hydrogen carbonate solution contains $\frac{1 \times 0.1}{1000}$ moles.

25.0 cm^3 of sodium hydrogen carbonate solution contains $\frac{25.0 \times 0.1}{1000}$
= 0.025 moles

From the mole ratio 1 mole of acid : 2 moles of base,

The number of moles of sulphuric acid that reacted with sodium hydrogen carbonate

= $\frac{1}{2}$ the number of moles of the base

= $\frac{1}{2} \times 0.025$

= 0.0125 moles

27.8 cm³ of sulphuric acid contains 0.0125 moles of the acid.

1.0 cm³ of sulphuric acid contains $\frac{0.0125}{27.80}$ moles.

1000 cm³ of sulphuric acid contains $\frac{0.0125 \times 1000}{27.80} = 0.45 \text{ M}$

CHAPTER SEVEN

CHLORINE AND ITS COMPOUNDS

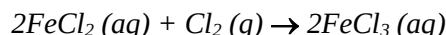
SELF-CHECK 7.1

1.	B.	2.	C.	3.	D.	4.	A.	5.	B.
6.	A.	7.	C.	8.	D.	9.	D.		

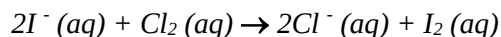
SELF-CHECK 7.2

1. (a) (i) A. Potassium Permanganate
B. Water
C. Concentrated Sulphuric acid
(ii) To remove unchanged hydrogen chloride.
(b) The hydrochloric acid must be concentrated.
(c) $2\text{KMnO}_4(\text{s}) + 16\text{HCl}(\text{aq}) \rightarrow 2\text{KCl}(\text{aq}) + 2\text{MnCl}_2(\text{aq}) + 8\text{H}_2\text{O}(\text{l}) + 5\text{Cl}_2(\text{g})$
2. (a) (i) Potassium permanganate
(ii) $2\text{KMnO}_4(\text{s}) + 16\text{HCl}(\text{aq}) \rightarrow 2\text{KCl}(\text{aq}) + 2\text{MnCl}(\text{aq}) + 8\text{H}_2\text{O}(\text{l}) + 5\text{Cl}_2(\text{g})$
(b) (i) A pale yellow solution is observed.
(ii) $\text{Cl}_2(\text{g}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{NaOCl}(\text{aq}) + \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
3. (a) (i) **See diagram on page 154 or 155**
- Chlorine is prepared in the laboratory by dropping conc. hydrochloric acid on to Potassium manganate crystals in a flat bottom flask.
- The chlorine produced is passed through water to remove any unchanged hydrogen chloride, and then through concentrated sulphuric acid to dry it. The gas is then collected by downward delivery.
(ii) $2\text{KMnO}_4(\text{s}) + 16\text{HCl}(\text{aq}) \rightarrow 2\text{KCl}(\text{aq}) + 2\text{MnCl}_2(\text{aq}) + 8\text{H}_2\text{O}(\text{l}) + 5\text{Cl}_2(\text{g})$
(iii) – As a bleaching agent in the pulp and textile industries
- In the treatment of sewerage and purification of water
- In the manufacture of plastics e.g. P.V.C. (Polyvinyl chloride)

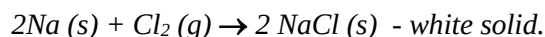
- (b) (i) The green solution of Iron (II) chloride turns to yellow/ brown solution due to oxidation of Iron (II) chloride to Iron (III) chloride solution by chlorine.



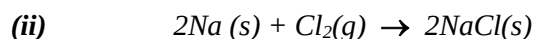
- (ii) The colourless potassium iodide solution turns to purple/ violet due to the displacement of I^- ions by chlorine to give Iodine.



- (c) (i) Sodium continues to burn in chlorine, producing white fumes of sodium chloride.



The reaction is vigorous because both sodium and chlorine are very reactive elements with sodium being a strong reducing agent while chlorine is a strong oxidizing agent.

