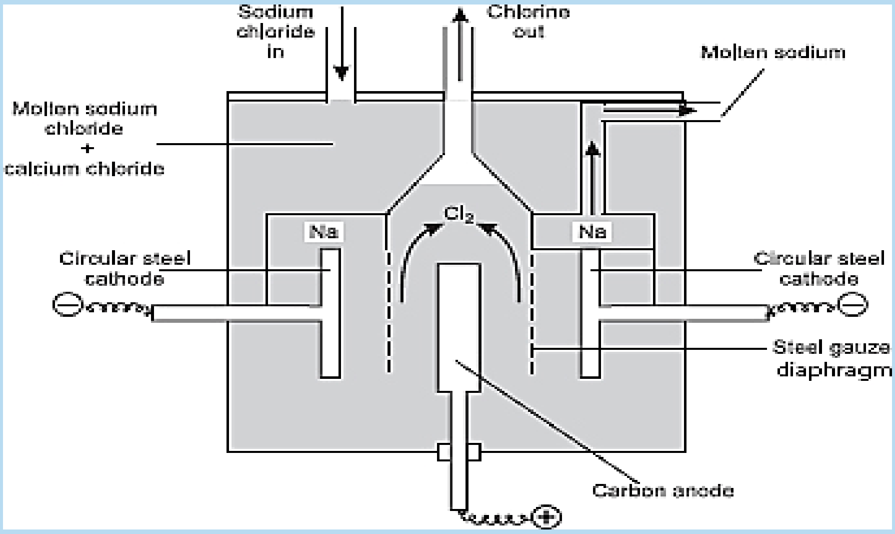
233

FORM FOUR CHEMISTRY HANDBOOK

With well drawn diagrams, solved examples and questions for exercise

(REVISED EDITION)



TR. BENARD CLAIN

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Table of Contents

Acknowledgement

First and foremost I thank the Almighty God for the gift of life, energy, knowledge and skills to pursue this work.

I am very grateful to the entire Swafaa Academy Secondary school fraternity for generously supporting me all round as I worked on this material. I must specifically appreciate the H.O.D Sciences Swafaa Academy Mr. Kipchumba Arusei for the support and encouragement.

The support by Muyeye Secondary School Science department members, Mr. Simiyu, Mr. Wafubwa and Mr. Ambrose of Physics, Mr. Cidini of Chem, Madam Irenel, Jane and Ruthl of Chemistry must be appreciated.

The care and best wishes I received from my mother Hellen Clain and my siblings deserve special attention. They were a great source of encouragement.

Lines that influence activities in my life

1. *God is always there to assist provided you ask for Him.*
2. *At its best, Chemistry eliminates complexity by revealing underlying simplicity.*
3. *There is no method of changing your fate except through hard work.*
4. *Cohesion with immediate neighbours and determination always betters your immediate environment.*

Brief Personal Profile

Benard Clain A. is an experienced teacher of Physics and Chemistry. He is a First Class Honors BSC graduate from Masinde Muliro University Of Science And Technology with an additional PGDE from the said University. Currently taking masters degree. He also has profound knowledge in computer applications and graphics.

TOPC ONE

ACIDS, BASES AND SALTS.

Table of Contents

Organizer

Objectives

By the end of the topic, the learner should be able to:

(a) Define acids, bases and salts.
(b) Explain the difference between a strong and weak acid and a strong and a weak base.
(c) Explain the role of a solvent in acid base character.
(d) Compare the strength of acids and bases in terms of pH and electrical conductivity.
(e) Write formula and ionic equations for acid-base and precipitation reactions.
(f) Identify the complex ions formed in specified cation-anion reactions.
(g) Define solubility and explain the use of solubility curves in salt extraction.
(h) State the types and causes of hardness of water and explain the methods of removing hardness.
(i) State some effects of hard water.

ACIDS, BASES AND SALTS.

Acids

An acid can be defined as *a substance that dissociates in water to give hydrogen ions as the only positively charged ions.*

Equations

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

$$H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$$

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

Strength of Acids

The pH value of a solution is related to the number of hydrogen ions present. The higher the number of the hydrogen ions the lower the pH value.

Acids whose aqueous solutions contain a high number of hydrogen ions hence lower pH values are called **strong acids** while those whose solutions have few hydrogen ions hence higher pH values are **weak acids**.

A strong acid dissociates completely and yields many hydrogen ions in aqueous solution.

A weak acid dissociates partially and yields few hydrogen ions in aqueous solution.

Examples of strong acids include **hydrochloric acid, sulphuric(VI) acid (H_2SO_4) and nitric(V) acid (HNO_3).**

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

$$H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$$

$$HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$$

The strong acids have **lower pH values** and are also **better conductors of electricity since conductivity is dependent on the number of ions present in the electrolyte.**

Weak acids have **higher pH values** and are **poor conductors** of electricity in their electrolytic state. They include **ethanoic acid, ethanedioic acid ($H_2C_2O_4$) and methanoic acid (HCOOH).**

$$CH_3COOH(aq) \longrightarrow H^+(aq) + CH_3COO^-(aq)$$

Bases

A **base** is defined as *a substance that dissociates in water to produce hydroxide* (OH^-) *as the only negatively charged ions.*

Strength of Bases.

A **strong base** is one that **ionises completely** in water to produce hydroxide ions. Strong bases have **high pH values** and are **good electrolytes,** eg **sodium hydroxide.**

When sodium hydroxide is dissolved in water, it completely ionises to form sodium and hydroxide ions.

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

Sodium hydroxide solution therefore contains many hydroxide ions.

Weak bases partially ionises to form hydroxide ions. They therefore have **lower pH values** and are **poor electrolytes,** eg **Ammonia solution.**

When ammonia gas dissolves in water, some of the ammonia molecules react with the water molecules to produce ammonium hydroxide.

$$NH_3(g) + H_2O(l) \longrightarrow NH_4OH\ (aq)$$

The ammonium hydroxide partially ionises to form ammonium and hydroxide ions.

$$NH_4OH(aq) \rightleftharpoons NH_4^+\ (aq) + OH^-\ (aq)$$

Ammonia solution therefore contains fewer hydroxide ions compared to sodium hydroxide solution.

Hydroxide ions are responsible for the alkaline properties. **The higher the number of hydroxide ions, the higher the pH value and the greater the conductivity.** Therefore, sodium hydroxide is a stronger alkali than ammonia solution.

Acid-base Reactions

Acids react with bases to form a salt and water.

(i) $NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$

$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

(ii) $CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$

$CuO\ (s) + 2H^+(aq) \longrightarrow Cu^{2+}(aq) + H_2O(l)$

(iii) $Ca(OH)_2(aq) + 2HNO_3\ (aq) \longrightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$

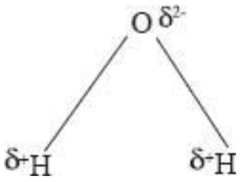
$2OH^-(aq) + 2H^+(aq) \longrightarrow 2H_2O(l)$

During the reactions, the hydrogen ions from the acid react with the hydroxide ions from the base to form water.

Effect of a Solvent on a Solute

A molecule which behaves as if it is negatively charged in one end and positively charged in the opposite end is said to be **polar**. Water is therefore a **polar solvent**.

A water molecule consists of one oxygen atom and two hydrogen atoms. Since the oxygen has a higher ability to attract bond electrons compared to hydrogen, the bond pairs of electrons are closer to the oxygen atom in the water molecule leading to the oxygen atom developing a partial negative charge (δ^-) while the hydrogen atoms develop a partial positive charge ($\delta+$).



The hydrogen chloride molecule is also polar. This is because chlorine has a higher ability to attract bond electrons compared to hydrogen. The chlorine end of the molecule develops a partial negative charge while the hydrogen end develops a partial positive charge.

$$^{\delta+}H - Cl^{\delta-}$$

When hydrogen chloride is dissolved in water, different water molecules attract both ends of the hydrogen chloride molecule and cause it to dissociate. This implies that the covalent bond is broken in such a way that the shared pair of electrons are left with the chlorine atom, thus creating a negatively charged chloride ion and a positively charged hydrogen ion.

Hydrogen chloride dissociates in water as follows;

$$HCl\ (s) \xrightarrow{water} H^+\ (aq) + Cl^-\ (aq)$$

The presence of the hydrogen ions in an aqueous solution of hydrogen chloride explains why it is acidic, i.e., turns blue litmus paper red and reacts with magnesium to liberate carbon(IV)oxide gas.

$$2H^+(aq) + Mg(s) \longrightarrow Mg^{2+}(aq) + H_2(g)$$

$$2H^+(aq) + CaCO_3(s) \longrightarrow Ca^{2+}(aq) + CO_2(g) + HO(l)$$

Methylbenzene is a non-polar solvent. When hydrogen chloride gas dissolves in methylbenzene, its molecules do not dissociate and therefore no hydrogen ions are present. Hence a solution of hydrogen chloride in methylbenzene does not exhibit acid properties. It has no effect on litmus paper and does not react with magnesium or calcium carbonate.

Effect of Polar and non-polar solvents: ammonia in water and in methylbenzene

Dry ammonia has no effect on dry litmus paper. A solution of ammonia in water turns red litmus paper blue while a solution of ammonia gas in methylbenzene has no effect on dry litmus paper.

When ammonia is dissolved in water, some of the ammonia molecules react with the polar water molecules to form ammonium hydroxide.

$$NH_3(g) + H_2O(l) \longrightarrow NH_4OH(aq)$$

The ammonium hydroxide ionises to form ammonium and hydroxide ions.

$$NH_4OH(aq) \longrightarrow NH_4^+(aq) + OH^-(aq)$$

Thus, the hydroxide ions are the ones that are responsible for the basic nature of the gas. A solution of dry ammonia gas in methylbenzene has no effect on dry litmus paper since no hydroxide ions are produced.

Ammonia gas dissolves in methylbenzene without undergoing any change because methylbenzene is non-polar. Hence there are no hydroxide ions in a solution of ammonia gas in methylbenzene. The solution therefore has no effect on litmus paper.

Amphoteric Oxides and Hydroxides

The metal oxides or hydroxides which react both as bases and acids are said to be **amphoteric.**

The oxides and hydroxides of aluminium, zinc and lead react with nitric(V) and sodium hydroxide to form colourless solutions.

Acids react with metal oxides and hydroxides to form salts and water (**neutralization reactions)** .

Aluminium oxide, zinc oxide, and lead(II) oxide react with the acid to form colourless solutions.

$$Al_2O_3\,(s) + 6HNO_3(aq) \longrightarrow 2Al(NO_3)_3(aq) + 3H_2O(l)$$

$$ZnO(s) + 2HNO_3(aq) \longrightarrow Zn(NO_3)_2(aq) + H_2O(l)$$

$$PbO(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + H_2O(l)$$

In these reactions, metal oxides are reacting as bases.

Aluminium oxide, zinc oxide, and lead(II) oxide also react as acids with sodium hydroxide to form colourless solutions.

$$Al_2O_3(s) + 2NaOH(aq) \longrightarrow 2NaAlO_2(aq) + H_2O(l)$$

$$PbO(s) + 2NaOH(aq) \longrightarrow Na_2PbO_2(aq) + H_2O(l)$$

$$ZnO(s) + 2NaOH(aq) \longrightarrow Na_2ZnO_2(aq) + H_2O(l)$$

Lead hydroxide, zinc hydroxide and aluminium hydroxide are also amphoteric.

$$Pb(OH)_2(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + 2H_2O(l)$$

$$Zn(OH)_2(s) + 2HNO_3(aq) \longrightarrow Zn(NO_3)_2(aq) + 2H_2O(l)$$

$$Al(OH)_3(s) + 3HNO_3(aq) \longrightarrow Al(NO_3)_3(aq) + 3H_2O(l)$$

In this reactions the hydroxides are reacting as bases.

Salts.

A salt is defined as an ionic compound formed when a cation derived from a base combines with an anion derived from an acid.

Precipitation Reactions

A **precipitate** is an insoluble solid that separates out from a solution during a chemical reaction.

Sodium carbonate, potassium carbonate and ammonium carbonate are soluble in water. Their solutions may be used to precipitate insoluble metal carbonates. The ionic equations showing the formation of calcium carbonate, zinc carbonate, Copper(II) carbonate, lead carbonate and barium carbonate are:

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)$$

$$Zn^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow ZnCO_3(s)$$

$$Cu^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CuCO_3(s)$$

$$Pb^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow PbCO_3(s)$$

$$Ba^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow BaCO_3(s)$$

Iron(III) and aluminium salts hydrolyse in water to give acidic solutions which react with carbonates to liberate carbon(IV) oxide gas. This explains the effervescence observed when sodium carbonate is added to salt solutions containing these ions.

Iron(III) and aluminium carbonates do not therefore exist. The precipitates formed in this case are hydroxides.

Test for sulphates, sulphites and chlorides (anions)

Most metal cations form soluble chlorides except lead(II) ions.

$$Pb^{2+}(aq) + 2Cl^{-}(aq) \longrightarrow PbCl_2(s)$$

However, lead(II) chloride is soluble in warm water. All sulphates of the cations listed are insoluble except barium sulphate and lead sulphate.

$$Pb^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s)$$

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$$

The sulphites of all the listed cations are insoluble. However, the sulphites react with dilute hydrochloric acid or nitric acid to produce sulphur(IV) oxide gas.

$$PbSO_3(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + SO_2(g) + H_2O(l)$$

$$BaSO_3(s) + 2HCl(aq) \longrightarrow BaCl_2(aq) + SO_2(g) + H_2O(l)$$

The reaction between a sulphite and hydrochloric or nitric(V) acid is used as a test to distinguish between sulphite and sulphate ions. The sulphite dissolves in the acid with the evolution of sulphur(IV) oxide, while the sulphate does not dissolve in the acid.

$$BaSO_3(s) + 2H^+(aq) \longrightarrow Ba^{2+}(aq) + SO_2(g) + H_2O(l)$$

Test for cations: using aqueous ammonia and sodium hydroxide.

Sodium hydroxide forms insoluble hydroxides with the ions of magnesium, zinc, aluminium, iron(II), iron(III), copper(II) and lead(II). These hydroxides have characteristic appearance which forms the basis of their identification. The ionic equations for the reactions are:

$$Mg^{2+}(aq) + 2OH^-(aq) \longrightarrow Mg((OH)_2(s)\ \textit{white precipitate}$$

$$Zn^{2+}(aq) + 2OH^-(aq) \longrightarrow Zn(OH)_2(s)\ \textit{white precipitate}$$

$$Al^{3+}(aq) + 3OH^-(aq) \longrightarrow Al(OH)_3(s)\ \textit{white precipitate}$$

$$Fe^{2+}(aq) + 2OH^-(aq) \longrightarrow Fe(OH)_2(s)\ \textit{green precipitate}$$

$$Fe^{3+}(aq) + 3OH^-(aq) \longrightarrow Fe(OH)_3(s)\ \textit{brown precipitate}$$

$$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu\ (OH)_2(s)\ \textit{blue precipitate}$$

$$Pb^{2+}(aq) + 2OH^-(aq) \longrightarrow Pb\ (OH)_2(s)\ \textit{white precipitate}$$

Complex Ions

The hydroxides of zinc, aluminium, and lead react with excess sodium hydroxide to form soluble complex ions. Thus:

$$Zn(OH)_2(s) + 2OH^-(aq) \longrightarrow [Zn(OH)_4]^{2-}\ \textit{colourless}$$

$$Al(OH)_3(s) + OH^-(aq) \longrightarrow [Al(OH)_4]^-(aq)\ \textit{colourless}$$

$$Pb(OH)_2(s) + 2OH^-(aq) \longrightarrow [Pb\ (OH)_4]^{2-}(aq)\ \textit{colourless}$$

Potassium hydroxide may be used in place of sodium hydroxide. Similar results are obtained when potassium hydroxide is used.

When ammonia solution is used in place of sodium hydroxide, insoluble metal hydroxides are formed in the same way. However, the hydroxides of zinc and copper dissolve in excess ammonia solution because of the formation of complex ions.

Thus:

$$\underset{white}{Zn(OH)_2(s)} + 4NH_3\ (aq) \longrightarrow \underset{colourless}{[Zn(NH_3)_4]^{2+}(aq)} + 2OH^-(aq)$$

$$\underset{blue}{Cu(OH)_2(s)} + 4NH_3\ (aq) \longrightarrow \underset{deep\ blue}{[Cu(NH_3)_4]^{2+}(aq)} + 2OH^-(aq)$$

It is observed that only zinc hydroxide dissolves in excess of both sodium hydroxide and ammonia solution.

In complex ions the ammonium molecule and the hydroxide groups form dative bonds with the metal ions.

Applications of Complex Ions

Complex ions are important in biology and industry. Chlorophyll and haemoglobin are complex ions.

In industry, complex ions play an important role in softening of hard water as used in ion exchangers and also in the extraction of some metals such as aluminium from their ores.

Solubility and Solubility Curves.

The **solubility** of a solute in a solvent is the maximum mass of solute required to saturate 100 g of the solvent at a particular temperature.

Solubility can also be expressed in moles per litre of solution at a given temperature.

Example

24 cm3 of water dissolves 6 g of a solute at 15°C. Determine the solubility of the solute in g/100g of water.

Since density of water is 1g/cm3, 24 cm3 of water =24g.

24 g of water dissolves 6g of the solute at 15°C.

This means that 100 g of solvent will dissolve:

$$\frac{100 \times 6}{24} = 25 \text{ g of solute.}$$

Thus the solubility of the salt is 25 g per 100 g of water at 15°C.

Therefore, in one litre (1000 cm^3) of water, the mass that will dissolve is:

$$\left[\frac{1000}{100} \times 25\right] \text{g} = 250 \text{ g}$$

Therefore moles per litre will be:

$$\frac{250 \text{ g}}{\text{molar mass}} = \frac{250}{10.1}$$
$$= 2.4755 \text{ moles/litre}$$

Examples

1.The solubility of $KClO_3$ at 100°C is 60g/100g water. What mass of $KClO_3$ will be deposited at:

(i)75 °C if the solubility is now 39g/100g water.

At 100°C = 60.0g

Less at 75°C = - 39.0g

Mass of crystallized out **21.0g**

(i)35 °C if the solubility is now 28 g/100g water.

At 100°C = 60.0g

Less at 35°C = - 28.0.0g

Mass of crystallized out **32.0g**

2. KNO_3 has a solubility of 42 g/100g water at 20°C.The salt was heated and added 38g more of the solute which dissolved at100°C. Calculate the solubility of KNO_3 at 100°C.

Solubility of KNO_3 at 100°C = solubility at 20°C + mass of KNO_3 added

=> 42g + 38g = **80g KNO_3 /100g H_2O**

3. A salt solution has a mass of 65g containing 5g of solute. The solubility of this salt is 25g per 100g water at 20°C. 60g of the salt are added to the solution at 20°C.Calculate the mass of the solute that remain undissolved.

Mass of solvent at 20°C = mass of solution – mass of solute

=> 65 - 5 = **60g**

Solubility before adding salt = $\frac{\text{mass of solute x 100}}{\text{Volume of solvent}}$

=> $\frac{5 \times 100}{60}$ = **8.3333g/100g** water

Mass of solute to equalize with solubility = 25 – 8.3333g = **16.6667g**

Mass of solute undissolved = 60.0 - 16.6667g = **43.3333** g

Effect of temperature on the solubility of salts.

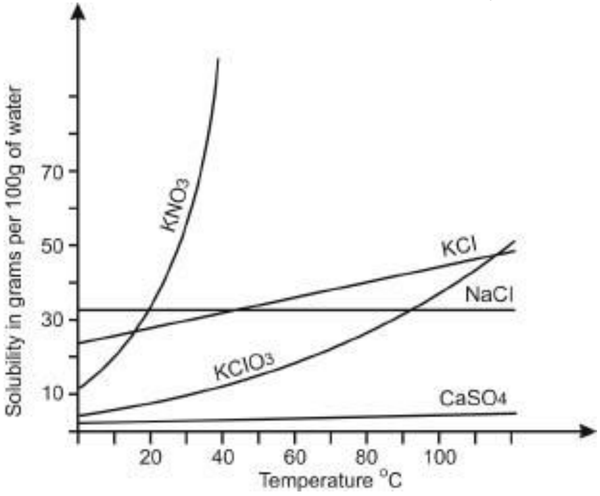
The solubility of a solute in a particular solvent depends on temperature.

For most salts, solubility increases as the temperature rises.