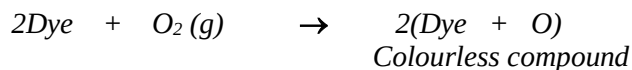


4. (a) - To remove fumes of hydrogen chloride.
- To dry the gas.

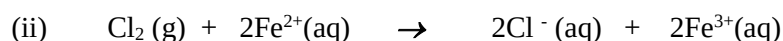
- (b) (i) $\text{Cl}_2(\text{g}) + \text{H}_2\text{O}(\text{aq}) \rightleftharpoons \text{HOCl}(\text{aq}) + \text{HCl}(\text{aq})$
(ii) Chlorine dissolved in water is a bleaching agent. The bleaching action of chlorine is due to the ion OCl^- in the hypochlorous acid. The hypochlorous acid is a very reactive compound and readily gives up its oxygen.



The oxygen given off reacts with the dye, to form a colourless compound.



- (c) (i) The green solution of iron (II) sulphate would turn to yellow in colour due to the formation of iron (III) chloride solution.



5. (a) (i) Chlorine.
(ii) $2\text{PbO}(\text{s}) + 8\text{HCl}(\text{aq}) \rightarrow 2\text{PbCl}_2(\text{aq}) + 4\text{H}_2\text{O}(\text{l}) + 2\text{Cl}_2(\text{g})$

- (b) (i) $2\text{KBr}(\text{aq}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{KCl}(\text{aq}) + \text{Br}_2(\text{aq})$

(ii) Red colour develops much faster than in (b) (i).



6. (a) $\text{MnO}_2(\text{s}) + 4\text{HCl}(\text{aq}) \rightarrow \text{MnCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) + \text{Cl}_2(\text{g})$

- (b) (i) $2\text{H}_2\text{O}(\text{l}) + \text{Cl}_2(\text{g}) \rightarrow \text{MnCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) + \text{Cl}_2(\text{g})$

(ii) The Chloric (I) acid, HClO formed is decomposed by sunlight, giving off oxygen.



- (c) (i) $2\text{KOH}(\text{aq}) + \text{Cl}_2(\text{g}) \rightarrow \text{KClO}(\text{aq}) + \text{KCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 Or $2\text{OH}^-(\text{aq}) + \text{Cl}_2(\text{g}) \rightarrow \text{ClO}^-(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 (ii) Potassium chlorate (I) (Hypochlorite); potassium chloride.
 (iii) It is used as a mild antiseptic.

SELF-CHECK 7.3

7. (a) $\text{H}_2\text{SO}_4(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{NaHSO}_4(\text{aq}) + \text{HCl}(\text{g})$
 (b) See diagram on page 159
 (c) (i) Effervescence of a colourless gas that turns lime water milky would be given off.

$$\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$$

 (ii) White precipitate would be formed

$$\text{AgNO}_3(\text{s}) + \text{HCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{HNO}_3(\text{aq})$$

 (iii) Effervescence of a colourless gas that burns with a pop sound would be evolved.

$$\text{Mg}(\text{s}) + 2\text{HCl}(\text{s}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(\text{g})$$

 (d) Copper is below hydrogen in the reactivity series. So can not displace hydrogen from hydrogen chloride.
8. (a) A dry sample of hydrogen chloride can be prepared by action of conc. Sulphuric acid on rock salt (impure sodium chloride).
 The gas hydrogen chloride gas is dried by passing it through concentrated Sulphuric acid and then collected by downward delivery method.
- Equation:** $\text{H}_2\text{SO}_4(\text{l}) + \text{NaCl}(\text{s}) \rightarrow \text{NaHSO}_4(\text{s}) + \text{HCl}(\text{g})$
- (b) Hydrochloric acid
- (c) (i) Acidified Silver nitrate solution
 (ii) A white precipitate is formed insoluble in the acid.
- (d) $\text{H}^+(\text{aq}) + \text{HCO}_3^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
- (e) $\text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) \rightarrow \text{PbCl}_2(\text{s})$
 Moles of $\text{Pb}(\text{NO}_3)_2$ that reacted = $\frac{25 \times 0.2}{1000} = 0.005$
 Moles of PbCl_2 formed = 0.005 since the mole ratio $\text{Pb}(\text{NO}_3)_2 : \text{PbCl}_2$

$$= 1 : 1$$

 R.f.m of $\text{PbCl}_2 = 207 + (35.5 \times 2) = 278$
 Mass of PbCl_2 formed = $278 \times 0.005 = 1.39\text{g}$
9. (a) See diagram on page 161
 (b) $\text{Fe}(\text{s}) + 2\text{HCl}(\text{g}) \rightarrow \text{FeCl}_2(\text{s}) + \text{H}_2(\text{g})$

- (c) (i) When aqueous sodium hydroxide was added to a solution of Iron (II) chloride, a green precipitate is formed which is insoluble in excess sodium hydroxide solution.
- (ii) $\text{FeCl}_2(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s}) + 2\text{NaCl}(\text{aq})$
- Or $\text{Fe}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s})$
- (d) (i) The green solution turned to a yellow solution when chlorine was passed through Iron (II) chloride solution because of the oxidation of Fe^{2+} (green) to Fe^{3+} (yellow/brown)
- (ii) **Equation:** $2\text{Fe}^{2+}(\text{aq}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{Fe}^{3+}(\text{aq}) + 2\text{Cl}^-(\text{aq})$
- (e) (i) To test for Fe^{3+} ions, use potassium hexacyanoferrate (II) – a dark blue precipitate is formed or use sodium hydroxide solution- a brown precipitate is formed.
- (ii) To test for Cl^- ions, use acidified silver nitrate solution – a white precipitate is formed, soluble in aqueous ammonia.

CHAPTER EIGHT

ION CHEMISTRY

SELF-CHECK 8.1

1. B.	2. D.	3. B.	4. B.	5. D.
6. A.	7. C.	8. B.	9. A.	10. A.
11. C.	12. C.	13. B.	14. C.	15. C.
16. C.	17. A.	18. B.	19. B.	20. B.
21. B.	22. B.	23. C	24. B.	25. C.

SELF-CHECK 8.2

1. (a) Copper (II) carbonate
- (b) (i) - Black solid formed
- A colourless gas that turns lime water milky was evolved.
- (ii) $\text{CuCO}_3(\text{s}) \rightarrow \text{CuO}(\text{s}) + \text{CO}_2(\text{g})$
- (c) (i) Barium nitrate solution followed by dilute nitric acid.
- (ii) $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$
2. (a) The hydroxide is highly soluble.
- (b) (i) HCO_3^- , CO_3^{2-} (ii) HCO_3^-
- (c) (i) MgCO_3



3. (a) Dilute sulphuric acid.

With: Pb^{2+} - White precipitate,
 Al^{3+} - No observable change.

(b) Dilute nitric acid:

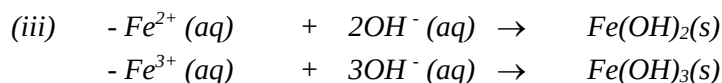
With: SO_4^{2-} - No observable change.
 CO_3^{2+} - A colourless gas that turns lime water milky evolved.

4. (a) (i) Green. (ii) Brown.

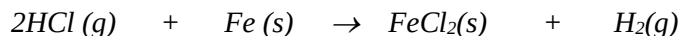
(b) (i) $FeSO_4$. (ii) $Fe_2(SO_4)_3$.

(c) (i) Dilute sodium hydroxide solution.

(ii) With: - Fe^{2+} - Green precipitate formed.
- Fe^{3+} - Brown precipitate formed.



(d) Pass dry hydrogen chloride over heated iron wool in a combustion tube.



5. (i) Pb^{2+} ions (ii) $Pb^{2+}(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s)$
(ii) Lead (II) hydroxide forms white precipitate which is insoluble in excess ammonium hydroxide solution.

6. (a) (i) Ammonium hydroxide:

With: Pb^{2+} ion - White precipitate insoluble in excess ammonium hydroxide.
 Zn^{2+} ion - White precipitate soluble in excess ammonium hydroxide forming a colourless solution.