However, the solubility of salts such as **sodium chloride and calcium sulphate remain almost constant with temperature change.**

The solubility of some salts such as **calcium ethanoate ($Ca(CH_3COO)_2$) decreases with increase in temperature.**

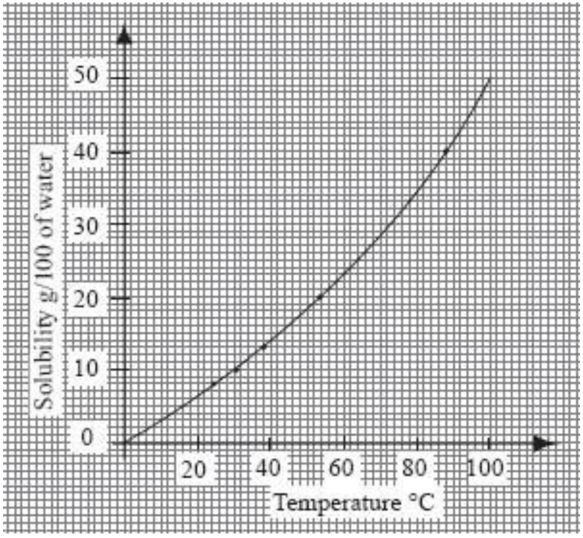
The figure below shows how the solubilities of some familiar salts vary with temperature.



The following results were obtained from an experiment to determine the solubility of potassium chlorate.

Volume of water added	**5**	**10**	**15**	**20**	**25**
Mass of $KClO_3$ in grams per 100 g of water	40.0	20.0	13.3	10.0	8.0
Temperature at which crystals appear	88	52	38	31	25

When solubility is plotted against temperature, the following graph is obtained.



From the graph it is clear that an increase in temperature results in an increase in solubility.

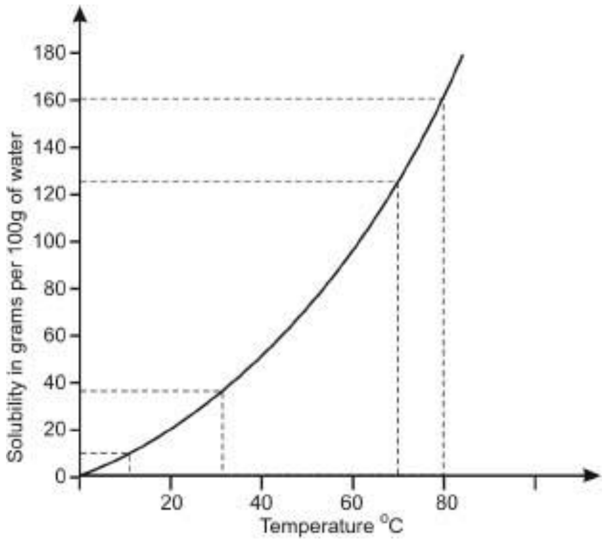
A graph of solubility against temperature for a salt solution is known as a solubility curve.

Importance of solubility curves

Crystallisation

Solubility curves are useful since they give information that can be used to show how crystals are obtained from saturated solutions.

If a saturated solution of potassium chlorate at 70°C is cooled to 30°C, crystallisation occurs. The mass of crystals is determined using the steps below:



(i) Mass of salt at 70°C = 124 g

(ii) Mass of salt at 30°C = 37 g

(iii) Mass of crystals deposited = 124 – 37
= 87 g

Fractional Crystallisation

The knowledge of solubility curves is also useful in the separation of a mixture of salts with different solubilites. The separation of different salts by making use of the differences in solubility in the same solvent is known as **fractional crystallisation**.

For example, if a mixture of salts contains 20 g of potassium nitrate and 18 g of potassium chlorate in 100 g of water at 50°C, one of the salts can be separated using the information in the table below

Salt	**Solubility at various temperatures**		
	50 °C	**20°C**	
	Potassium nitrate	86 g	31 g
	Potassium chlorate	18 g	8 g

If the mixture is cooled from 50°C to 20°C mass of potassium chlorate that crystallises = 18 – 8 = 10 g

But no potassium nitrate will crystallise.

The solubility of one salt has no effect on the solubility of the other.

The curves obtained when the solubility of various salts is against temperature is plotted can be used to determine the amount of salt obtained at a given temperature.

Effect of temperature on the solubility of KNO_3 and $Pb(NO)_3$

Temperature (°C)	0	20	40	60	80	100
Solubility of KNO_3 in 100 g of H_2O	12.5	32.5	62.5	110.0	137.5	
Solubility of $Pb(NO_3)_2$ in 100 g of H_2O	37.5	52.5	69.0	87.5	110.0	131.0

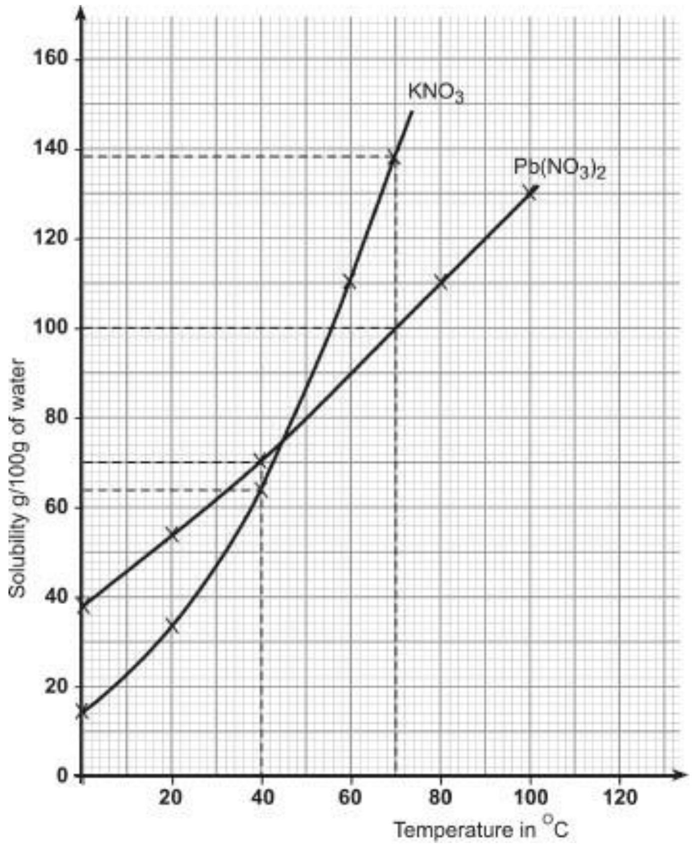
For example, when 100 cm^3 of a saturated solution of potassium nitrate and lead nitrate is cooled from 70°C to 40 °C, the amount of salt that crystallises out is calculated from the graph as follows:

At 70°C the mass of:

(i) Potassium nitrate = 138 g
(ii) Lead nitrate = 100 g

At 40°C, the mass of:

(i) Potassium nitrate = 64 g
(ii) Lead nitrate = 70 g



Therefore, the mass of potassium nitrate crystals deposited is 138 – 64 = 74 g The mass of lead nitrate crystals deposited is (100 – 70) = 30 g

Applications of fractional crystallization: Extraction of salt from sea water.

Lake Magadi contains trona, $Na_2CO_3NaHCO_3.2H_2O$ and common salt NaCl.

During the day when temperatures are **high, sodium carbonate** crystallises out because its **solubility is low at high temperatures** while common salt remains in solution.

At night when the temperatures are **low**, **sodium chloride** crystallises out while sodium **carbonate remains in solution.**

At Ngomeni salt works, sea water is pumped into shallow ponds and evaporation is allowed to take place naturally. As the amount of water reduces, sodium chloride crystallises out because of the increase in its concentration.

Hardness of Water

Hard water *is water that does not readily lather with soap.*

Soft water readily lathers with soap.

Calcium hydrogen carbonate, calcium and magnesium sulphate are the substances that are responsible for hardness of water.

Water containing Ca^{2+} and Mg^{2+} ions does not lather readily with soap. The ions of calcium and magnesium react with soap to form an insoluble substance called **scum.**

Temporary and permanent hardness of water.

Water hardness that is removed by boiling is **temporary hardness**. Temporary hardness of water is caused by the presence of the hydrogen carbonates of calcium and magnesium.

When an aqueous solution of calcium hydrogen carbonate is boiled, it decomposes into calcium carbonate which precipitates out. Similarly, when a solution containing magnesium hydrogen carbonate is boiled, it decomposes to magnesium carbonate.

The resulting solutions form lather readily with soap.

$$Ca(HCO_3)_2(aq) \longrightarrow CaCO_3(s) + H_2O(l) + CO_2(g)$$

$$Mg(HCO_3)_2(aq) \longrightarrow MgCO_3(s) + H_2O(l) + CO_2(g)$$

The above carbonates, magnesium and calcium hydrogen carbonate exist only in solution.

Hardness that cannot be removed by boiling is called **permanent hardness. Permanent hardness** of water is caused by the presence of the **sulphates of magnesium and calcium** in water. **Calcium sulphate and magnesium sulphate do not decompose on heating thus boiling cannot be used to remove permanent hardness of water.**

Disadvantages of Hard Water

(a) Wastes soap.

(b) Stains white clothes. The soap precipitates out Ca^{2+} and Mg^{2+} as scum which form stains on the clothes.

(c) Deposition of fur (calcium carbonate) in kettles, pipes and boilers. This reduces efficiency in boilers, and may make hot water pipes to burst.

Advantages of Hard Water

(a) It provides calcium which is essential for strengthening bones.

(b) Hard water does not dissolve lead from lead pipes because a layer of carbonate is formed as a lining therefore there is no contact between the water flowing in the pipes and lead.

(c) It is used in beer brewing.

Methods of Removing Hardness of Water

Hardness can be removed by various methods, which include:

(a) **Boiling**: This method removes temporary hardness only. It cannot be used to remove permanent hardness since heating has no effect on the sulphates of calcium and magnesium.

During boiling, calcium hydrogen carbonate and magnesium hydrogen carbonate in the water are decomposed to produce the insoluble carbonates.

$$Ca(HCO_3)_2(aq) \longrightarrow CaCO_3(s) + CO_2(g) + H_2O(l)$$

$$Mg(HCO_3)_2(aq) \longrightarrow MgCO_3(s) + CO_2(g) + H_2O(l)$$

(b) **Distillation**: This method removes both types of hardness but it is expensive. During distillation, the water is allowed to evaporate, then condensed and collected as the distillate. The salt is left behind as the residue.

(c) **Addition of sodium carbonate**: The Ca^{2+} and Mg^{2+} ions are precipitated out. This method removes both types of harness.

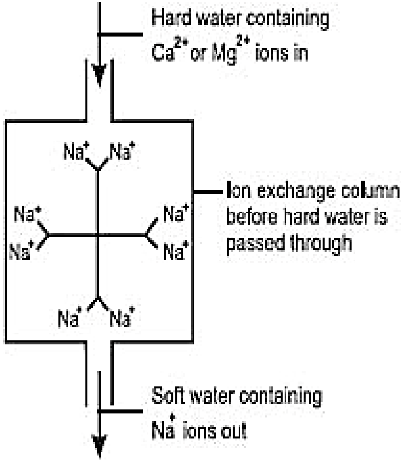
$$Ca^{2+}(aq) + CO^{2-}(aq) \longrightarrow CaCO_3(s)$$

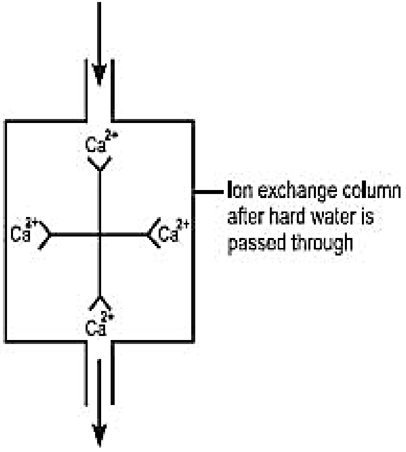
$$Mg^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)$$

(d) **Ion exchange**: In this process hard water is passed through a column filled with some complex sodium compounds which can be represented by the formula Na_2X where X is a complex.

$$Na_2X(s) + Ca^{2+}(aq) \longrightarrow CaX(s) + 2Na^{+}(aq)$$

The Ca^{2+} and Mg^{2+} ions are precipitated as CaX and MgX which remain in the column as shown below.





The column gradually becomes converted to calcium and magnesium complex, and loses the ability to remove calcium and magnesium ions. It can be recovered by **flushing** it with a **concentrated solution of sodium chloride (brine)** to form **calcium chloride and magnesium chloride which are washed away**. This way, the column is recharged or regenerated. The ion exchange method removes **both types of hardness**.

$$CaX(s) + 2NaCl(aq) \longrightarrow CaCl_2(aq) + Na_2X(s)$$

$$MgX(s) + 2NaCl(aq) \longrightarrow MgCl_2(aq) + Na_2X(s)$$

(e) **Addition of calcium hydroxide**: In this process, a known amount of calcium hydroxide is added to water containing calcium hydrogen carbonate, calcium carbonate is precipitated.

$$Ca(HCO_3)_2(s) + Ca(OH)_2(aq) \longrightarrow 2CaCO_3(s) + 2H_2O(l)$$

(f) **Addition of ammonia solution**. Aqueous ammonia is added to water containing calcium hydrogen carbonate which precipitates out as calcium ions.

$$Ca(HCO_3)_2(aq) + 2NH_4OH(aq) \longrightarrow CaCO_3(s) + 2H_2O(l) + (NH_4)_2CO_3(aq)$$

This method removes temporary hardness only.

Review Questions

1. 2006 Q 11 P1

(a) Water from a town in Kenya is suspected to contain chloride ions but not sulphate ions. Describe how the presence of the chloride ions in the water can be shown. (1 mark)

(b) State one advantage of drinking hard water rather than soft water. (1 mark)

2. 2006Q 14 P1

Below is a list of oxides.

MgO, N_2O, K_2O, CaO, and Al_2O_3.

Select:

(a) A neutral oxide (1 mark)
(b) A highly water-soluble basic oxide (1 mark)
(c) An oxide which can react with both sodium hydroxide solution and dilute hydrochloric acid.
(1 mark)

3. 2006 Q 24P1

(a) Complete the table below to show the colour of the given indicator in acidic and basic solutions. (1 mark)

Indicator	Colour in	
	Acid solution	Basic solution
Methyl orange		Yellow
Phenolphthalein	Colourless	

(b) How does the PH value of 0.1M potassium hydroxide solution compare with that of 0.1M aqueous ammonia? Explain. (2 marks)

4. 2007 Q 15 P1

(a) Explain why permanent hardness in water cannot be removed by boiling.

(2marks)

(b) Name two methods that can be used to remove permanent hardness from water.

(1mark)

5. 2007 Q 16 P1

The table below shows the tests that were carried out on solid N and the observations made.

	Test	Observations
I	Solid N was heated	Solid N turned from white to yellow.
II	Dilute hydrochloric acid was added to solid N.	A colourless solution was formed.
III	To the colourless solution obtained in test I, excess sodium hydroxide solution was added.	A white precipitate was formed which dissolved to form a colourless solution.

Write the formula of the anion in;

(a) Solid **N** (1mark)

(b) The colourless solution formed in test **II** (1mark)

6. 2007 Q 6 P2

(a) The elements nitrogen, phosphorous and potassium are essential for plant growth. Potassium in fertilizers may be in the form of potassium nitrate.

(i) Describe how a sample of a fertilizer may be tested to find out if it contained nitrate ions. (2marks)

(ii) Calculate the mass of nitrogen present if a 25kg bag contained pure ammonium phosphate, $(NH_4)_2HPO_4$. (N = 14.0, H=1.0, P = 31.0, O = 16.0) (2 marks)

(b) The table below shows the solubility of ammonium phosphate in water at different temperatures.

Temperature (°C)	Solubility of ammonium phosphate in g/100g water
10	63.0
20	69.0
30	75.0
40	82.0
50	89.0
60	97.0

(i) On the grid provided, draw the solubility curve of ammonium phosphate (Temperature on x - axis) (3 marks)

(ii) Using the graph, determine the solubility of ammonium phosphate at 25 °C. (1 mark)

(iii) 100g of a saturated solution of ammonium phosphate was prepared at 25 °C

I. What is meant by a saturated solution? (1 mark)

II. Calculate the mass of ammonium phosphate which was used to prepare the saturated solution. (2 marks)

7. 2008 Q 15 P1

The table below gives the solubilities of substances J, K and L at different temperatures

Substance	Solubility in grams per 100 g water at			
	0 °C	20°C	40°C	60°C
J	0.334	0.16	0.097	0.0058
K	27.60	34.0	40.0	45.5
L	35.70	36.0	40.0	37.3

Select the substance which, when dissolved in water, heat is given out. Give a reason. (2 marks)

8. 2008 Q 17 P1, 2015 Q17 P1

A compound whose general formula is $M(OH)_3$ reacts as shown by the equation below.

$M(OH)_3(s) + OH^-(aq) \longrightarrow M(OH)_4^-(aq)$

$M(OH)_3(s) + 3H^+(aq) \longrightarrow M^{3+}(aq) + 3H_2O(l)$

(a) What name is given to compounds which behave like $M(OH)_3$ in the two reactions (1 mark)

(b) Name two elements whose hydroxides behave like that of M. (2 marks)

9. 2009 Q 2 P1

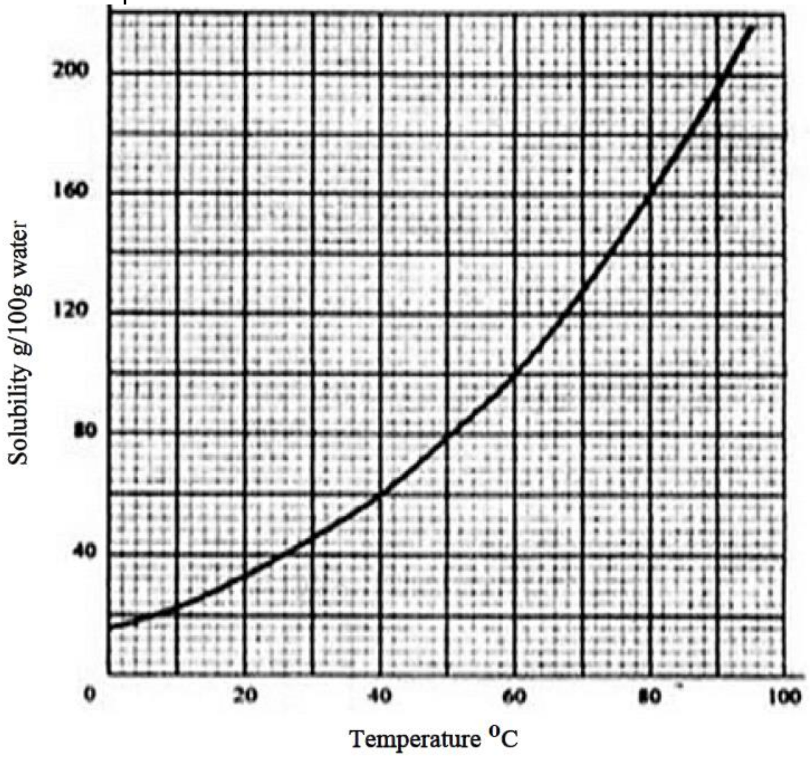
Hardness of water may be removed by either boiling or addition of chemicals.

(a) Write an equation to show how boiling removes hardness of water. (1 mark)

(b) Name two chemicals that are used to remove hardness of water. (2 marks)

10. 2009 Q 16 P1

The solubility curve of potassium nitrate is shown below.

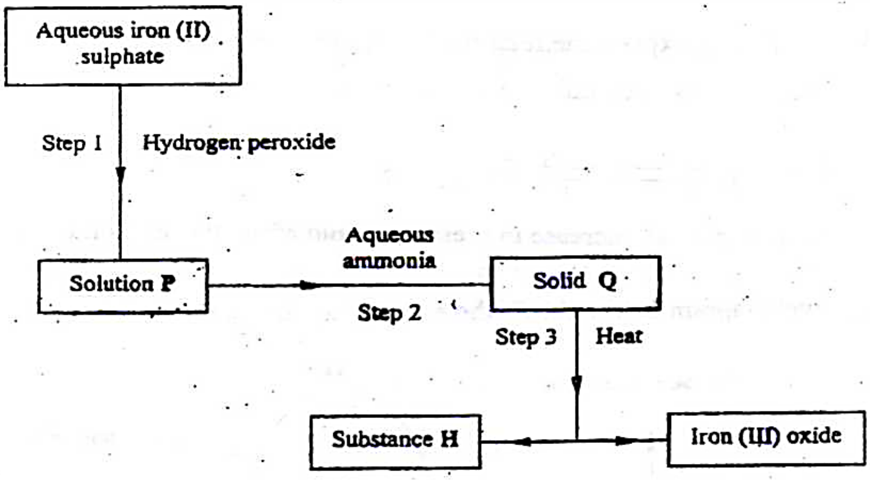


(a) Determine the solubility of potassium nitrate at 50°C. (1 mark)

(b) Determine the molar concentration of saturated potassium nitrate at 50°C. (K = 39.0, O = 16.0, N = 14.0 and density of water = 1 g/cm^3). (2 marks)

11. 2009 Q 23 P1

Use the flow chart below to answer the questions that follow.



(a) What observation would be made in step **1**? (1 mark)

(b) Name another substance that could be used in step **2** (1 mark)

(c) Give the name of substance H. (1 mark)

12. 2009 Q 25 P1

For each of the following reactions, state the observation and write the formula of the compound responsible for the observation.

(a) Bromide water is added to aqueous potassium iodine (1½ mark)

(b) Excess aqueous ammonia is added to copper (II) hydroxide (precipitate)

(1½ mark)

13. 2010 Q 16 P1

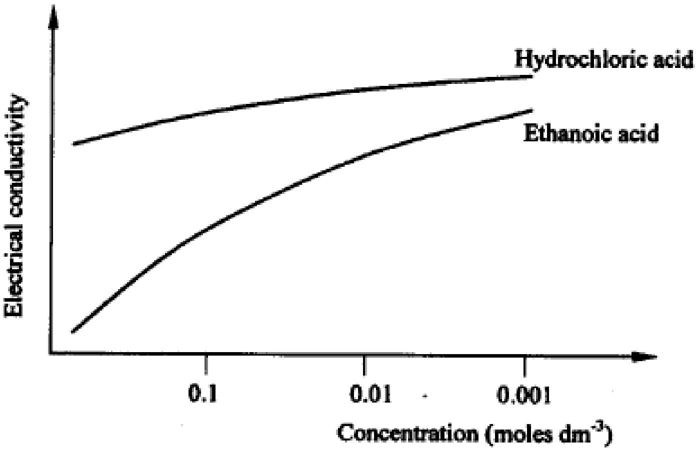
A sample of fertilizer is suspected to be calcium ammonium nitrate. Describe chemical tests for each of the following ions in the sample:

(a) Calcium ions; (2 marks)

(b) Ammonium ions. (1 mark)

14. 2010 Q 23 P1

The curves below shows how the electronic conductivity of hydrochloric and ethanoic acids vary with concentration.



Explain why the electrical conductivity of 0.01M hydrochloric acids is higher than that of 0.01M ethanoic acid. (2 marks)

15. 2010 Q 25 P1

A sample of river water was divided into three portions. The table below shows the test carried out on the portions and the observations made.

Test	**Observation**	**Inference**
To the first portion, 1 cm^3 of soap solution was added	No lather formed.	
The second portion was boiled, cooled and 1 cm^3 of soap solution was added.	No lather formed.	
To the third portion, 3 cm^3 of aqueous sodium carbonate was added, the mixture filtered and 1 cm^3 of soap solution added to filtrate	Lather formed immediately.	

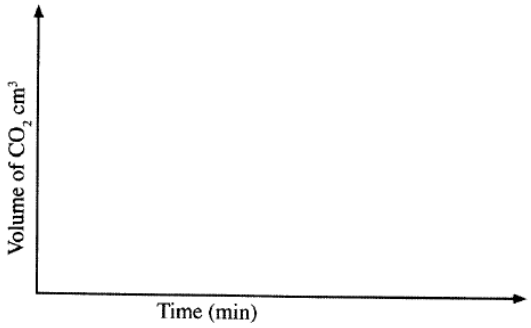
Complete the table by filling in the inferences (3 marks)

16. 2011 Q 8 P1

(a) What is meant by a strong acid? (1 mark)

(b) In an experiment, 40 cm^3 of 0.5 M hydrochloric acid was reacted with excess sodium carbonate and the volume of carbon (IV) oxide produced recorded with time. In another experiment, the same volume and concentration of ethanoic acid was also reacted with excess sodium carbonate and the volume of carbon (IV) oxide produced recorded with time.

On the grid below, sketch and label the curves if the volumes of carbon (IV) oxide were plotted against time. (2 marks)



17. 2011 Q 15 P1

Soap dissolves in water according to the equation below

$NaSt(aq) \longrightarrow Na^+(aq) + St^-$ where St^- is the stearate ion.

(a) Write the formula of the scum formed when soap is used in hard water. (1 mark)

(b) Write the ionic equation for the reaction that occurs when sodium carbonate is used to remove hardness in water. (3 marks)

18. 2011 Q 30 P1

A sample of river water is suspected to contain zinc ions. Describe how the presence of zinc ions and sulphate ions can be established. (3 marks)

19. 2012 Q6 P1

Study the information in the table below and answer the questions that follow:

Salt	Solubility (g/100g water)	
	At 40°C	At 60°C
$CuSO_4$ $Pb(NO_3)_2$	28 79	38 98

A mixture containing 35g of $CuSO_4$ and 78g of $Pb(NO_3)_2$ in 100g of water at 60°C was cooled to 40°C.

(a) Which salt crystallised out? Give a reason. (2 marks)

(b) Calculate the mass of the salt that crystallised out. (1 mark)

20. 2012 Q10 P1

(a) Name two cations that are present in hard water. (1 mark)

(b) Explain how the ion exchange resin softens hard water. (2 marks)

21. 2013 Q9 P1

Aqueous hydrogen chloride reacts with potassium manganate (VII) to produce chlorine gas, while a solution of hydrogen chloride in methylbenzene has no effect on potassium manganate (VII). Explain this observation.
(2 marks)

22. 2013 Q10 P1

The table below gives the solubilities of substances T and U at 10°C and 40°C.

Substance	**Solubility g/100g water**	
	10 °C	40 °C
T	40	65
U	15	17

When an aqueous mixture containing 55g of T and 12g of U at 80 °C was cooled to 10 °C, crystals formed.

(a) Identify the crystals formed . (1 mark)

(b) Determined the mass of the crystals formed. (1 mark)

(c) Name the method used to obtain the crystals. (1 mark)

23. 2014 Q12 P1

The table below shows the pH values of solutions A, B, C and D

Solution	A	B	C	D
pH	2	7	11	14

Select solutions in which a sample of lead (II) hydroxide is likely to dissolve. Give reasons for each solution selected. (3 marks)

24. 2014 Q7 P2

(a) Dissolving of potassium nitrate in water is an endothermic process. Explain the effect of increase in temperature on the solubility of potassium nitrate.
(2 marks)

(b) The table below shows the solubilities of potassium sulphate and potassium chlorate (V) at different temperatures.

Temperature (°C)	0	20	40	60	80	100
Solubility of K_2SO_4 g/100 g water	8.0	10.0	14.0	17.5	20.0	22.0
Solubility of $KClO_3$ g/100g water	3.0	5.0	15.5	24.0	38.0	53.0

(i) Draw the solubility curves for both salts on the same axis. (Temperature on the X-axis (3 marks)

(ii) A solution of potassium sulphate contains 20g of the salt dissolved in 100 g of water at 100 °C. This solution is allowed to cool to 25 °C

I) At what temperature will crystals first appear?

II) What mass of crystals will be present at 25 °C? (1mark)

(iii) Which of the two salts is more soluble at 30 °C? (1mark)

(iv) Determine the concentration of potassium sulphate in moles per litre when the solubility of the two salts are the same (K= 39.0, O=16.0; S=32.0) (3 marks)

(v) 100 g of water at 100 °C contains 19g of potassium sulphate and 19 g of potassium chlorate (V). Describe how a solid sample of potassium sulphate at 60 °C can be obtained. (2 marks)

25. 2015 Q6 P1

Describe how sodium carbonate is used to remove water hardness (2 marks)

26. 2015 Q6c P1

Two different samples of water (**I** and **II**) were tested with soap solution. Sample **II** was further subjected to two other processes before adding soap. 20 cm^3 of each sample of water was shaken with soap solution in a boiling tube until a permanent lather was obtained. The results are shown in the table below

Water sample	**Volume of soap solution needed (cm^3)**	
	before boiling	**after boiling**
I	10	5
II	6	6
II after filtering	6	6
II after distilling	2	2

(i) Identify the water sample that had temporary hardness. Explain your answer. (2 marks)

(ii) Explain why the results for sample **II** are different after distilling but remain unchanged after filtering. (2 marks)

(iii) State two disadvantages of using both water samples for domestic purposes. (2 marks)

27. 2017 P1 Q9.

Copper(II) ions react with excess aqueous ammonia to form a complex ion.

(a) (i) Write an equation for the reaction that forms the complex ion. (1 mark)

(ii) Name the complex ion. (1 mark)

(b) Explain why CH_4 is not acidic while HCl is acidic yet both compounds contain hydrogen. (1 mark)

28. 2017 P1 Q14.

Using the elements chlorine, calcium and phosphorus:

(a) Select elements that will form an oxide whose aqueous solution has a pH less than 7. (1 mark)

(b) Write an equation for the reaction between calcium oxide and dilute hydrochloric acid. (1 mark)

(c) Give one use of calcium oxide. (1 mark)

29. 2018 P1 Q1.

(a) Define a soluble base. (1 mark)

(b) Aqueous solutions of 2M ethanoic acid and 2M nitric(V) acid were tested for electrical conductivity. Which solution is a better conductor of electricity? Explain. (2 marks)

30. 2018 P2 Q4

(a) An experiment was carried out to prepare crystals of magnesium sulphate. Excess magnesium powder was added to 100 cm3 of dilute sulphuric(VI) acid in a beaker and warmed until no further reaction took place.
The mixture was filtered and the filtrate evaporated to saturation, then left to cool for crystals to form.

(i) Write an equation for the reaction. (1 mark)
(ii) Explain why excess magnesium powder was used. (1 mark)
(iii) State how completion of the reaction was determined. (1 mark)
(iv) What is meant by a saturated solution? (1 mark)
(v) Explain why the filtrate was not evaporated to dryness. (2 mark)

31. 2019 P1 Q6.

A farmer intended to plant cabbages in his farm. He first tested the pH of the soil and found it to be 3.0.If cabbages do well in alkaline soils, explain the advice that would be given to the farmer in order to realise a high yield. (2 marks)

32. 2019 P1 Q16.

(a) Complete the following table. (2 mark)

Solution	pH	Nature of solution
H	1.0	
I		Neutral
J		Weak acid
K	13.0	

(b) Explain why a solution of ammonia in methylbenzene has no effects on red litmus paper while in aqueous ammonia red litmus paper turns blue. (1 mark)

33. 2019 P1 Q25.

Chemical tests were carried out on separate samples of water drawn from the same source. The observations made were recorded as shown in Table 4.

Test	Observation
(i) Addition of aqueous calcium chloride	No white precipitate
(ii) Addition of dilute sulphuric (VI) acid	No effervescence, colourless solution
(iii) Addition of a few drops of acidified barium nitrate	No white precipitate
(iv) Addition of aqueous ammonia	White precipitate dissolves

State the inferences made in reactions:

(i).......................... (1 mark)

(ii)............................ (1 mark)

(iii)............................. (1 mark)

34. 2019 P2 Q4.

(a) Explain the following observations:

(i) The colour of aqueous copper (II) sulphate fades when a piece of magnesium metal is dropped into the solution. (2 marks)

(ii) A piece of iron bar is coated with a brown substance when left in the open on a rainy day. (2 marks)

(b) A sample of water is suspected to contain aluminium ions (Al^{3+}). Describe a laboratory experiment that can be carried out to show that Al^{3+} ions are present in the water sample. (3 marks)

(c) In an experiment to determine the number of moles of water of crystallisation of a hydrated compound $Na_2SO_4.XH_2O$, 5g of the compound were heated strongly to a constant mass.

(i) Explain how a constant mass was obtained. (2 marks)

(ii) During the experiment, the mass of the residue was found to be 2.205 determine the number of moles of water of crystallisation in the compound.
(Na = 23.0; O = 16.0; S = 32.0; H = 1.0) (3 marks)

TOPIC TWO

ENERGY CHANGES IN PHYSICAL AND CHEMICAL PROCESSES.

Table of Contents

Organizer

Objectives

By the end of this topic, the learner should be able to:

(a) Define endothermic and exothermic reactions using the ΔH notation.

(b) Draw energy level and energy cycle diagrams.

(c) Explain fusion and vaporisation as evidence of inter-particle forces.

(d) Explain that energy changes in chemical reactions are due to bond breaking and bond formation.

(e) Define and explain various types of heat changes.

(f) Carry out some experiments to determine enthalpy changes for some reactions.

(g) Write correct simple thermochemical equations.

(h) State Hess's Law and carry out related calculations.

(i) State and explain the factors that influence the choice of a fuel.

(j) Explain the environmental effects of fuels.

TOPIC TWO

ENERGY CHANGES IN PHYSICAL AND CHEMICAL PROCESSES.

Energy is the ability to do work. There are many forms of energy such as chemical, electrical, heat, kinetic and potential energy.

Energy can neither be created nor destroyed but can only be converted from one form to another. For example, heat energy to kinetic energy.

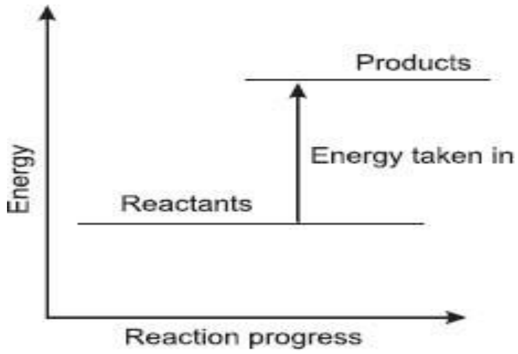
Many physical and chemical changes are accompanied by energy changes.

Endothermic and Exothermic Reactions

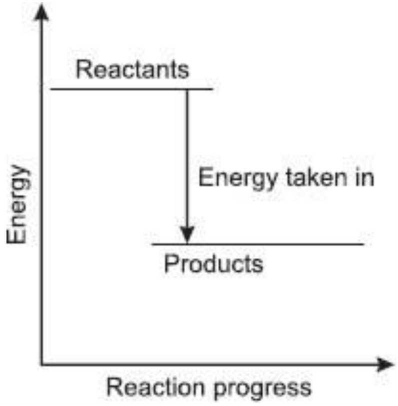
A process that is accompanied by **absorption of heat** is called an endothermic reaction, for example dissolving of ammonium nitrate in water.

The reactants and products in a reaction form a system.

In an endothermic reaction, the system **absorbs heat energy from the surrounding. Therefore, the products have a higher energy compared to the reactants** as shown below:

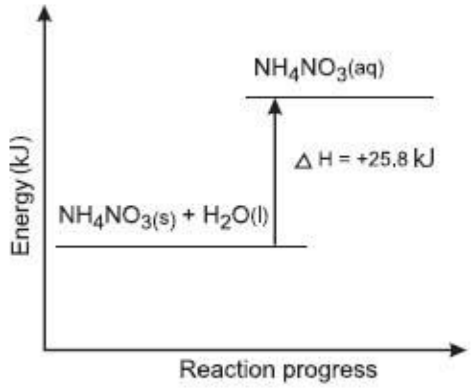


A process that is accompanied by production of heat is called an exothermic reaction, for example dissolving of sodium hydroxide in water. In an exothermic reaction, heat is **lost from the system.** The products therefore have a **lower energy content compared to the reactants** as shown below.

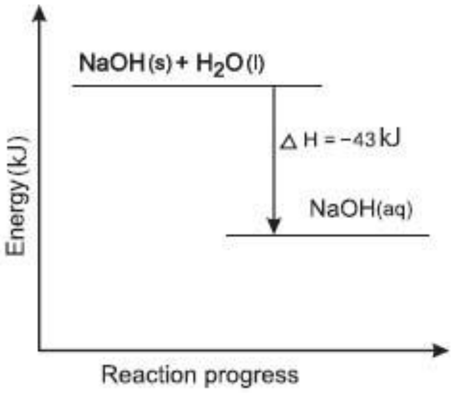


Illustrations which show the relative energies of the products and reactants in a reaction are called energy level diagrams.

The energy level diagrams for ammonium nitrate and sodium hydroxide dissolving in water respectively are:



(a) Ammonium nitrate



(b) Sodium nitrate

Enthalpy Notation

The **heat content** in a chemical is called its **enthalpy**. It is assigned the symbol **H**.

Enthalpy change is a **change in heat content** and is denoted by the symbol **ΔH**.

The enthalpy change during a reaction is given by the expression:

Enthalpy change (ΔH) = Enthalpy of products – Enthalpy of reactants.

In an **endothermic reaction**, the enthalpy of products is **higher** than that of the reactants. The **enthalpy change** is therefore **positive**.

In an **exothermic reaction**, the products have **lower** enthalpy compared to the reactants. The **enthalpy change is** therefore **negative.**

The energy change associated with a chemical change is usually shown **at the end** of the balanced chemical equation. For example,

$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$ ΔH= –728 kJ mol^{-1}

$$NaNO_3(s) + (aq) \xrightarrow{H_2O} NaNO_3(aq) \quad \Delta H = +21 \text{ kJ mol}^{-1}$$

Such a chemical equation is called a **thermochemical equation**.

Bond Breaking and Bond Formation in Physical and Chemical Processes

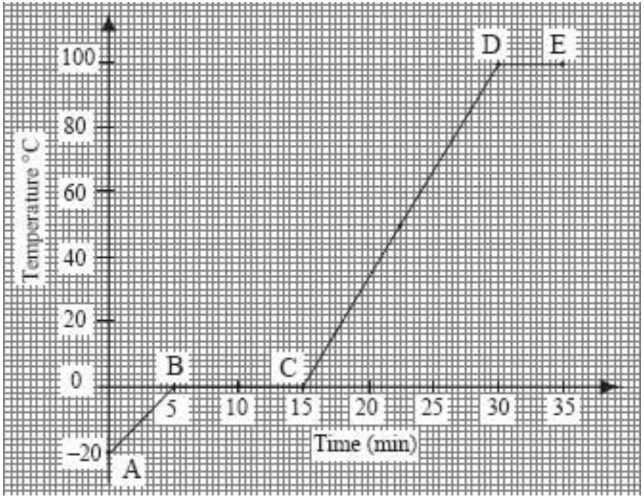
During a physical change such as melting and evaporation, energy changes are involved.

For pure substances, temperature remains constant at the melting and boiling points respectively. The heat energy supplied during melting and boiling is only used to overcome the forces holding the solid lattice together and the forces of attraction between the liquid molecules respectively. It does not therefore result in temperature rise.

The amount of heat energy required to convert a given amount of a solid substance to a liquid at its melting point is called latent heat of fusion.

The amount of heat energy required to convert one mole of a solid substance at its melting point to a liquid and vise versa is called molar heat of fusion.

The heat absorbed by a substance when changing from the liquid state to the gaseous state at constant temperature is known as latent heat of vaporisation. The amount of heat energy required to convert one mole of a liquid substance at its boiling point to its gaseous state is called molar heat of vaporisation



Heating curve of pure ice

If the forces holding the particles together in the solid structure are strong, then the molar heat of fusion is high. This is reflected in the high melting point of the substance. Similarly, when the forces holding the particles in the liquid are strong, the molar heat of vaporisation is high. This is reflected in the high boiling point of the substance.

The atoms in elements or compounds are held together by chemical bonds. For reactions to occur between substances, the bonds must be broken and new bonds formed as the constituent atoms recombine to form new substances.

Bond breaking and bond formation involve energy changes in the substances involved. For example, methane reacts with chlorine as shown below:

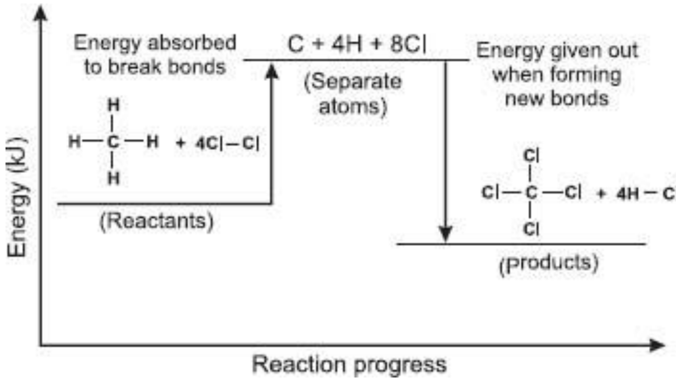
$$CH_4(g) + 4Cl_2(g) \longrightarrow CCl_{4+}\ 4HCl(g).$$

For a reaction to occur, the covalent bonds in both methane and chlorine **must be broken** to obtain separate atoms. Energy is **required** to break these bonds. Thus **bond breaking is an endothermic process.**

Once the bonds are broken, the atoms **rearrange themselves** and **new bonds are formed** in the products. In the process of bond formation, **energy is given out**. Therefore, **bond formation is an exothermic process.**

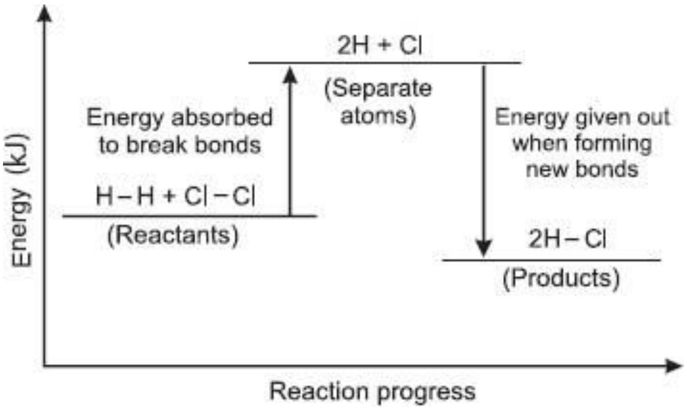
The energy **spent in breaking the bonds** in the methane and chlorine molecules is **less** than the energy **given out when the products are formed.** Overall, the reaction therefore **gives out energy.**

These changes can be illustrated using the energy level diagram shown below for the reaction between methane and chlorine.