(ii) Silver nitrate:

With: CO_3^{2-} ion: - No observable change.
 Cl^- ion: - White precipitate which turns to grey on standing.

CHAPTER NINE

OXIDATION AND REDUCTION

SELF-CHECK 9.1

1.	C.	2.	A.	3.	B.	4.	C.
5.	B.	6.	A.	7.	C	8.	C

SELF-CHECK 9.2

7. (a) Oxidation is the loss of electrons
- (b) (i) $H^+ + 2e^- \rightarrow H_2$ Reduction
- (ii) $Fe^{2+} - e^- \rightarrow Fe^{3+}$ Oxidation
- (iii) $Cl_2 - 2e^- \rightarrow 2Cl^-$ Oxidation
- (c) (i) the elements on the right hand of the periodic table are good oxidizing agents.

Second period - Oxygen and fluorine

Third period - Chlorine

(ii) The elements on the left hand of the periodic table are good reducing agents.

Second period - lithium

Third period - sodium and magnesium.

(iii) Good oxidizing agents have almost filled valence orbitals; Small atomic radius (less electropositive element).

(iv) Good reducing agents have a few electrons more than an inert gas electronic configuration; greater atomic radius (more electro positive element)

CHAPTER TEN

ENERGY CHANGES IN CHEMICAL REACTIONS

SELF-CHECK 10.1

1. A.	2. B.	3. D.	4. C.	5. A.
6. C.	7. A.	8. D.	9. A.	10. C.
11. B.	12. D.	13. D.	14. B.	15. A.

SELF-CHECK 10.2

1. (a) Enthalpy of combustion is the enthalpy change that occurs when one mole of a substance is completely burnt in (pure) oxygen.
- (b) **Apparatus:**
Retort stand/clamp, a thin walled tin can, thermometer, water, ethanol, and a spirit lamp.

Procedure:

- Weigh a spirit burner containing pure propanol.
- Fill a thin-walled tin can with a known volume of water.
- Clamp the tin can with its content above the lamp as shown in the diagram below.
- Insert the thermometer into the water.
- Read and record the initial temperature of the water.

See diagram on page 193

- Heat the water while carefully stirring with the thermometer.
- Reweigh the lamp and its content.

Results:

Volume of cold water	=	V_1	cm^3
Initial temperature of water	=	t_1	$^{\circ}\text{C}$
Final temperature of water	=	t_2	$^{\circ}\text{C}$
Mass of spirit burner before burning	=	m_1	g
Mass of spirit burner after burning	=	m_2	g
Specific heat capacity of water	=	4.2	$\text{J/g } ^{\circ}\text{C}$

Calculation

$$\begin{aligned}
 \text{Heat gained by water} &= mc\Delta t \\
 &= \left(\text{Mass of water} \right) \times \left(\text{Sp. ht. cap of water} \right) \times \left(\text{Change in Temperature} \right) \\
 &= (\text{Vol.} \times \text{Density}) \times 4.2 \times (t_2 - t_1) \\
 &= (\text{Vol.} \times 4.2 \times (t_2 - t_1) \text{ Joules}
 \end{aligned}$$

Heat lost/ produced by propanol is calculated as follows:

The relative molecular mass of ethanol, $\text{C}_3\text{H}_7\text{OH}$,

$$\begin{aligned}
 &= (3 \times 12) + (1 \times 8) + 16 \\
 &= 36 + 8 + 16 \\
 &= 60 \text{ g}
 \end{aligned}$$

$$\begin{aligned}
 \text{Mass of ethanol used up (burnt)} &= \left(\text{Mass of spirit lamp before burning} \right) - \left(\text{Mass of spirit lamp after burning} \right) \\
 &= (m_1 - m_2) \text{ g}
 \end{aligned}$$

Assuming heat gained by water = Heat given out by ethanol

$(m_1 - m_2) \text{ g}$ of ethanol produces, $V\rho \times 4.2 \times (t_2 - t_1)$

$$60 \text{ g of ethanol produce will produce} = \frac{60}{m_1 - m_2} \times (V \times 4.2) \times (t_2 - t_1) \text{ KJ mol}^{-1}$$

- (c) Mass of propanol burnt = 0.54 g, $\Delta t = 21.5 ^{\circ}\text{C}$, Volume of water = 150 cm^3
 Density of water = 1 g/cm^3 , sp. ht. cap of water = $4.2 \text{ Jg}^{-1}\text{K}^{-1}$

$$\begin{aligned}
 \text{Heat energy gained by water} &= mc\Delta t \\
 &= V\rho \times 4.2 \times (t_2 - t_1) \\
 &= 150 \times 1 \times 4.2 \times 21.5 \\
 &= 13545 \text{ J} \\
 &= 13.545 \text{ KJ}
 \end{aligned}$$

0.54 g of propanol produces 13.545 KJ of heat energy.

1 g of propanol produces $\frac{13.545}{0.54}$ KJ of heat energy.

60 g (1 mole) of propanol will produce $\frac{60}{0.54} \times 13.545 = 1505 \text{ KJ mol}^{-1}$

2. (a) Enthalpy of neutralization is the enthalpy change that occurs when an acid reacts with an alkali to produce a salt and a mole of water.

(b) Final volume of solution = Vol. of acid + Vol. of base
 $= (50 + 50) \text{ cm}^3$
 $= 100 \text{ cm}^3$

Mass of solution = Volume $\times$ density
 $= 100 \times 1$
 $= 100 \text{ g}$

Change in temperature, Δt = (Final - Initial) temperature
 $= 30.8 - 27.5$
 $= 3.3^\circ\text{C}$

Given the specific heat capacity of water is 4.2 Jg^{-1} ,

Heat evolved = Heat gained by the water formed
 $= \text{Mass} \times \text{sp. heat cap} \times \text{temperature rise}$
 $= 100 \times 4.2 \times 3.3 \text{ J}$

$\therefore$ Heat evolved = **1386 J**

Moles of hydrochloric acid = Molarity $\times$ volume in litres
 $= \frac{0.5 \times 50}{1000}$
 $= 0.025 \text{ moles of acid}$

From $\text{KOH(aq)} + \text{HCl(aq)} \rightarrow \text{KCl(aq)} + \text{H}_2\text{O(l)}$

1 mole of acid produces 1 mole of water, therefore, the number of moles of water produced
 $= 0.025$

0.025 moles of water is produced with the evolution of **1386 J** of heat energy.

$\therefore$ 1 mole of water will be produced with the evolution of $\frac{1386 \times 1}{0.025}$
 $= 55440 \text{ J}$
 $= 55.44 \text{ KJ mol}^{-1}$.

CHAPTER ELEVEN

EXTRACTION OF METALS

SELF-CHECK 11.1

1.	B.	2.	C.	3.	A.	4.	C.
5.	A.	6.	C.	7.	C.	8.	D.

SELF-CHECK 11.2

1. (a) (i) Sodium - Sodium Chloride (NaCl) and
Iron - Hematite (Fe₂O₃)

(ii) Extraction of Sodium by electrolysis

Sodium is extracted by the electrolysis of molten sodium chloride in the Downs Cell. The cell consists of a cylindrical steel cathode and graphite anode.

Sodium chloride is heated and melts at 801 °C. Calcium chloride is added as an impurity to lower the melting point to about 600 °C. A very high current is passed through the electrolyte to keep it in molten state.

Ions present: From NaCl (s) → Na⁺(l) + Cl⁻(l)

From CaCl₂ (s) → Ca²⁺(l) + 2Cl⁻(l)

Equations at the electrodes.

At the anode: 2Cl⁻(l) → Cl₂(g) + 2e⁻

Or 2Cl⁻(l) - 2e⁻ → Cl₂(g)

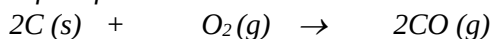
At the cathode: Na⁺(l) + e⁻ → Na (l)

The liquid Sodium is collected under dry nitrogen gas. This prevents the metal from reacting with the atmosphere.