Energy level diagram for the reaction between methane and chlorine

The reaction between hydrogen gas and chlorine gas to form hydrogen chloride can be similarly illustrated as shown below:



Energy level diagram for the reaction between hydrogen and chlorine

In reactions where **more energy is spent in breaking the bonds** in the reactants **than** is given out when **new bonds are formed** in the products, the overall process is **endothermic**.

If the bond strengths in reactants and products are known, then the enthalpy change for the reaction can be determined as follows:

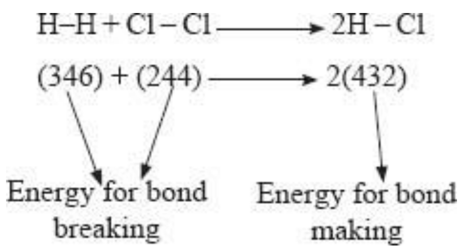
Example

Use the following bond energies to determine whether the reaction below is exothermic or endothermic:

$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$.

Bond	Energy in kJ mol^{-1}
H–H	+436
Cl–Cl	+244
H–Cl	+432

Working



Heat of reaction = Bond breaking energy + Bond formation energy

= 680 kJ + – 864 kJ

= –184 kJ

Since the overall energy is negative (–) the formation of hydrogen chloride (HCl) gas from hydrogen gas and chlorine gas is an exothermic reaction.

Determination of Enthalpy Changes

Enthalpy of solution.

The enthalpy change that occurs when a substance is dissolved in a solvent to give an infinitely dilute solution is called the **enthalpy of solution**.

The molar heat of solution is the enthalpy change that occurs one mole of a substance is dissolved in a solvent to give an infinitely dilute solution.

Ammonium nitrate

To determine the molar enthalpy of solution of ammonium nitrate:

Measure 100 ml distilled water into a clean 250 ml plastic beaker. Note the temperature of the water. Weigh accurately 2.0 g of ammonium nitrate. Add all the ammonium nitrate at once into the water in the beaker. Stir gently using the thermometer to dissolve the ammonium nitrate. Note and record the steady temperature of the resulting solution.

Discussion Questions

1. **State whether the changes in the experiment are exothermic or endothermic.**

Ammonium nitrate dissolves in water with the **absorption of heat (a drop in temperature).** This change is **endothermic**.

2. **Calculate the enthalpy change in each experiment.**

When 2 g of ammonium nitrate is dissolved in 100 ml distilled water, a temperature change of about 1.5°C (1.5 K) is noted. The enthalpy change is calculated using the formula:

Enthalpy change = Mass of solution × Specific heat capacity × Temperature change

ΔH = MCΔT where M = Mass

C = Specific capacity

ΔT = Change in temperature in Kelvin

In this case, it is **assumed** that the **mass of the solution is equal to that of a similar amount for water, i.e., 100 g (0.1 kg). The specific heat capacity of water is used in the calculation.** Therefore:

Heat absorbed = 0.1kg × 4.2 kJ Kg^{-1} K^{-1} × 1.5 K

= 0.63 kJ

3. **Calculate the number of moles of solute dissolved in each experiment.**

The number of moles of ammonium nitrate is calculated as follows:

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{2.0\text{ g}}{80\text{ g mol}^{-1}} = 0.025\text{ mol}$$

4. **Calculate the enthalpy change that would occur if one mole of each solute is dissolved.**

when 0.025 mol/ of ammonium nitrate dissolves, 0.63 kJ of heat is absorbed. Therefore the enthalpy change when 1 mol of ammonium nitrate dissolves is:

$$= \frac{1\text{ mol} \times 0.63\text{ kJ}}{0.025\text{ mol}} = 25.2\text{ kJ}$$

Thus when 1 mol of ammonium nitrate dissolves 25.2 kJ of heat is absorbed

5. **Write the equation for each reaction. Include the energy change.**

The equation for the reaction including the enthalpy change is:

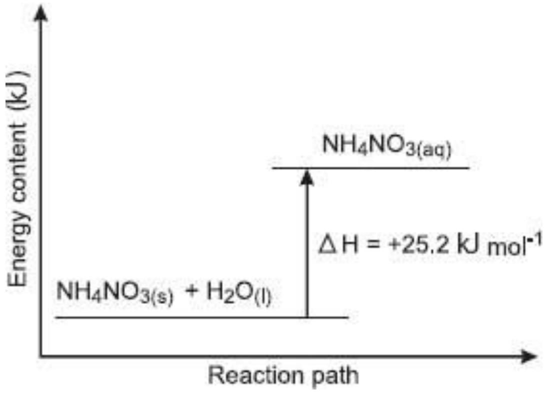
$NH_4NO_3(s) \longrightarrow NH_4NO_3(aq)$ ΔH = + 25.2 kJ mol^{-1}

This value: + 25.22 kJ mol^{-1} is called the <u>molar heat of solution</u> of ammonium nitrate.

NB: A mole of solute should be dissolved in a sufficient amount of solvent to ensure complete dissolution.

A chemical equation which shows the enthalpy change during reaction is called a <u>thermochemical equation</u>.

A thermochemical equation is illustrated on an energy level diagram as shown below:



Energy level diagram for dissolving ammonium nitrate

Sodium hydroxide

Sodium hydroxide dissolves in water with the **evolution of heat**. A temperature **rise** of about **5°C** is recorded when 2 g of sodium hydroxide dissolves in 100 ml distilled water.

Heat evolved = Mass × Specific heat capacity × Temperature change

*= 0.1 Kg × 4.2 kJ Kg**1** K**1** × 5 K*

= 2.1 kJ

Moles of sodium hydroxide used is given by:

$$\text{No. of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

$$= \frac{2.0\text{ g}}{40\text{ g mol}^{-1}}$$

$$= 0.05\text{ mol}$$

Therefore, 0.05 mol liberates 2.1 kJ of heat.

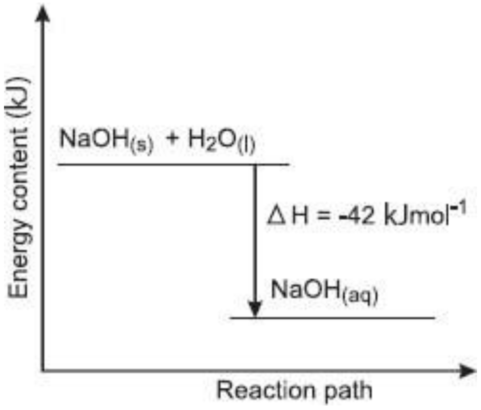
$$1\text{ mol liberates: } \frac{1.0 \times 2.1}{0.05}\text{ kJ of heat} = 42\text{ kJ}$$

The thermochemical equation for the reaction is:

$NaOH(s) \longrightarrow NaOH(aq)$ ΔH= –42 kJ mol-**1**

This (i.e., – 42 kJ mol-**1**) is the **molar heat of solution of sodium hydroxide.**

The thermochemical equation can be illustrated on an energy level diagram.



Energy level diagram for dissolving sodium hydroxide.

Concentrated sulphuric(VI) acid?

Wrap a clean 250 ml plastic beaker with tissue paper. Secure the tissue paper with a rubber band. Measure exactly 98 cm**3** of distilled water into the beaker. Note the steady temperature of the water. Measure 2 cm**3** of concentrated sulphuric(VI) acid. Hold the beaker containing distilled water in a tilted position and carefully pour the 2 cm**3** of concentrated sulphuric(VI) acid into the beaker.

Discussion Questions

1. **State whether the change is exothermic or endothermic.**

Concentrated sulphuric(VI) acid dissolves in water with the evolution of heat. The reaction is exothermic.

2. **Determine the temperature change for the reaction.**

When 2cm^3 of concentrated sulphuric(VI) acid is dissolved in 98 cm^3 of water, a temperature rise of 4.5°C is recorded.

3. **Calculate the enthalpy change for the reaction.**

The total volume of the solution = (2 + 98) cm^3 = 100 cm^3

The mass of solution is given by = Volume × Density

= 100 cm^3× 1.0 g cm^{-3}

= 100g (0.1g)

Heat evolved = Mass × Specific heat capacity × Temperature change

= 0.1 Kg × 4.2 kJ Kg^{-1} K^{-1} × 4.5 K = 1.89 kJ

4. **Determine the mass of the concentrated acid used hence the moles of sulphuric(VI) acid. (Concentrated sulphuric(VI) acid is 98% pure and has a specific density of 1.84 (cm**-3**).**

The mass of the concentrated acid is calculated from the formula:

Mass = Specific density × Volume

= 1.84g cm^{-3} × 2 cm^3 = 3.68 g

Thus the 2 cm^3 of concentrated sulphuric(VI) acid has a mass of 3.68 g. But since the acid is 98% pure, the actual mass of sulphuric(VI) acid is calculated by:

Mass of acid = Mass of conc. acid × percentage purity

$$= 3.68\text{ g} \times \frac{98}{100} = 3.6064\text{ g}$$

The moles of the acid is given by:

$$= \frac{3.6064 \text{ g}}{98 \text{ g mol}^{-1}}$$

$$= 0.0368 \text{ mol}$$

5. **Calculate the enthalpy change when one mole of concentrated sulphuric(VI) acid dissolves.**

when 0.0368 mol of sulphuric(VI) acid dissolves, 1.89 kJ of heat is evolved.

When 1 mol of the sulphuric(VI) acid dissolves, $\frac{1.0 \times 1.89 \text{ kJ}}{0.0368}$ of heat is evolved.

= 51.358 kJ of heat.

Thus the molar heat of solution of sulphuric(VI) acid

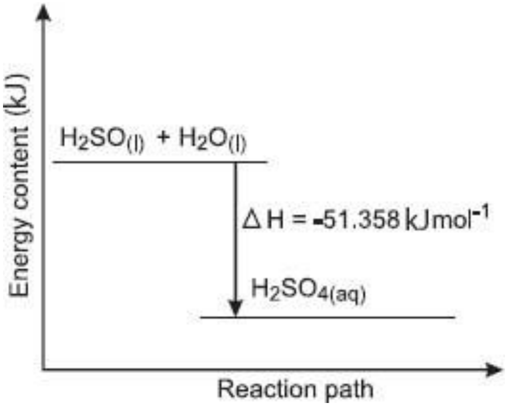
= –51.358 kJ mol^{-1}

The negative sign is added to show that the process is exothermic.

6. **Write down the thermochemical equation for the reaction**.

$$H_2SO_4\,(l) \xrightarrow{H_2O} H_2SO_4(aq) \qquad \Delta H = -51.358 \text{ kJ mol}^{-1}$$

The thermochemical equation is illustrated on an energy level diagram as shown.



Energy level diagram for dissolving concentrated sulphuric(VI) acid.

NB:The **experimental values** of molar heat of solution obtained in experiments are **usually lower than those quoted in data books**, since the **heat gained by the apparatus** and that **lost to the surrounding** during an experiment is **not accounted for in the calculations.**

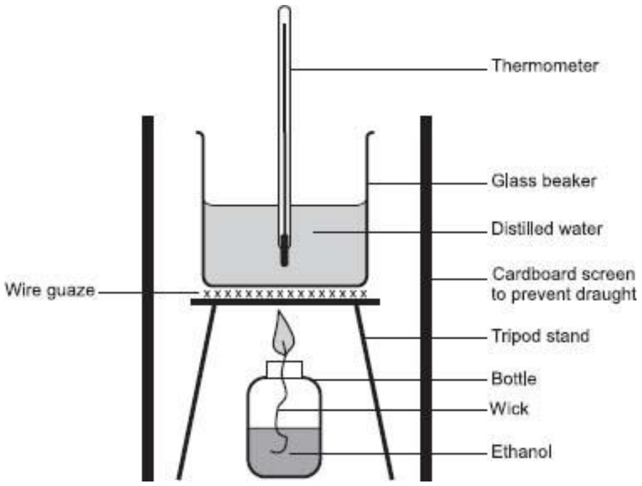
Enthalpy of combustion.

Enthalpy of combustion is the enthalpy change that occurs when a substance burns completely in oxygen.

Molar heat of combustion is the enthalpy change that occurs when one mole of a substance is completely burned in oxygen.

To determine the molar heat of combustion of ethanol,

Pour exactly 100 cm**3** of distilled water into a 250 ml glass beaker. Record the temperature of the water. Half fill a small bottle with ethanol and close the bottle with a lid fitted with a clean wick.



Weigh the bottle and its contents and arrange the apparatus as shown.

Light the wick to start heating the water. Stir the water carefully with the thermometer and extinguish the flame when the temperature of the water has risen by 15°C. Record the temperature of the water. Weigh the bottle and its contents again. Record the results obtained.

The following results were obtained in a similar experiment.

Initial mass of bottle + contents = 28.25 g

Final mass of bottle + contents = 28.02 g

Mass of ethanol burnt = 0.23 g

Final temperature of water = 43.0°C

Initial temperature of water = 28.0°C

Rise in temperature of water = 15°C (15 K)

QUESTIONs

1. **Determine the number of moles of the ethanol burned.**

Formula mass of ethanol, $CH_3CH_2OH = (12 \times 2) + 16 + 6 = 46$ g

$$\text{Moles of ethanol burned} = \frac{0.23 \text{ g}}{46 \text{ g mol}^{-1}}$$

= 0.005 mol

2. **Determine the amount of heat given out.**

The heat evolved by the burning ethanol is absorbed by the water. It is this heat that raises the temperature of water. Therefore, the heat evolved is determined thus:

Heat evolved = Mass of solution × Specific heat capacity × Temperature change

$= 0.1 \text{ kg} \times 4.2 \text{ kJ g}^{-1} \text{ K}^{-1} \times 15 \text{ k} = 6.3 \text{ kJ}$

3. **Using your answers to questions 1 and 2, calculate the molar enthalpy of combustion of ethanol.**

0.005 moles of ethanol liberate 6.3 kJ of heat.

Therefore 1 mol of ethanol will liberate:

$$\frac{6.3 \text{ kJ} \times 1.0 \text{ mol}}{0.005 \text{ mol}}$$

= 1260 Kg of heat

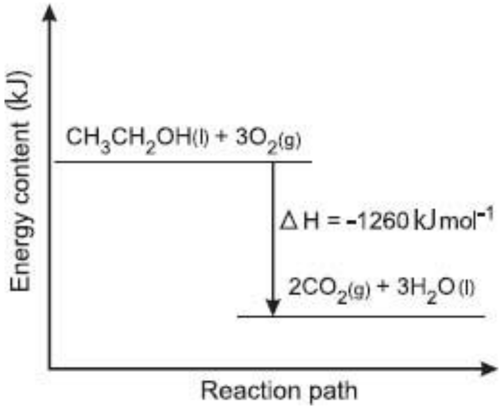
4. **The theoretical enthalpy of combustion of ethanol is –1368 kJ mol-1, why does the value calculated from experimental results differ from this?**

Experimentally determined heats of combustion are usually lower than the theoretical values because the heat lost to the surrounding is not accounted for in the calculations (experimental error).

5. **Write the thermochemical equation for the reaction.**

$CH_3CH_2OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$ $\Delta H = -1260$ kJ mol^{-1}

The thermochemical equation is illustrated on an energy level diagram as shown.



Energy level diagram for the combustion of ethanol

The negative sign is added to show that the process is exothermic.

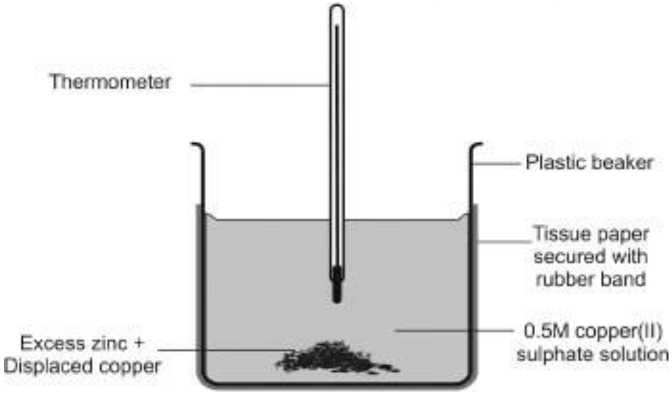
3.Enthalpy of Displacement

The enthalpy of displacement is the enthalpy change that occurs when a substance is displaced from a solution of its ions.

The Molar heat of displacement ***is the enthalpy change that occurs when one mole of a subtance is diplaced from a solution of its ions.***

To determine the enthalpy of displacement of copper,

Wrap a 250 cm^3 plastic beaker with tissue paper. Measure 100 cm^3 of 0.5 M copper (II) sulphate into the beaker. Note the steady temperature of the solution. Carefully, transfer 4.0 g of zinc powder into the plastic beaker and stir carefully with the thermometer. Record the highest temperature attained by the solution.



Questions

1. **State whether the reaction is exothermic or endothermic**.

When zinc is added to copper(II) sulphate solution, an **exothermic** reaction occurs.

2. **What other observations were made? Explain.**

The **blue colour of the solution fades** as a **brown solid is deposited**.

During the reaction, the blue copper(II) ions in the solution are **reduced to copper metal** which is the **brown solid** deposited at the bottom of the beaker.The zinc atoms are **oxidised to zinc ions which are colourless**

3. **Write the equation for the reaction that takes place.**

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

4. **Determine the temperature change for the reaction.**

The expected change in temperature for this experiment is 24.5°C (24.5 K).

5. **Calculate the number of moles of zinc used.**

$$\text{Moles of zinc used} = \frac{4\text{ g}}{65\text{ g mol}^{-1}}$$

$$= 0.062\text{ mol}$$

6. **Calculate the number of moles of copper.**

$$\text{Moles of copper (II) ions} = \frac{0.5\text{ mol} \times 100\text{ cm}^3}{1000\text{ cm}^3}$$

$$= 0.05\text{ mol}$$

7. **Why is it necessary to use excess zinc powder, magnesium or iron filings in this reaction?**

Excess zinc is used in this experiment in order to ensure that all the copper(II) ions are changed to copper metal.

8. **Calculate the molar heat of displacement of copper(II) ions with zinc. Given that:**

(a) Density of solution is 1 g cm-3

(b) The volume of the solution remains unchanged after the reaction.

(c) The specific heat capacity of the solution is 4.2 kJ g -1 **K**-1 **(Cu= 63.5, S=32.0, O =16.0, Zn = 65.0)**

From the equation the reaction ratio between metal and copper(II) ions is 1:1.

Heat evolved = Mass × Specific heat capacity × Temperature change

$= 0.1\text{ kg} \times 4.2\text{ kJ Kg}^1\text{ K}^1 \times 24.5\text{ K}$

= 10.29 kJ

Thus, when 0.05 mol of copper(II) ions are displaced from solution by zinc, 10.29 kJ of heat is evolved.

When 1 mol of Copper (II)ions are displaced, $\frac{10.29 \times 1}{0.05}$ = 205.8 kJ of heat is evolved.

The molar heat of displacement of copper by zinc therefore, is –205.8 kJ mol^{-1}

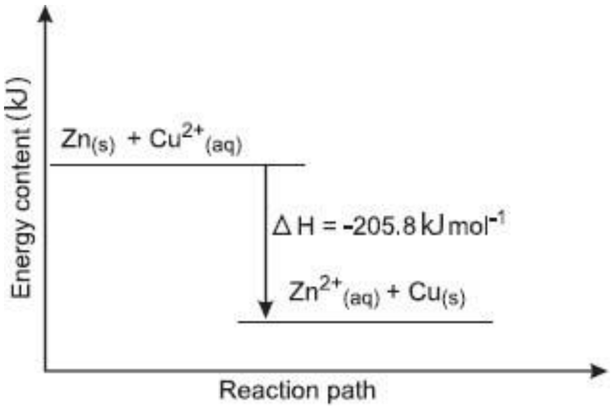
9. **Comment on your results given that the molar heat of displacement of copper (II) ions and zinc is 216 kJ mol**-1

The experimental value obtained in this experiment is low because the heat lost to the surrounding and that absorbed by the apparatus is not accounted for in the calculation.

10. **Write the thermochemical equation for the reaction and represent the information in an energy level diagram.**

Thermochemical equation for the reaction is:

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+} + Cu(s)\ \Delta H = -205.8\text{ kJ mol}^{-1}$$



Energy level diagram for displacement of copper(II) ions from solution

The value obtained when magnesium is used is higher than that of zinc and iron, while that of zinc is higher than that of iron. **The further apart the two metals are in the reactivity series, the higher the molar heat of displacement.**

4.Enthalpy of neutralisation

The enthalpy of neutralization is the enthalpy change that occurs when an acid and a base react to form water and a salt.

The molar heat of neutralisation is *the enthalpy change that occurs when an acid and a base react to produce one mole of water.*

Strong acids and strong alkalis dissociate completely in water into their ions. During neutralisation, hydrogen ions react with hydroxide ions to form water molecules. Heat energy is liberated in the process.

To determine the enthalpy of neutralization of sodium hydroxide with hydrochloric acid,

Wrap a clean 250 ml plastic beaker with tissue paper. Secure the tissue paper with a rubber band. Transfer exactly 50 cm**3** of 2 M hydrochloric acid solution into the beaker. Note the steady temperature T**1** of the hydrochloric acid solution. Using a clean dry measuring cylinder, measure exactly 50 cm**3** of 2 M sodium hydroxide solution and note its steady temperature T**2**. Add the 50 cm**3** of 2 M hydrochloric acid. Carefully, stir the contents of the beaker with the thermometer while adding the sodium hydroxide solution. Note the highest temperature T**4** attained by the resulting solution. Record the results.

A sample calculation on enthalpy is shown below.

Temperature of hydrochloric acid, t**1** = 22.5°C

Temperature of sodium hydroxide solution t**2** = 22.5°C

Average temperature of the acid and alkali,

$$t_3 = \frac{t_1 + t_2}{2} = \frac{22.5 + 22.5}{2} = 22.5\ °C$$

The highest temperature of the mixture t**4**

= 35.5 °C

Temperature change ΔT = t**4** – t**3**

= 35.5 – 22.5

= 13.0 °C

The specific heat capacity of the solution is 4.2 kJ Kg**-1** K**-1**. If in the experiment, 50 cm**3** of 2 M hydrochloric acid are neutralised by 50 cm**3** of 2 M sodium hydroxide, then the final volume of the resulting solution

= 50 + 50 = 100 cm**3**

If the density of the resulting solution is taken to be 1 g/cm**3** then:

The mass of the solution = 100 × 1 = 100 g

Heat evolved = MC ΔT

$$= \frac{100 \times 4200\ \text{J} \times 13.0}{1000}$$

= 5460 J = 5.46 kJ

Mol of hydrochloric acid used = $\frac{2 \times 50}{1000}$ = 0.1 mol. Therefore,

mol of H^+ ions = 0.1 mol since one mole of HCl produces one mole of H^+ ions.

Mol of sodium hydroxide used = $\frac{2 \times 50}{1000}$ = 0.1mol

Mol of OH^- ions = 0.1 mol.

The equation of the reaction is

$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

Thus 1 mol of H^+ reacts with 1 mol of OH^- to form 1 mol of water.

In the reaction between the acid and the base, 0.1 mol of water are formed.

When 0.1 mol of water are formed 5.46 kJ of heat energy is evolved. Therefore, when 1.0 mol of water is formed

$$= \frac{5.46 \times 1}{0.1}$$

= 54.6 kJ of heat energy is evolved.

This is the molar heat of neutralisation of hydrochloric acid by sodium hydroxide.

The enthalpy change can be incorporated in the neutralisation equation as follows:

$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)\ \Delta H_{neut} = -54.6\ \text{kJ mol}^{-1}$

Strong acids and strong bases react to liberate about 57.2 kJ when they form one mol of water during neutralisation.

When <u>**one of the reactants or both are weak,**</u> the enthalpy of neutralisation is **less than in** the case when **strong acids are used with strong bases,** for example, when one **mole of hydrochloric acid (strong acid) is neutralised by one mole of ammonium hydroxide (weak base), about 51.4 kJ of heat is evolved for every one mole of water formed.**

$HCl(aq) + NH_4OH(aq) \longrightarrow NH_4Cl(aq) + H_2O(l)\ \Delta H = -51.4\ \text{kJ mol}^{-1}$

Similarly if **ethanoic acid (weak acid)** is neutralised by **sodium hydroxide (strong base)** only **55.2 kJ** of heat is produced for every mole of water formed.

$C_2H_4CO_2(aq) + NaOH(aq) \longrightarrow C_2H_3COONa(aq) + H_2O(l)\ \Delta H = -55.2\ \text{kJ mol}^{-1}$

Weak acids such as ethanoic acid or weak bases such as ammonia solution, are **only partially ionised in solution**. Therefore, <u>**some energy is used up to ionise them**</u> before the neutralisation takes place.

The enthalpy change obtained when a **weak acid reacts with a weak base is even lower**. For example, when ethanoic acid reacts with ammonia solution, only 50.4 kJ of heat is evolved.

$CH_3COOH(aq) + NH_4OH(aq) \longrightarrow CH_3COONH_4(aq) + H_2O(l)\ \Delta H = -50.4\ \text{kJ mol}^{-1}$

Standard Conditions For Measuring Enthalpy Changes

The standard conditions of temperature and pressure adopted are 25°C (298 K) and one atmosphere pressure (101.325 kPa) respectively. Values of enthalpy changes that are measured at these conditions are called standard enthalpy changes.

Standard enthalpy changes are given a special symbol; **ΔHθ**. The symbol **'θ'**denotes '**standard'** and it implies a pressure of 101.325 kPa (i.e. one atmospheric pressure) and a temperature of 298 K (i.e., 25°C). A **subscript** is also added to the symbol to indicate the **type of enthalpy change involved.** Thus:

- **ΔHθc r**efers to the standard **molar enthalpy change of combustion,** e.g.,

$CH4(g) + 2O2\ O \longrightarrow CO2(g) + 2H2(l)\ \Delta H\vartheta C(CH4) = -890\ kJ\ mol{-1}$

- **ΔHθneut** Refers to the standard **molar enthalpy change of neutralisaiton** e.g.

$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H2O(l)\ \Delta H\vartheta neut = -58\ kJ\ mol{-1}$

- **ΔHθsoln** refers to the standard **molar enthalpy change of dissolution (solution)**

e.g., $NaNO3\ (s) + aq \longrightarrow NaNO3\ (aq)\ \Delta H\vartheta Soln(NaNO3) = +21\ kJ\ mol{-1}$

The value of **ΔH** in a thermochemical equation refers to **molar quantities** shown in the equation. For example, the following equations show the standard molar enthalpy change of formation of water.

$HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H2O\ (l)$

$H+(aq) + OH-(aq) \longrightarrow H2O(l)\ \Delta H = -57.2\ kJ\ mol-1$

Note that **only one mole of H2O** is shown in the above equation.

$H2SO4(aq) + NaOH(aq) \longrightarrow Na2SO4(aq) + 2H2O(l)$

The equation for the formation of two moles of would be:

$2H+(aq) + 2OH-(aq) \longrightarrow 2H2O(l)\ \Delta H = -114.4\ kJ-1$

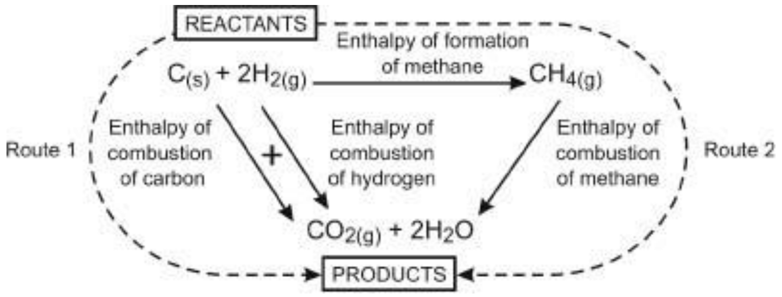
The **amount of heat evolved when two moles are formed is therefore twice what is evolved when one mole is formed.** Note that the units for **ΔHθ is the kilo joule per mole (kJ mol**-1**).**

Hess's Law

Hess's Law of constant heat summation states that *The energy changes in converting reactants to products is the same regardless of the route by which the chemical change occurs.*

This law is applied in determining theoretically the enthalpy change for some reactions that cannot be determined experimentally because the reactions cannot take place under normal conditions.

An **energy cycle diagram** that links the reactants and products is used. This provides various routes for which a reaction can occur and can be used to determine various enthalpies, for example the enthalpy of formation of methane.



Energy cycle diagram incorporating enthalpy of formation of methane

In the energy cycle above, there is **more than one way** of converting carbon and hydrogen into carbon(IV) oxide and water respectively. Carbon and hydrogen can either be **burned directly (route 1)** or they can **first be combined to form methane which can then be burned (route 2).**

Whether **route 1** or **route 2** is followed, the **end product is the same**. This means that the **energy changes for route 1 are equal to the energy changes for route 2.** This must be so because of the **law of conservation of energy**. The enthalpy of formation of methane can then be calculated from the expression:

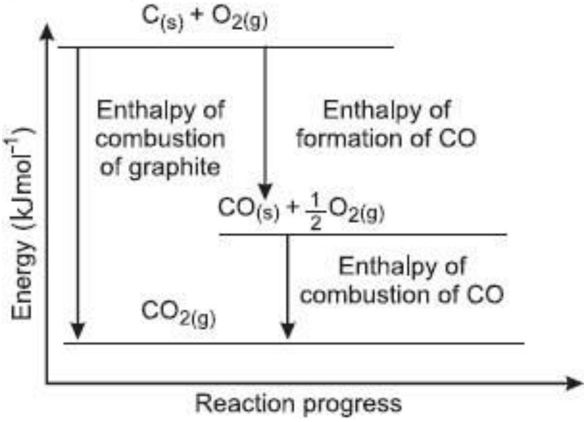
$$\boxed{\text{Enthalpy of formation of methane}} = \boxed{\text{Enthalpy of combustion of carbon}} + \boxed{\text{Enthalpy of combustion of hydrogen} \times 2} - \boxed{\text{Enthalpy of combustion of methane}}$$

= – 393 + (–286 × 2) – (–890)

= –965 + 890

= –75 kJ mol-**1**

An **energy level diagram** can also be used. For example, to determine the enthalpy of formation of **carbon (II) oxide**, the following energy level diagram can be used.



Energy level diagram for formation of carbon(II) oxide

From the energy level diagram, the enthalpy of formation of carbon(II) oxide would be given by:

Enthalpy = Enthalpy of combustion of C(s) – Enthalpy of combustion of CO(g)

ΔHϑf CO(g) = ΔHϑC C(s) - ΔHϑC CO(g)

= –393 – –283

= – 170 kJ mol-*1*

Application of Hess's law relies on the **information provided by thermochemical equations for the reaction in question.**

Information provided by the thermochemical equations can be used to **draw energy cycle and energy level diagrams** linking the equations for a reaction as well as determine the **enthalpy change** for the reaction.

Worked Examples

1. The thermochemical equation for the combustion of carbon (graphite) and carbon(II) oxide are as follows:

$C(s) + O_2(g) \longrightarrow CO_2(g)$ $\Delta H = -393$ kJ mol^{-1}

$CO(g) + \frac{1}{2} O_2(g) \longrightarrow CO(g)$ $\Delta H = -283$ kJ mol^{-1}

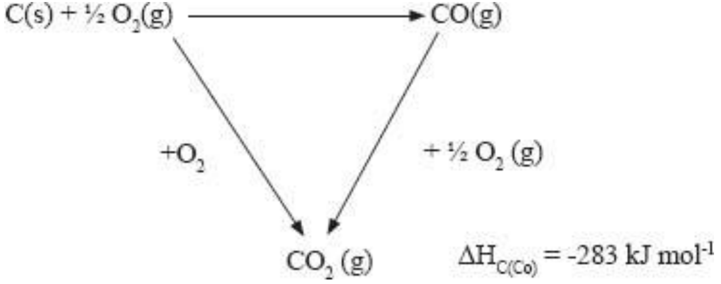
(a) Use the information provided to draw:

(i) An energy cycle diagram.

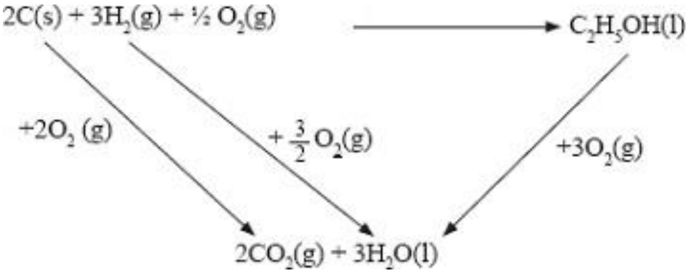
(ii) An energy level diagram for the reaction.

(b) Use the energy cycle and energy level diagrams drawn in (a) to determine the enthalpy of formation of carbon(II) oxide.

These equations can be linked in an energy cycle diagram as follows:



The energy cycle diagram for the formation and combustion of ethanol is :



From the energy level diagram:

Enthalpy of formation of ethanol = 2× Enthalpy of combustion of carbon + 3× Enthalpy of combustion of hydrogen – Enthalpy of combustion of ethanol

$\Delta H^{\theta}_{f}(C_2H_5OH(l)) = 2\Delta H^{\theta}_{c}\,(C(s) + 3\Delta H^{\theta}_{c}\,(H_2(s) - \Delta H_{c}(C_2H_5OH(l)$

$$= 2 \times (-393) + 3 \times (-286) - (-1368)$$

$$= -786 + (-858) - (-1368)$$

$$= -1644 + 1368$$

$$= -276 \text{ kJ mol}^{-1}$$

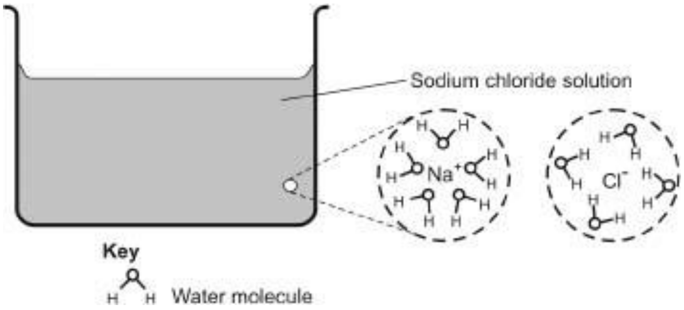
Relationship Between Heat of Solution, Hydration and Lattice Energy

Lattice energy is *the energy change when one mole of an ionic compound is formed from its constituent ions in the gaseous state.*

When an ionic substance dissolves, energy equivalent to the lattice energy is absorbed. For example, when sodium chloride is dissolved in water, the ionic bonds are first broken to release the gaseous ions.

$$NaCl(s) \xrightarrow{water} Na^+(g) + Cl^-(g) \qquad \Delta H^\circ = +776\ kJ\ mol^{-1}$$

The sodium and chloride, ions get surrounded by several water molecules. The water molecules surrounding the sodium ions arrange themselves in such a way that the partially negative oxygen is attracted to the positive sodium ion and the partially positive hydrogen is attracted to the negative chloride ions and the ions thus become hydrated.



The hydration process involves the formation of new bonds between the polar water molecules and the ions. The process is always exothermic. The energy given out is known as hydration energy.

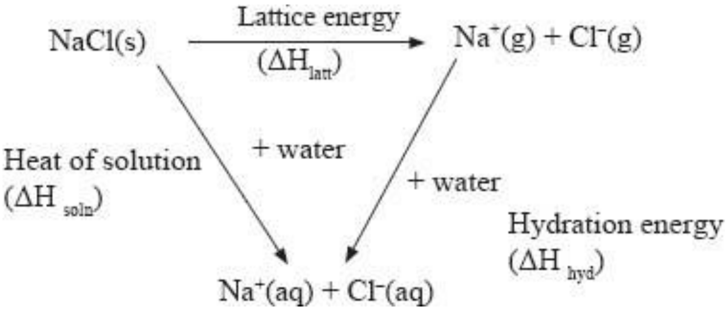
$$Na^+(g) \xrightarrow{H_2O} N^+(aq) \qquad \Delta H^\ominus = -390\ kJ\ mol^{-1}$$

$$Cl^-(g) \xrightarrow{H_2O} Cl(aq) \qquad \Delta H^\ominus = -384\ kJ\ mol^{-1}$$

Hydration energy is defined as the energy change that occurs when one mole of gaseous ions become hydrated.

The overall enthalpy change of solution depends on whether the endothermic or the exothermic process is larger.

For sodium chloride, the endothermic process is greater than the exothermic process. Therefore, the enthalpy of solution of sodium chloride has a positive value. The following energy cycle diagram can be used to calculate the heat of solution of sodium chloride.



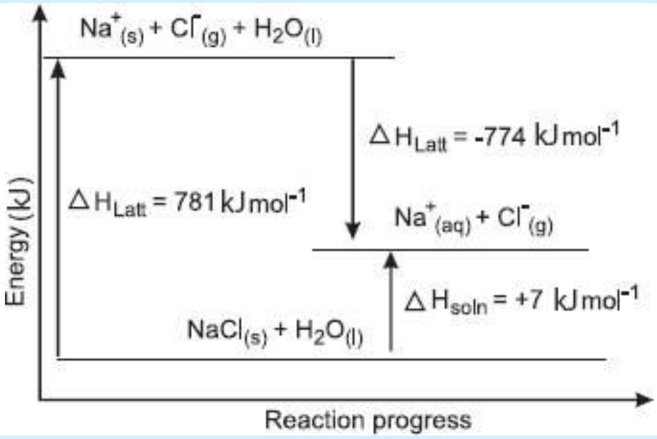
From the energy cycle diagram,

Heat of solution = Lattice energy + hydration energy

ΔHsol = **ΔH** = **ΔH**hyd

= 781 + (-774)

= + 7 kJ mol**-1**



Energy level diagram for dissolving sodium chloride

The lattice energy is given positive values in the calculations above because bond breaking is an endothermic process.

The hydration energy for sodium chloride is the sum of the separate hydration energies of sodium ions and chloride ions, i.e.,

ΔH**hyd **NaCl = ΔH**hyd **(Na+) + ΔH**hyd **(Cl-)

= –390 + (–384)

= –774 kJ

When ions are hydrated, the amount of energy released depends on the size of the ions and the charge on the ions. .

Enthalpies of hydration of some common ions

Ion	ΔH_{hyd} (kJ mol^{-1})	Ion	ΔH_{hyd} (kJ mol^{-1})
H^+	−1075	F^-	−457
Li^+	−449	Cl^-	−384
Na^+	−390	Br^-	−351
K^+	−305	I^-	−307
NH^{+4}	−281	OH^-	−460
Mg^{2+}	−1891		
Ca^{2+}	−1562		
Al^{3+}	−4613		

Lattice energies of some ionic common compounds

Compound	***ΔH kJ mol**-1*
NaF	–915
NaCl	–781
Na Br	–743
NaI	–699
MgCl**2**	–2489
MgO	–3933
AgCl	–890

Fuels

A fuel *is a substance that produces useful energy when it undergoes a chemical or nuclear reaction.*

Fuels can be **solids** such as **nuclear fuel, coke, coal, charcoals and wood.**

Other fuels are **liquids** such as **petrol, kerosene and diesel oil,** or **gases** like **natural gas, biogas, water gas and petroleum gas.**

Heating Values of Fuels

The **heating value of a fuel** is ***the amount of heat energy given out when a unit mass, or a unit volume of a fuel is completely burned in oxygen.***

The heating value of a fuel is obtained by **dividing the molar enthalpy of combustion by the formula mass of the fuel.**

Heating value has the units kJ g**-1** (kilojoules per gram).

For example, **ethanol** has an enthalpy of combustion of **-1360 kJ/mol.** On dividing this by the formula mass of ethanol, **46**, we get the heating value of ethanol as **30 kJ g**-1**.**

Heating values of various fuels

Fuels	***Heating value in J/g***
Solid Fuels	
Charcoal	33
Coal	29
Wood	17
Liquid fuels	
Ethanol	30
Fuel oil	45
Paraffin	48
Gaseous Fuels	
Methane (natural gas)	55
Propane	50
Butane	48

How to choose a fuel

The choice of a fuel depends on the purpose it is to be used for. The following points have to be considered in each case:

- Heating value.
- Ease and rate of combustion.
- Availability.
- Ease of transportation.
- Ease of storage.
- Environmental effects.
- Cost.

For example, wood and charcoal are chosen for domestic heating because:

- They are cheap.
- The are readily available.
- They can be easily transported.
- When they burn they do not produce poisonous products.
- They burn slowly.

Methylhydrazine (CH**3**NHNH**3**) is used for rocket propulsion because:

- It burns very rapidly producing large amounts of gases which in turn create a huge thrust as they escape.
- It has a very high heat of combustion (4740 kJ mol**–1**).
- It ignites easily.