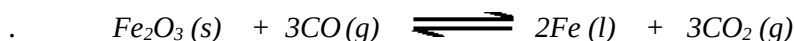
Extraction of Iron

By reduction of its ore hematite (Fe₂O₃) in a blast furnace by carbon monoxide
A mixture of iron ore (concentrated, crushed and roasted), coke and limestone is fed into a blast furnace from the top of the furnace.

A blast of preheated air is injected into the furnace through the pipes called tuyeres at the lower part of the furnace. The air oxidizes the hot coke to carbon monoxide.

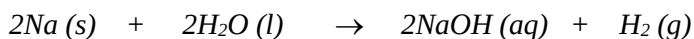


At the upper part of the furnace, the carbon monoxide formed reduces the iron (II) oxide to iron according to:

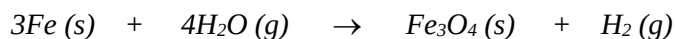


The iron absorbs carbon from the coke, its melting point is lowered. It melts and then sinks and collects at the bottom of the furnace, where it is tapped off and solidified into blocks called “pig iron” or cast iron.

- (b) Sodium reacts with water at ordinary room temperature.



Iron does not react with cold water, but reacts with steam.



2. (a) Sodium chloride (NaCl)

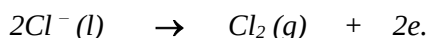
(b) **Extraction of sodium**

Sodium is extracted by electrolysis of molten or fused sodium chloride in the Down's cell using of steel cathode and carbon anode.

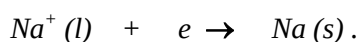
A very high current is passed to keep the electrolyte in liquid state.

– Calcium chloride is added to lower the melting point of sodium chloride.

– Chlorine gas is liberated at the anode as a byproduct.



Molten sodium is produced at the cathode.



– The molten sodium is collected under nitrogen to prevent it from reacting with air.

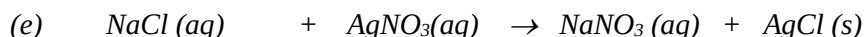
(c) Source of the raw material,

Availability of: - Power,

- Market and

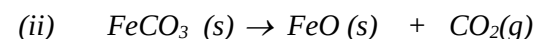
- labour

(d) (i) The greenish colour of chlorine disappears and
A white solid formed.



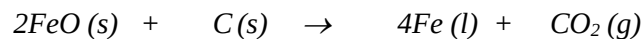
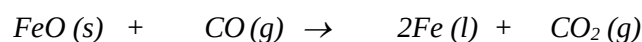
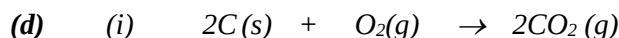
3. (a) Silicon (IV) oxide.

(b) (i) Siderite (Iron (II)carbonate.



(c) (i) Iron ore, lime and coke.

(ii) A blast of hot air

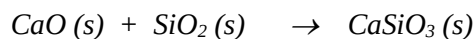


(ii) **Remember that the role of lime stone is to remove the main impurity.**

It first decomposes to calcium oxide (quick lime CaO) according to:



The calcium oxide then combines with Silicon (IV) oxide to form Calcium silicate (CaSiO₃), the slag according to the equation:

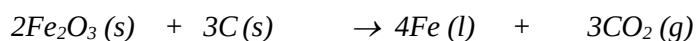
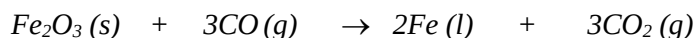


(e) *Iron and Carbon.*

4. (a) *Magnetite - Fe₃O₄*

(b) *To remove the main impurity, silicon (IV) oxide.*

(c) *Carbon monoxide and carbon reduce the iron (II) oxide to iron at temperature ranges of (500 – 800) °C according to the equations:*



(d) (i) *A colourless gas that burns with a pop sound will be given off and a green solution is formed.*

(ii) *A black solid, iron (II) oxide, will be formed and a colourless gas that burns with a pop sound will be given off.*