Precautions Necessary When Using Fuels

Improper handling of fuels can result it death and destruction of property. Each type of fuel should be handled with the necessary care and precautions:

- Charcoal stoves should be operated in well ventilated rooms to avoid poisoning by carbon(II) oxide.
- Vehicle engines should not be left running in closed garages to avoid poisoning by carbon(II) oxide.
- Gas cylinders should be stored in well ventilated rooms far from heat sources. Only certified cylinders should be used for gas transportation.
- Fuel storage facilities should be located far away from populated areas.
- People should keep off from fuel spilled from tankers.

Environmental Effects of Fuels

Fossil fuels such as coal and petroleum contain carbon, nitrogen and sulphur compounds which on burning produce poisonous gases such as sulphur(IV) oxide, sulphur(VI) oxide, carbon(II) oxide and nitrogen(IV) oxide.

The sulphur and nitrogen oxides dissolve in rain-water to produce acid rain. Acid rain:

- Wears limestone buildings and statues.
- Corrodes iron sheets, iron gates and other metallic structures.
- Acidifies lakes leading to death of plants and animals in the lakes.
- Leaches nutrients from plant leaves leading to their death.
- Leaches minerals from the soil leading to poor soil for agricultural activities.

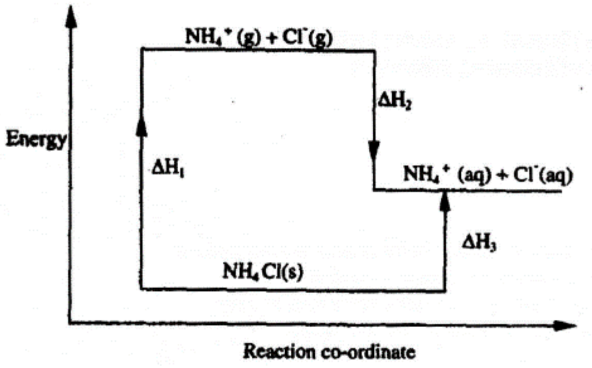
Carbon(IV) oxide and unburned hydrocarbons from fuels contribute to global warming. Global warming causes polar ice to melt causing inundation of low lying coastal lands. Other effects of global warming include unpredictable weather patterns and climatic changes.

Measures that are being taken to reduce pollution:

- Designing zero emission vehicles which use solar energy or electrical energy.
- Fitting catalytic converters to vehicle exhaust systems. These catalytic converters convert nitrogen oxides to harmless nitrogen, carbon(II) oxide to carbon(IV) oxides, unburned hydrocarbons to water and carbon(IV) oxide.
- Designing engines which uses unleaded petrol.
- Electronically controlling the quantity of air mixed with fuel to ensure more complete combustion.
- Adding fuels which contain oxygen in their molecules to petrol. For example, methanol and ethanol may be added to petrol. These fuels reduce the quantities of carbon(II) oxide and unburned hydrocarbons emitted.
- Encourage other means of transport such as the use of bicycle and electric trains.

1. 2006 Q 28 P1

Study the diagram below and answer the questions that follow.



(a) What do ΔH_1 and ΔH_2 represent? (2 marks)

(b) Write an expression to show the relationship between ΔH_1, ΔH_2 and ΔH_3. (1 mark)

2. 2006 Q 2a P2

In an experiment to determine the molar heat of reaction when magnesium displaces copper, 0.15g of magnesium powder were added to 25.0cm^3 of 2.0M copper (II) chloride solution. The temperature of copper (II) chloride solution was 25 °C, while that of the mixture was 43 °C.

(i) Other than increase in temperature, state and explain the observations which were made during the reaction. (3 marks)

(ii) Calculate the heat change during the reaction (specific heat capacity of the solution= 4.2jg^{-1}K^{-1} and the density of the solution = 1g/cm^3)

(2 marks)

(iii) Determine the molar heat of displacement of copper by magnesium. (Mg=24.0). Write the ionic equation for the reaction.
(1 mark)

(iv) Sketch an energy level diagram for the reaction. (2 marks)

3. 2007 Q 10 P1

The thermo chemical equations for the formation of hydrogen peroxide under standard conditions are:

$H_2(g) + O_2(g) \longrightarrow H_2O_2(g)$; $\Delta H^\theta_f = -133 kJmol^{-1}$

$H_2(g) + O_2(g) \longrightarrow H_2O(l)$; $\Delta H^\theta_f = -188 kJmol^{-1}$

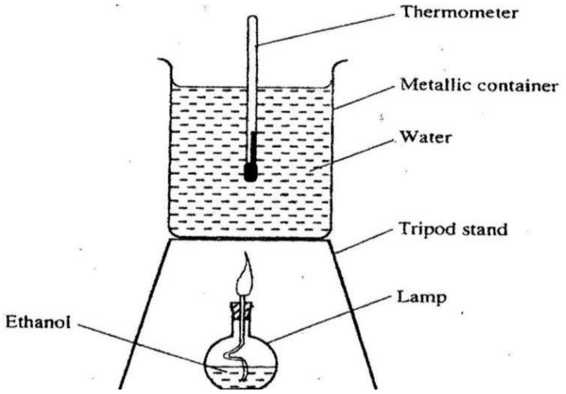
Write the thermo chemical equation for the molar heat of vaporization of hydrogen peroxide.

(2marks)

4. 2007 Q 1 P2

(a) State two factors that should be considered when choosing fuel for cooking. (2 marks)

(b) The diagram below represents a set–up that was used to determine the molar heat of combustion of ethanol.



During the experiment, the data given below was recorded:

Volume of water	***450 cm^3***
Initial temperature of water	***25.0 °C***
Final temperature of water	***46.5 °C***
Mass of ethanol + Lamp before burning	***125.5 g***
Mass of ethanol + lamp after burning	***124.0 g***

Calculate the:

(i) Heat evolved during the experiment. (density of water = 1g/cm^3, specific heat capacity of water = 4.2 $Jg^{-1}K^{-1}$)
(3 marks)

(ii) Molar heat of combustion of ethanol. (C = 12.0, O = 16.0, H=1.0) (2 marks)

(c) Write the equation for the complete combustion of ethanol. (1 mark)

(d) The value of the molar heat of combustion of ethanol obtained in (b) (ii) above is lower than the theoretical value. State two sources of error in the experiment.

(2 marks)

5. 2008 Q 7 P2

(a) Define the standard enthalpy of formation of a substance (1 mark)

(b) Use the thermo chemical equations below to answer the questions that follow.

1. $C_2H_6(g) + {}^{7}/_{2}O_2(g) \longrightarrow 2CO_2(g) + 3H_2O\ (L);\ \Delta H_1 - 1560kjmol^{-1}$
2. $C_{(graphite)} + O_2\ (g) \longrightarrow CO_2(g);$ $\Delta H_2 = -394kJmol^{-1}$
3. $H_2(g) + {}^{1}/_{2}O_2\ (g) \longrightarrow H_2O(g)$ $\Delta H_3 = -286kJmol^{-1}$

(i) Name two types of heat changes represented by ΔH_3 (2 marks)

(ii) Draw an energy level diagram for the reaction represented by equation 1. (2 marks)

(iii) Calculate the standard enthalpy of formation of ethane (2 marks)

(iv) When a sample of ethane was burnt, the heat produced raised the temperature of 500g of water by 21.5 K, (specific heat capacity of water = $4.2Jg^{-1}K^{-1}$). Calculate the:

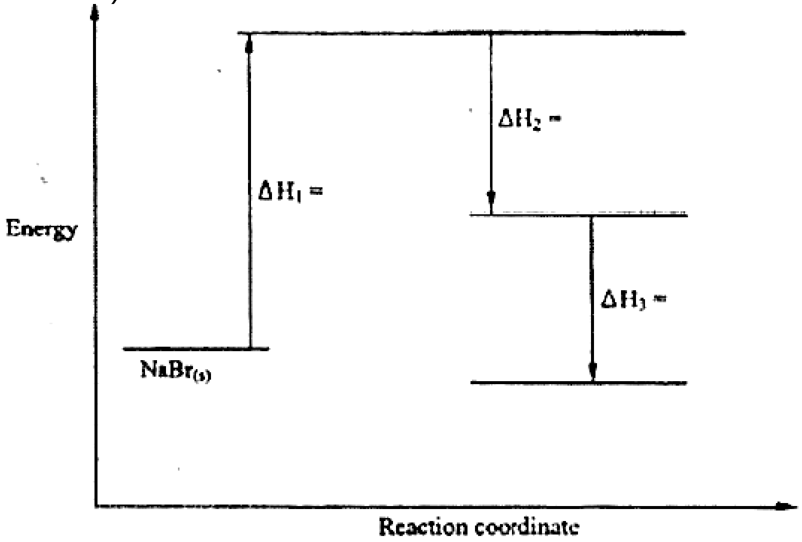
I. Heat change for the reaction (2 marks)
II. Mass of ethane was burnt. (R.F.M of ethane= 30) (2 marks)

6. 2009 Q 9 P1

(a) What is meant by molar heat of solution? (1 mark)

(b) The lattice energy of sodium bromide and hydration energies of sodium and bromide ions are: 733, 406 and 335 $kJmol^{-1}$ respectively

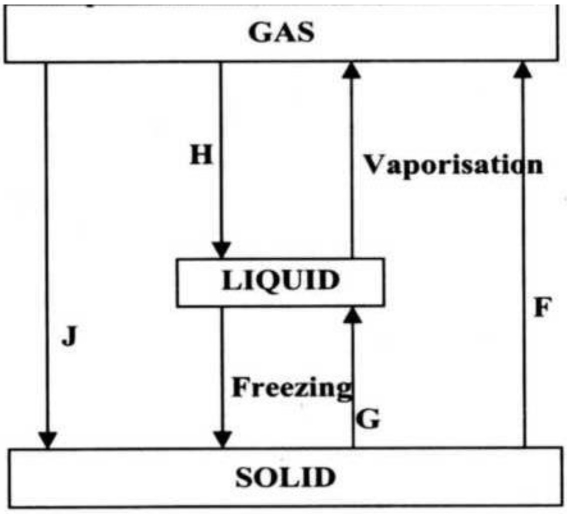
(i) Complete the energy cycle diagram above by inserting the values of ΔH_1, ΔH_2 and ΔH_3. (1½ marks)



(ii) Determine the molar heat of solution of solid sodium bromide. (½ mark)

7. 2009 Q 5 P2

(a) **Figure 3** shows the changes that take place between state of matter. Some of them Have been identified and others labelled.



(i) Give the names of the process:

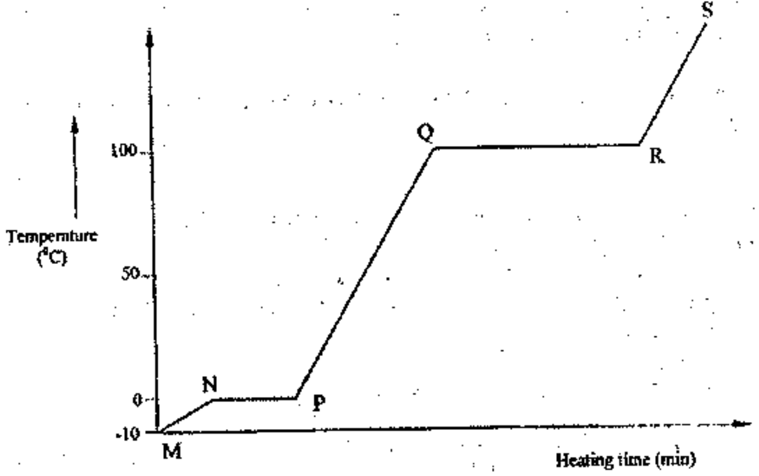
I. H (1 mark)

II. G (1 mark)

(ii) Name one substance that can undergo process **F** When left in an open container in the laboratory. (1 mark)

(iii) The process J is called deposition. Using water is an example; write an equation that represents the process of deposition. (1 mark)

(b) Figure 4 shows the heating curve for water



(i) Give the names of the intermolecular forces of attraction in the segments.

I. **MN** (1 mark)

II. **RS** (1 mark)

(ii) The heat of fusion and evaporation of water are 334.4 Jg^{-} and 1159.4 Jg^{-} respectively.

I. Explain why there is a big difference between the two. (2 marks)

II. How is the difference reflected in the curve? (1 mark)

(b) Coal, oil and natural gas are major sources of energy. They are known as fossil Hydrogen is also a source of energy.

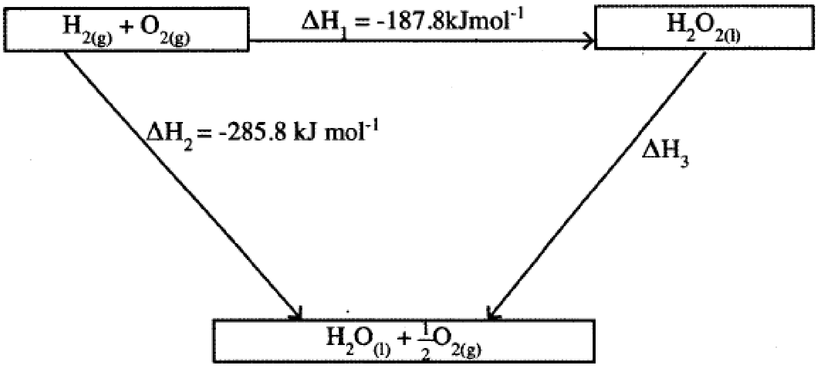
(i) State and explain two reasons why hydrogen is a very attractive fuel compared to fossil fuels (3 marks)

(ii) State one disadvantage of using hydrogen fuel instead of fossil fuel.

(1mark)

8. 2010 Q 10 P1

The figure below shows an energy cycle.



(a) Give the name of the enthalpy change ΔH_1. (1 mark)

(b) Determine the value of ΔH_3. (2 marks)

9. 2010 Q 12 P1

A beaker contained 75.0cm^3 of aqueous copper (II) sulphate at 23.7 °C when scrap iron metal was added to the solution, the temperature rose to 29.3 °C.

(a) Write an ionic equation for the reaction that took place. (1 mark)

(b) Given that the mass of copper deposited was 5.83g, calculate the molar enthalpy change in kJmol-1. (Specific heat capacity of solution = 4.2Jg^{-1} K^{-1}, density of solution 1.0gcm^{-3}, Cu = 63.5) (2 marks)

10. 2010 Q 4 P2

(a) 50cm^3 of 1M copper (II)sulphate solution was placed in a 100cm^3 plastic beaker. The temperature of the solution was measured. Excess metal A powder was added to the solution, the mixture stirred and the maximum temperature was repeated using powder of metals B and C. The results obtained are given in the table below:

	A	**B**	**C**
Maximum temperature (°C)	26.3	31.7	22.0
Initial temperature (°C)	22.0	22.0	22.0

(i) Arrange the metal **A**, **B**, **C** and copper in order of reactivity starting with the least reactive. Give reasons for the order. (2 marks)

(ii) Other than temperature change, state one other observation that was made when the most reactive metal was added to the copper (II) sulphate solution. (1 mark)

(b) The standard enthalpy change of formation of methanol is -239 kJmol^{-1}.

(i) Write the thermal chemical equation for the standard enthalpy change of formation of methanol. (1 mark)

(ii) Methanol is manufactured by reacting carbon (II)oxide with hydrogen at 300 °C and a pressure of 250 atmospheres.

The equation for the reaction is:

$$CO_{(g)} + 2H_{2(g)} \rightleftharpoons CH_3OH_{(g)}$$

(I) How would the yield of methanol be affected if the manufacturing process above is carried out at 300 °C and a pressure of 400 atmosphere? Explain (2 marks)

(II) Use the following data to calculate the enthalpy change for the manufacture of methanol from carbon(II)oxide and hydrogen. (3 marks)

$$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g);\ \Delta H^\theta = -283kJmol^{-1}$$

$$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(i);\ \Delta H^\theta = -286kJmol^{-1}$$

$$CH_2OH(l) + \frac{3}{2}O_2(g) \longrightarrow CO_2(g) + 2H_2O(l);\ \Delta H^\theta = -715kJmol^{-1}$$

(iii) The calculated enthalpy change in part b(ii) (II) above differ from the standard enthalpy change of formation of methanol. Give a reason. (1 mark)

11. 2011 Q 23 P1

The thermal chemical reaction between carbon and sulphur is as shown by the equation below

$$C(s) + 2S(s) \longrightarrow CS_2(l) \qquad \Delta H = +117.0\ kJmol^{-1}$$

On the grid below, sketch and label the energy diagram for the reaction. (2 marks)

12. 2011 Q 7 P2

(a) What is meant by molar heat of combustion? (1 mark)

(b) State the Hess's Law. (1 mark)

(c) Use the following standard enthalpies of combustion of graphite, hydrogen and enthalpy of formation of propane.

ΔH^θ_c (Graphite) = -393 kJmol^{-1}
ΔH^θ_c ($H_2(g)$) = -286 kJmol^{-1}
ΔH^θ_f (C_3H_8 (g)) = -104 kJmol^{-1}

(i) Write the equation for the formation of propane. (1 mark)

(ii) Draw an energy cycle diagram that links the heat of formation of propane with its heat of combustion of graphite and hydrogen. (3 marks)

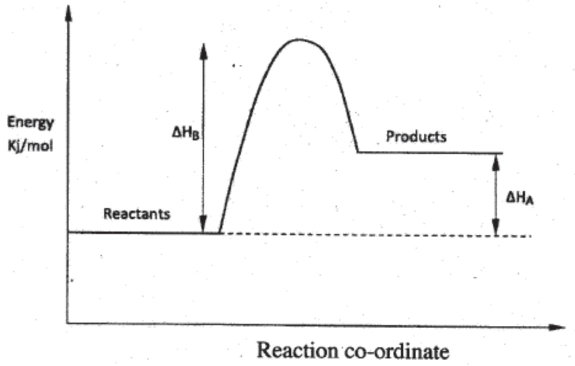
(iii) Calculate the standard heat of combustion of propane. (2 marks)

(d) Other than the enthalpy of combustion, state one factor which should be considered when choosing a fuel. (1 mark)

(e) The molar enthalpies of neutralization for dilute hydrochloric acid and dilute nitric (V) acid are -57.22 kJmol- while that of ethanoic acid is -55.2 kJmol-. Explain this observation. (2 marks)

13. 2012 Q17 P1

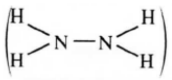
Study the energy level diagram below and answer the questions that follow.



(a) Give the name of ΔH_a (2 marks)

(b) How can ΔH_b be reduced? Give a reason (2 marks)

14. 2013 Q11 P1

Hydrazine gas,  burns in oxygen to form nitrogen gas and steam.

(a) Write an equation for the reaction (1 mark)

(b) Using the bond energies given below, calculate the enthalpy change for the reaction in (a) above. (2 marks)

Bond	**Bond energy (kJ per mole)**
N ≡ M	944
N - N	163
N - H	388
O = O	496
H - O	463

15. 2013 Q27 P1

A student investigated a property of acids M and N by reacting equal volumes of acid M and N of the same concentration with equal volumes of 2M potassium hydroxide. The results were recorded in the table below.

Acid	Rise in temperature(Δ)K
M	4
N	2

(a) Which of the acids is likely to be a weak acid? Explain. (2 marks)

(b) Write the equation for the reaction between ethanoic acid and potassium hydroxide. (1 mark)

16. 2013 Q2 P2

(a) (i) what is meant by the term Enthalpy of formation? (1 mark)

(ii) The enthalpies of combustion of carbon, methane and hydrogen are indicated below:

$C(s) + O_2(g) \longrightarrow CO_2(g)$; $\Delta H = -393\ kJ\ mol^{-1}$

$H_2(g) + O_2(g) \longrightarrow H_2O\ (l)$; $\Delta H = -286\ kJ\ mol^{-1}$

Enthalpy of combustion of $CH_4 = -890\ kJ\ mol^{-1}$

I. Draw an energy cycle diagram that links the enthalpy of formation of methane to enthalpies of combustion of carbon, hydrogen and methane (2 marks)

II. Determine the enthalpy of formation of methane (2 marks)

(b) An experiment was carried out where different volumes of dilute hydrochloric acid and aqueous sodium hydroxide both at 250c were mixed and stirred with a thermometer. The highest temperature reached by each mixture was recorded in the table below

Volume of hydrochloric acid (cm^3)	5	10	15	20	25	30	35	40	45
Volume of sodium hydroxide (cm^3)	45	40	35	30	25	20	15	10	5
Highest temperature of mixture(°C)	27.2	29.4	31.6	33.8	33.6	31.8	30.0	28.4	26.6

(i) On the grid provided. plot a graph of highest temperature (vertical axis) against volume of hydrochloric acid (horizontal axis) (3 marks)

(ii) Using your graph, determine the

I. Highest temperature reached; (½ mark)

II. Volume of acid and base reacting when highest temperature is reached. (½ mark)

(iii) Calculate the amount of heat liberated during the neutralization process. (specify heat capacity is 4.2 $Jg^{-1}K^{-1}$ and the density of solution is 1.0 gcm^{-3}) (2 marks)

(c) The molar enthalpy of neutralization between hydrochloric acid and ammonia solution was found to be -52.2 $kJmol^{-1}$, while that of hydrochloric acid and sodium hydroxide was - 57.1 $kJmol^{-1}$. Explain the difference in these values. (2 marks)

17. 2014 Q14 P1

When 20cm^3 of 1 M sodium hydroxide was mixed with 20cm^3 of 1 M hydrochloric acid, the temperature rose by 6.7 °C. Assuming the density of the solution is 1 g/cm^3 and the specific heat capacity of the solution is 4.2$Jg^{-1}K^{-1}$;

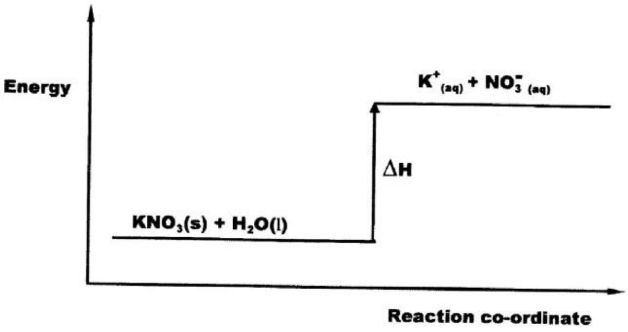
(a) Calculate the molar heat of neutralization; (2 marks)

(b) When the experiment was repeated with 1 M ethanoic acid, the temperature changes was found to be lower than that with 1 M hydrochloric acid. Explain. (1 mark)

18. 2015 Q3 P1

(a) What is meant by lattice energy? (1 mark)

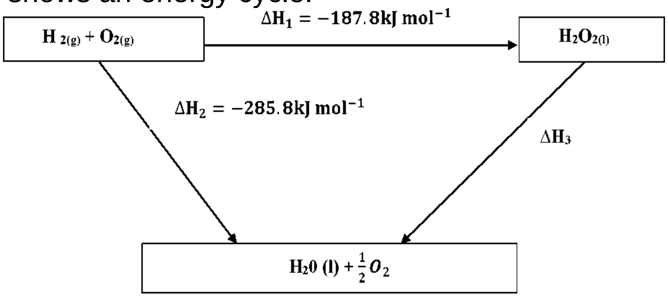
(b) Study the energy level diagram below and answer the question that follows



What type of reaction is represented by the diagram? (1 mark)

19. 2016 Q3 P1

The figure below shows an energy cycle.



(a)Give the name of the enthalpy change ΔH_1. (1 mark)

(b) Determine the value of ΔH_3. (2 Marks)

20. 2017 P1 Q19.

The following procedure was used to investigate the temperature changes that occur when sodium hydroxide solution is added to dilute hydrochloric acid.

(i) Place the acid in a glass beaker and record its temperature.
(ii) Add a known volume of sodium hydroxide solution.
(iii) Stir the mixture and record the highest temperature reached.
(iv) Repeat steps (ii) and (iii) with different volumes of sodium hydroxide solution.

(a) State two factors that must be kept constant in this experiment. (1 mark)

(b) Explain how the use of a polystyrene cup will affect the results. (1 mark)

21. 2018 P1 Q 18.

(a) Define molar heat of displacement. (1 mark)

(b) The following ionic equation represents the reaction between metal Y and an aqueous solution of Z^{2+}.

Z^{2+} (aq) + Y(s) Z(s) + Y2 (aq); ΔH = -ve

Draw an energy level diagram to represent the reaction. (2 marks)

22. 2019 P1 Q17.
The heat of solution and hydration energy of potassium chloride is -17.2 kJ and -689 kJ respectively.
Calculate the lattice energy of potassium chloride. (2 marks)

23. 2019 P1 Q 18.
Use the information in **Table 2** to answer the questions that follow.

Bond	**Bond energy (KJ mol⁻)**
C-H	412
Cl-Cl	242
C-Cl	338
H-Cl	431

(a) State what is meant by heat of reaction. (1 mark)

(b) Calculate the heat change when one mole of methane reacts completely with excess chlorine in the presence of UV light.
(2 marks)

24. 2019 P2 Q5.
(a) What is meant by a molar heat of neutralisation? (1 mark)
(b) In an experiment to determine the molar heat of neutralisation, 50 cm^3 of 1 M hydrochloric acid was neutralised by adding 10 cm^3 Onions Of dilute sodium hydroxide.
During the experiment, the data in **Table 1** was obtained.

Table 1

Volume of sodium hydroxide (cm^3)	0	10	20	30	40	50	60
Temperature of mixture (°)	25.0	27.0	29.0	31.0	31.0	30.0	29.0

(i) Write the equation for the reaction in this experiment. (1 mark)

(ii) On the grid provided, plot a graph of temperature (Y-axis) against volume of sodium hydroxide (X-axis) added. (3 marks)
(iii) Determine from the graph the:
I. volume of sodium hydroxide which completely neutralises 50 cm^3 of 1M hydrochloric acid.
(1 mark)
II. change in temperature, ∆T, when complete neutralisation occurred. (1 mark)

(iv) Calculate:
I. The heat change, ∆H, when complete neutralisation occurred. (Specific heat capacity = 4.2 Jg^{-1} K^{-1} density of solution 1.0 gcm^{-3}) (2 marks)
II. Molar heat of neutralisation of hydrochloric acid with sodium hydroxide. (1 mark)

(v) How would the value of molar heat differ if 50 cm^3 of 1M ethanoic acid was used instead of 1M hydrochloric acid? Give a reason.
(2 marks)

TOPIC THREE
REACTION RATES AND REVERSIBLE REACTIONS

Table of Contents

Objectives

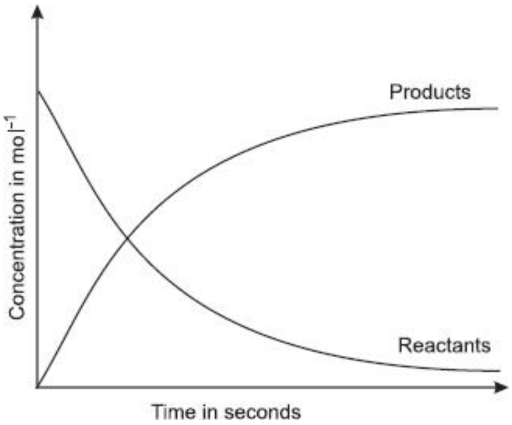
By the end of the topic, the learner should be able to:

(a) Define rate of reaction and explain the term activation energy.
(b) Describe some methods used to measure the rate of reaction.
(c) Explain the effect of different factors on reaction rates.
(d) Illustrate reaction rates graphically and interpret experimental data.
(e) State examples of simple reversible reactions.
(f) Explain chemical equilibrium and the effect of different factors on the position of equilibrium.

REACTION RATES AND REVERSIBLE REACTIONS

Reaction Rates

Reaction rate *is a measure of how much of the reactants are consumed or how much of the products are formed per unit time.*



Relationship between the concentration of reactants and products during a reaction

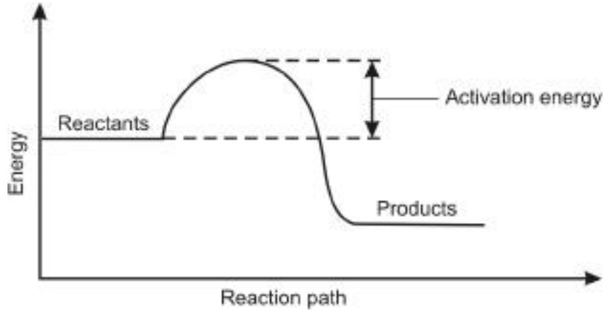
Collision Theory and Activation Energy

In order for a reaction to occur, reacting particles must collide.

Not all collisions results into products. For effective collisions, the reacting particles must have a certain minimum energy.

The minimum energy required by the reacting particles to cause a successful collision to form products is referred to as **Activation Energy**.

This information can be represented in a diagram as shown.



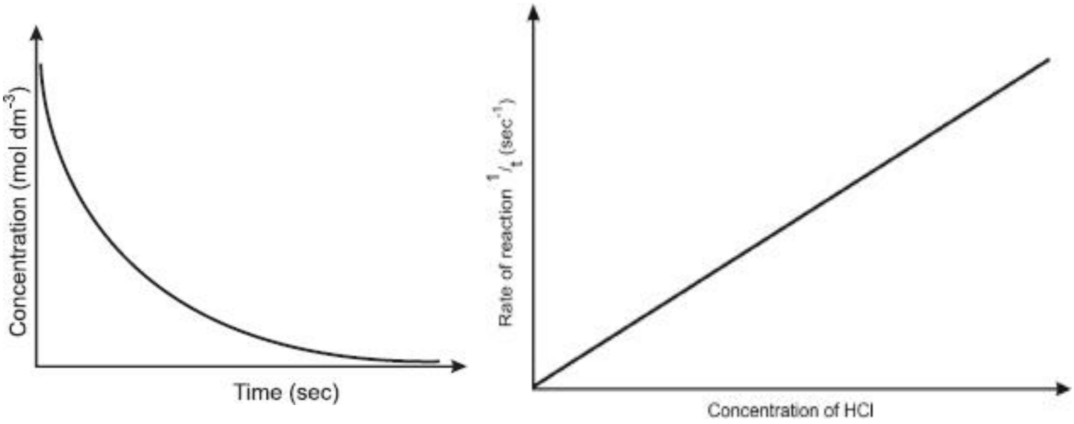
Energy changes during a reaction

Factors Affecting the Rate of a Reaction.

The rate of a reaction depends on concentration, pressure (for gaseous reactants), temperature, surface area(for solid reactants) and catalyst.

(d) Concentration.

When the concentration of the particles of the reactants is high, the number of collisions is also high. This increases the number of successful collisions and hence, increases the rate of a reaction.



Graph of (a) concentration versus time and (b) rate of reaction vs concentration

In a reaction between an acid, say HCl, and a metal eg Magnesium ribbon, it follows that the higher the concentration of the acid, the faster the reaction.

A graph of $\frac{1}{t}$ (s^{-1}) against concentration of the hydrochloric acid shows that the rate increases with increase in concentration

In the experiment, the rate is measured by the time taken for the piece of magnesium ribbon to disappear.

(e) Time.

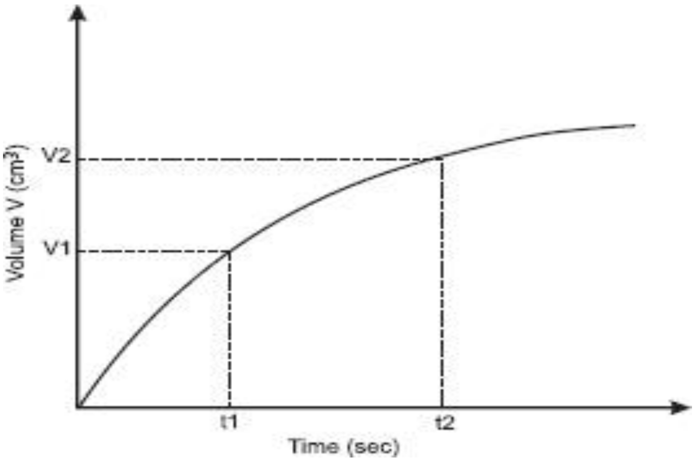
The rate of reaction is inversely proportional to the time taken.

$$\text{Rate} \propto \frac{1}{\text{time}}$$

Consider the reaction between magnesium and dilute hydrochloric acid whichproceeds with the production of hydrogen gas.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

The graph obtained when the volume of the gas produced is plotted against time is a curve.



It is clear that as the reaction progresses, a lower volume of hydrogen gas is liberated within a given time interval. Both the amount of magnesium and the concentration of the acid are decreasing during the experiment. The rate of production of hydrogen gas is highest at the start of the experiment, then it slows down and eventually stops when one or all the reactants are used up.

The average rates of reactionscan be worked out by dividing the change in volume of gas at selected time intervals as follows;

$$\text{Av. rate} = \frac{\text{Volume of hydrogen produced}}{\text{Time taken}}$$

$$= \frac{(V_2 - V_1)\ \text{cm}^3}{t_2 - t_1\ \text{sec}}$$

(f) Temperature

The rate of a reaction increases with a rise in temperature.

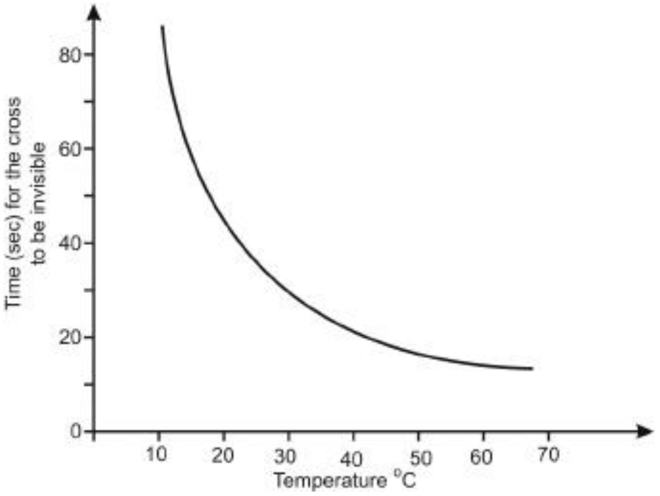
During a reaction, an increase in temperature results in an **increase in the kinetic energy of the particles.** This makes the particles to **move faster and collide more frequently** with sufficient energy to cause more effective collisions per given time. Thus the reaction takes place faster resulting in faster formation of the products.

For example, the reaction between sodium thiosulphate and hydrochloric acid forms a yellow precipitate of sulphur.

$$Na_2S_2O_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + SO_2(g) + H_2O(l) + S(s)$$

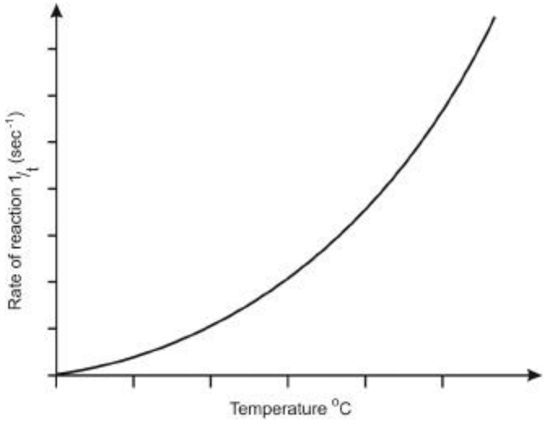
As more and more sulphur is formed, the cross becomes less and less visible until it disappears. The time taken for the cross to be invisible is shorter at higher temperature. Therefore the reaction is faster at higher temperature.

The graph of time versus temperature is a curve. At high temperature less time is required to produce the same amount of sulphur.



It can be seen that the rate of reaction almost doubles with every 10°C temperature rise.

The graph of $\frac{1}{t}$ (sec^{-1}) versus temperature t (°C) shows that the rate increases with a rise in temperature.



(g) Surface area.

The **greater the surface area of contact between reacting particles, the higher the rate of reaction**.

An **increases in the surface area** of the particles **increases the area of contact** of the reacting particlesand hence the reaction proceeds faster.

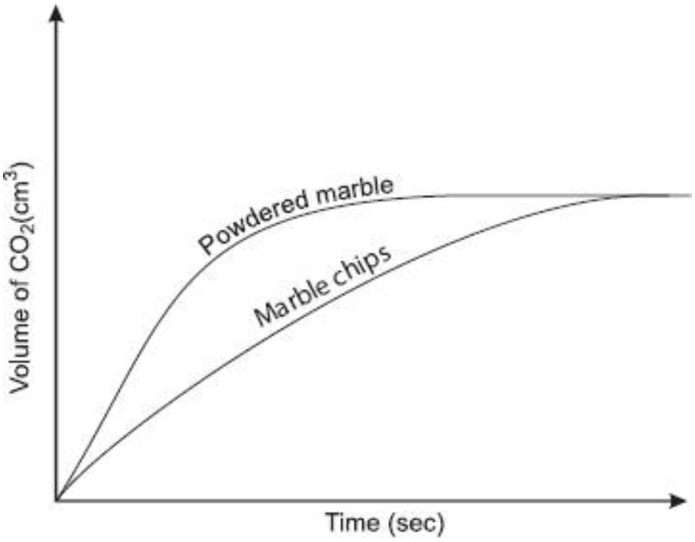
For example, dilute hydrochloric acid reacts with marble to liberate carbon(IV) oxide gas.

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

(iii) The volume of the gas produced per unit time is used to determine the rate of the reaction.

(iv) The reaction is faster at the start due to the high concentration of the reactants.

(v) When a graphs of volume of CO_2 produced is plotted against time for marble chips and powdered marble, it is observed that the curve for powdered marble is steeper than that of the marble chips. This means that carbon(IV) oxide gas is being produced faster with marble powder.



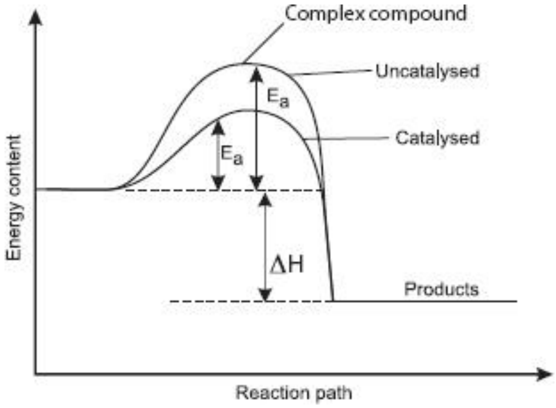
(vi) Both reactions slow down as the reactants are used up and finally both curves flatten out at the same final volume. The total volume of carbon(IV) oxide gas produced in both experiments is the same because the mass of marble used is the same

6. When solid substances are ground into a powder**, the sizes of the particles are reduced. This increases the surface area of the particles thus increasing the area of contact of the reacting particles.** The greater the surface area of contact between reacting particles, the higher the rate of reaction. This is why the reaction of powdered marble with acid takes place more vigorously and reaches completion faster.

(h) Catalysts

A catalystis a substance which alters the rate of a chemical reaction but remains chemically unchanged at the end of the reaction.

Catalysts **lower the activation energy** of the of reaction.They **provide a shorter route** for reactions, hence **minimising the energy required for the reaction to proceed** and hence increases the initial rate of a reaction.

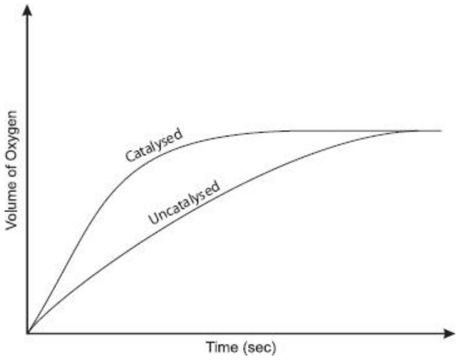


Comparison of catalysed and uncatalysed reactions

For example, Hydrogen peroxide decomposes to form water and oxygen gas.

$$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$$

The rate of the decomposition can be determined by using the volume of oxygen produced per unit time. The production of oxygen is very slow at room temperature. However, addition of manganese(IV) oxide,which acts as a catalyst, makes the reaction faster and a large volume of oxygen is collected within a short time. If the volume of oxygen collected in each case is plotted against time, the curves obtained are as shown below.



The mass of manganese(IV) oxide remains unchanged at the end of the experiment since it is not used up during the reaction.

A particular catalyst alters the rates of certain reactions but not all reactions. Some common catalysts are shown below.

Catalyst	**Industrial process**
Iron	Manufacture of ammonia in Haber process
Vanadium(V) oxide platinum	Manufacture of sulphuric acid contact process
Platinum	Manufacture of nitric(V) acid

The energies of the reactants and the products remain the same but the activation energy is lowered.

(i) Light

Light influences the rate of some chemical reactions by **energising some of the particles** involved. This **increases the number of effective collisions per unit time thus increasing the rate of reaction.**

For example, Potassium bromide reacts with silver nitrate to form a pale yellow precipitate of silver bromide.

$$KBr(aq) + AgNO_3(aq)) \longrightarrow AgBr(s) + KNO_3(aq)$$

In the presence of light, silver bromide decomposes to form metallic silver and bromine. Inthe presence of light, the colour of the precipitate changes from pale yellow to grey due to the formation of metallic silver. In the absence of light, there is no noticeable colour change of the precipitate .

The degree of change depends on the intensity of light falling on the test-tube. The rate of decomposition of the silver bromide is greater when exposed to more light.

The following are some examples of photochemical reactions.

(a) **Photography**

Photographic plates are normally coated with silver bromide. Processing of films is done in a dark room to avoid decomposition of the silver bromide.

(b) **Reaction between chlorine and hydrogen**

Under ordinary temperature and pressure, chlorine and hydrogen do not react in darkness. However exposed to direct light, the reaction is explosive.

$$Cl_2(g) + H_2(g) \xrightarrow{UV\ light} 2HCl(g)$$

(c) **Reaction between bromine and methane**

Bromination of methane does not take place in darkness. But in the presence of light the reaction takes place rapidly.

$$CH_4(g) + Br_2(l) \xrightarrow{UV\ light} CH_3Br(g) + HBr(g)$$

(d) **Photosynthesis**

During photosynthesis, carbon(IV) oxide and water are converted into sugar and oxygen which are useful for plant growth.

(j) Effect of Pressure on the Rate of Reaction

Pressure **has no effect** on reactions involving solids and liquids. It **only affects reactions in which either the reactant or products are gases.**

Increasing the pressure of reacting gases by compressing **brings the gaseous particles closer** thus increasing the rate of effective collisions. This results in an **increase in the rate of a reaction.**

$$6CO_2(g) + 6H_2O\ (l) \xrightarrow{Light} C_6H_{12}O_6(aq) + 6O_2(g)$$

Reversible Reactions

Some chemical reactions are reversible, i.e., they can go forward or backward depending on the prevailing conditions.

A reaction which can proceed in both directions is called a **reversible** reaction.

When white solid ammonium chloride is warmed gently in a test-tube, it decomposes into ammonia and hydrogen chloride gases. The process is called **thermal dissociation**.

$$NH_4Cl(s) \longrightarrow NH_3(g) + HCl\ (g)$$

On the cooler part of the test-tube, ammonia and hydrogen chloride **recombine** to form ammonia chloride.

$$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$$

Production of ammonia chloride when ammonia and hydrogen chloride recombine is a **reverse reaction** to the thermal dissociation of ammonia chloride. The **forward** and the **reverse** (backward) are represented by **double arrows** in an equation.

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl\ (g)$$

When hydrated copper(II) sulphate crystals are heated, a white powder is formed. The crystals lose water of crystallisation.

$$\underset{\text{Blue crystals}}{CuSO_4.5H_2O(s)} \longrightarrow \underset{\text{White powder}}{CuSO_4} + \underset{\text{Colourless liquid}}{5H_2O(l)}$$

When the colourless liquid obtained is poured back into the solid residue, the white anhydrous copper(II) sulpahte turns blue.

$$\underset{\text{White powder}}{CuSO_4(s)} + \underset{\text{water}}{5H_2O(l)} \longrightarrow \underset{\text{Blue crystals}}{CuSO_4.5H_2O(s)}$$

A reversible reaction like the one above is represented by double arrows ($\rightleftharpoons$).

$$CuSO_4.5H_2O(s) \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} CuSO_4\ (s) + 5H_2O(l)$$

Otherexamples of reversible reactions are:

(i) $\underset{\text{pink}}{COCl_2.6H_2O(s)} \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \underset{\text{blue}}{COCl_2} + 6H_2O(g)$

(ii) $\underset{\text{grey}}{3Fe(s)} + 4H_2O(g) \rightleftharpoons \underset{\text{brown}}{Fe_3O_4(s)} + 4H_2(g)$

(iii) $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

Equilibrium

Equilibrium is a **state of balance** that occurs when both the forward and backward reactions are taking place at the same rate but in opposite directions.

A dynamic equilibrium is attained when the rate of the forward reaction is equal to that of the backward reaction.

If a sample of ammonium chloride is heated in a closed container, both the forward and backward reactions take place.

$$NH_4Cl(s) \rightleftarrows NH_3(g) + HCl(g)$$

If the reactions attain equilibrium, the sign $\rightleftharpoons$ replaces the sign $\rightleftarrows$ for reversible reactions which cannot attain equilibrium.

The equations of the reaction can now be written as,

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)?$$

Factors Affecting Equilibrium

The state of balance in reversible reactions can be affected by changing physical conditions such as temperature, pressure and concentration. They affect the equilibrium by either favouring the forward or backward reaction.

Le Chatelier's principle.

"When a change in conditions is applied to a system in equilibrium, the system moves so as to oppose that change."

Effect of Change in Concentration on Equilibrium

If substance A and B react to form product C and D, then changing the concentration of any of the components will shift the equilibrium in such a way as to oppose the change.

$$A(aq) + B(aq) \rightleftharpoons C(aq) + D(aq)$$

If the concentration of B is increased, then the reaction will favour the direction which helps in reducing the concentration of B. This is achieved by more of B reacting with A to form C and D. If the concentration of D is increased, the backward reaction will be favoured.

The effect of change of concentration on a reaction at equilibrium can be demonstrated using the chromate/dichromate reaction.

$$\underset{\textit{Yellow}}{2Cr_2O_4^{2-}(aq)} + 2H^+(aq) \rightleftharpoons \underset{\textit{Orange}}{Cr_2O_7^{2-}(aq)} + H_2O(l)$$

Addition of acid causes more chromate (CrO_4^{2-}) ions to change to dichromate ($Cr_2O_7^{2}$) ions making the solution to become orange. Addition of hydroxide ions reverses the reaction leading to the formation of the yellow chromate solution.

$$Cr_2O_7^{2-}(aq) + OH^-(aq) \rightleftharpoons 2CrO_4^{2-}(aq) + H_2O(l)$$

Altering the concentration of any one of the components of the equilibrium mixture disturbs the equilibrium by **making the reaction rate in one direction greater than the other, hence the reaction proceeds predominantly in that direction until the equilibrium is re-established.**

The observations are in agreement with **Le Chatelier's Principle.**

Effect of Pressure on Equilibrium

A change in pressure will only affect equilibrium in reactions involving gases. It has no effect on those reactions in which both the reactants and products are either solids or liquids since they cannot be compressed or expanded.

Pressure is caused by collisions of the gas particles with the walls of the reaction vessel. The more the number of molecules present, the higher the number of collisions and hence the higher the pressure. **If the pressure of a system in equilibrium is increased, the system will favour the direction that has fewer molecules.**

$$\begin{array}{llcl}
3A(g) + & B(g) & \longrightarrow & 2C(g) \\
3 \text{ moles} & 1 \text{ mole} & & 2 \text{ moles} \\
(3+1) \text{ moles} & & & 2 \text{ moles} \\
4 \text{ moles} & & & 2 \text{ moles} \\
4 \text{ vol} & & & 2 \text{ vol}
\end{array}$$

Consider the following reaction involving gases

The left-hand side has higher pressure (4 molecules) than the right hand side (molecules). Applying more pressure will favour the forward reaction since it reduces the pressure. A decrease in pressure favours the direction of reaction that forms more molecules.

The effect of changing pressure can be illustrated using the equilibrium established between nitrogen(IV) oxide and dinitrogen tetroxide.

$$\begin{array}{ccc}
2NO_2(g) & \rightleftharpoons & N_2O_4(g) \\
\textit{Dark-brown} & & \textit{Pale yellow}
\end{array}$$

Consider what happens if the two gases are enclosed in a syringe. If the piston is pushed in, the volume of the gases is reduced hence increasing pressure. The contents of the syringe appear pale yellow. This shows that an increase in pressure favours the forward reaction which forms fewer molecules.

When the piston is withdrawn slowly, the contents of the syringe darken as the volume is increased therefore lowering the pressure. This shows that backward reaction which forms more molecules is favoured by lowering the pressure.

If the number of molecules of reactants equals number of molecules of products, a change in pressure does not have any effect on the equilibrium.

For example:

$$\begin{array}{llcl}
I_2(g) + & H_2(g) & \longrightarrow & 2HI(g) \\
\textit{1 mol} & \textit{1 mol} & & \textit{2 mol}
\end{array}$$

Change in pressure will not affect the position of equilibrium but will help in the quick attainment of the equilibrium. The effects of change in pressure on equilibrium may be summarised as in table below.

Effect of pressure on equilibrium

Reaction	**Effect of pressure equilibrium change on the mixture**	
$N2O4(g) \rightleftharpoons 2NO2(g)$	**Increase**	**Decrease**
	More N2O4 formed (backward reaction favoured).	More NO2 formed (forward reaction favoured).
$N2(g) + 2H3(g) \rightleftharpoons 2NH2(g)$	More of the reactants formed due to increase in no. of molecules.	More NH3 formed due to decrease in volume.
$2SO2(g) O2(g) \rightleftharpoons 2SO2(g)$	More SO2 and O2 formed. Backward reaction forms more molecules.	More SO3 formed Forward reaction forms fewer molecules.
$4NH3(g) + 5O2(g) \rightleftharpoons 4NO(g) + 6H2O(g)$	More NO and H2O formed backward reaction accompanied by decrease in volume.	More NH3 and O2 formed. Forward reaction accompanied by increase in the number of molecules.
$H2(g) + Cl2(g) \rightleftharpoons 2HCl(g)$	No effect. No change in number of molecules.	No effect because number of molecules remain the same.

Effect of change of temperature on equilibrium.

A change in temperature alters the position of a chemical equilibrium.

A decrease in temperature favours the reaction which liberates heat (exothermic reaction). Subsequently, an increase in temperature favors a reaction which proceeds with absorption of heat (endothermic reaction).

The effect of change in temperature on equilibrium can be summarized by considering the reaction $A + B \rightleftharpoons C + D \ \Delta H = -ve$

Effect of temperature on equilibrium

Reaction	**Effect of change in temperature on equilibrium**	
	Increase	*Decrease*
Exothermic (ΔH -ve)	Shifts to the left and more of A and B would be formed.	Shifts to the right and more of C and D would be formed.
Endothermic (ΔH +ve)	Shifts to the right and more of C and D would be formed.	Shifts to the left and more of A and B would be formed.

For example, Copper turnings react with concentrated nitric(V) acid to produce brown nitrogen(IV) oxide gas ($NO_2(g)$).

$$Cu(s) + 4HNO_3(l) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

When the flask is cooled, the colour of the gas inside changes from brown to yellow. This is because on cooling, the brown nitrogen(IV) oxide (NO_2) dimerises (i.e., two molecules of the gas combine to form one molecule) to form dinitrogen tetra oxide (N_2O_4). At room temperature and pressure an equilibrium between the two gases is established.

$$\underset{\text{Pale yellow}}{N_2O_4\,(g)} \underset{cool}{\overset{heat}{\rightleftharpoons}} \underset{\text{Dar-brown}}{2NO_2(g)}$$

When the mixture is heated, the dinitrogen tetraoxide molecules break up to form nitrogen (IV) oxide molecules. The mixture turns a darker brown colour. The forward reaction in the equation is endothermic since rising the temperature will favour the process that absorbs heat. In this case, the equilibrium shifts from left to right and that is why the mixture becomes dark-brown.

A decrease in temperature favours the reaction which liberates heat. Since backward reaction is exothermic, the colour of the mixture becomes pale-yellow.

Catalysts and Equilibrium.

A catalyst has no effect on the position of equilibrium, it only alters the rate at which the equilibrium is attained.

Some Industrial Applications of Chemical Equilibrium

The conditions required to obtain the highest yield of the products in the shortest time at minimum cost are referred to as optimum conditions.

The position of equilibrium is normally changed by removal of one of the products or through variation of pressure and/or temperature.

The Haber Process

The reaction between nitrogen and hydrogen in the manufacture of ammonia is a reversible reactions. It provides a good example of the application of the equilibrium principle.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)\ \Delta H = -92\ kJ\ mol^{-1}$$

The forward reaction is exothermic and therefore an increase in temperature favours a shift in the equilibrium from right to left, i.e., ammonia breaks down to form nitrogen and hydrogen. This implies that the yield of ammonia is less at high temperatures. Lowering the temperature will favour production of more ammonia. However, the rate at which ammonia is produced will be too slow and therefore uneconomical. An optimum temperature of about 450°C is normally used.

In addition, the forward reaction is accompanied by a decrease in the number of molecules, i.e., volume. Increase in pressure therefore, favours the process of reaction that will tend to lower the pressure. Therefore high pressure favours the production of more ammonia. High pressure also increases the rate of reaction. However the cost of producing and maintaining high pressure in a system is quite enormous. Therefore an optimum pressure of 200 atmospheres is normally applied.

If some of the ammonia is removed from the system at equilibrium, more of the nitrogen and hydrogen will react to produce ammonia to replace that which is removed. Normally ammonia is removed as soon as it is produced so that more of the nitrogen and hydrogen can continue reacting.

The catalyst used for this process is iron. Though platinum is a better catalyst, it is very expensive and easily poisoned by impurities.

The Contact Process

In the manufacture of sulphuric(VI) acid, the main reaction taking place between sulphur (IV) oxide and oxygen to produce sulphur(VI) oxide is also a reversible reaction:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \; \Delta H = -197 \text{ kJ mol}^{-1}$$

The forward reaction is exothermic therefore an increase in temperature favours the backward reaction. Therefore, at high temperature less sulphur(VI) oxide will be formed. Low temperatures will favour the forward reaction but the rate will be slow. For maximum yield, an optimum temperature of 450°C is used.

The forward reaction is accompanied by a decrease in the number of molecules from 3 to 2. Therefore an increase in pressure will favour production of more sulphur(IV) oxide. The optimum pressure used is atmospheric pressure. This gives a percentage conversion of sulphur(IV) oxide to sulphur(VI) oxide of about 96%.

Vanadium(V) oxide is used as a catalyst to increase the rate of the reaction.

- 2006 Q 9 P1
  At 20 °C, NO_2 and N_2O_4 gases exist in equilibrium as shown in the equation below

  $$\underset{\text{Brown}}{2\,NO_2(g)} \rightleftharpoons \underset{\text{pale yellow}}{N_2O_4\,(g)}; \quad \Delta H = -ve$$

  State and explain the observation that would be made when;
    A syringe containing the mixture at 20 °C is immersed in ice-cold water; (1 mark)
    The volume of the gaseous mixture in a syringe is reduced. (1 mark)
- 2006 Q 7c-g P2
  The table below shows the volumes of nitrogen dioxide gas produced when different volume of 1M nitric acid were each reacted with 2.07 g of lead at room temperature.

Volume of 1 M nitric acid (cm^3)	Volume of nitrogen dioxide gas (cm^3)
5	60
15	180
25	300
35	420
45	480
55	480

III. Give a reason why nitric acid is not used to prepare hydrogen gas. (1 mark)

IV. Explain how the rate of the reaction between lead and nitric acid would be affected if the temperature of the reaction mixture was raised. (2 marks)

V. On the grid provided below, plot a graph of the volume of the gas produced (Vertical axis) against volume of acid. (3 marks)

VI. Using the graph, determine the volume of:

   25. Nitrogen dioxide produced when 30cm^3 of 1 M nitric acid were reacted with 2.07 g of lead (1 mark)

   26. 1M nitric acid which would react completely with 2.07g of lead. (1 mark)

VII. Using the answer in d(i) above, determine:

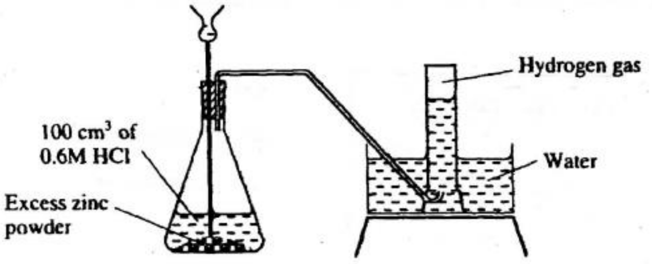
   2. The volume of 1M nitric acid that would react completely with one mole of lead (Pb=207) (2 marks)
   3. The volume of nitrogen dioxide gas produced when one mole of lead reacts with excess 1 M nitric room temperature. (1 mark)

VIII. Calculate the number of moles of:

   - 1M nitric acid that reacted with one mole of lead (1 mark)
   - Nitrogen dioxide produced when one mole of lead were reacted with excess nitric acid. (Molar gas volume of 2400cm^3) (1 mark)

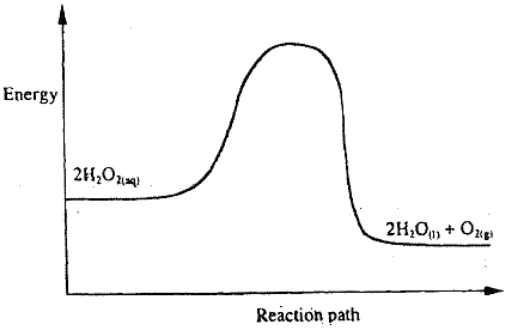
IX. Using the answers obtained in (f) (i) and (ii) above, write the equation for the reaction between lead and nitric acid given that one mole of lead nitrate and two moles of water were also produced. (1 mark)

- 2007 Q 25 P1
  The diagram below shows a student's set-up for the preparation and collection of hydrogen gas.



How would the final volume of hydrogen gas produced be affected if 80cm^3 of 0.75 M hydrochloric acid was used? (1 mark)
Give a reason why helium is increasingly being preferred to hydrogen in weather balloons. (1 mark)

- 2007 Q 27
  The diagram below is a sketch of the graph of the non-catalysed decomposition of hydrogen peroxide. (1 mark)



On the same axis, sketch the graph for the decomposition of hydrogen peroxide when manganese (IV) oxide is added. (2 marks)

- 2007 Q 4 P2, 2016 P2 Q3.

(c) Methanol is manufactured from carbon (IV) oxide and hydrogen gas according to the equation:

$$CO_2 (g) + 3H_2 (g) \rightleftharpoons CH_3OH (g) + H_2O (g)$$

The reaction is carried out in the presence of a chromium catalyst at 700K and 30kPa. Under these conditions, equilibrium is reacted when 2% of the carbon (IV) oxide is converted to methanol.

(i) How does the rate of the forward reaction compare with that of the reverse reaction when 2% of the carbon (IV) oxide is converted to methanol? (1 mark)

(ii) Explain how each of the following would affect the yield of methanol:
   I. Reduction (2 marks)
   II. Using a more efficient catalyst (2 marks)

(iii) If the reaction is carried out at 500K and 30kPa, the percentage of carbon (IV) oxide converted to methanol is higher than 2%
   I. What is the sign of ΔH for the reaction? Give a reason (2 marks)
   II. Explain why in practice the reaction is carried out at 700K but NOT at 500K (1mark)

(d) Hydrogen peroxide decomposes according to the following equation:

$$2H_2O_2 (aq) \longrightarrow 2H_2O (l) + O_2 (g)$$

In an experiment, the rate of decomposition of hydrogen peroxide was found to be 6.0×10^{-8} mol $dm^{-3}s^{-1}$.

(iii) Calculate the number of moles per dm^3 of hydrogen peroxide that had decomposed within the first 2 minutes. (2 marks)

(iv) In another experiment, the rate of decomposition was found to be 1.8×10^{-7} mol $dm^{-3}s^{-1}$. The difference in two rates could have been caused by addition of a catalyst. State, giving reasons, one other factor that may have caused the difference in two rates of decomposition. (2 marks)

- 2008 Q 23

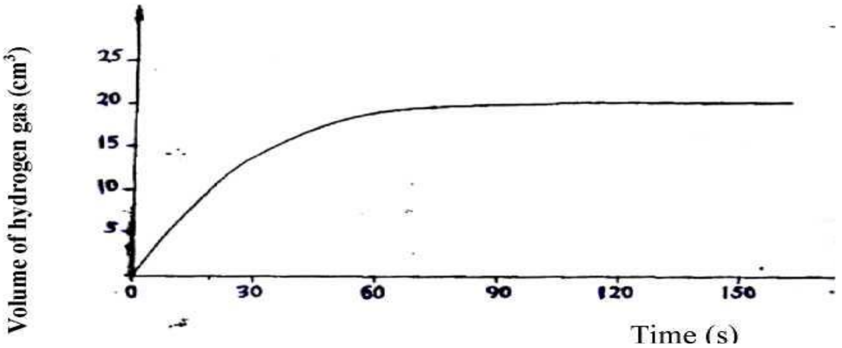
In a closed system, aqueous iron (III) chloride reacts with sulphide gas as shown in the equation below.

$$2FeCl_3 (aq) + H_2S (g) \rightleftharpoons 2FeCl_2 (aq) + 2HCl (aq) + S(s)$$

State and explain the observation that would be made if dilute hydrochloric acid is added to the system at equilibrium. (2 marks)

- 2008 Q 29

A certain mass of a metal E_1 reacted with excess dilute hydrochloric acid at 25 °C. The volume of hydrogen gas liberated was measured after every 30 seconds. The results were presented as shown in the graph below.



i. Name one piece of apparatus that may have been used to measure the volume of gas liberated. (1 mark)

ii. (i)On the same axis, sketch the curve that would be obtained if the experiment was repeated at 35 °C. (1 mark)

(ii) Explain the shape of your curve in b(i) above. (1 mark)

- 2008 Q4 P2

(a) (i) State the Le chatelier's principle. (1mark)

(ii) Carbon (II) oxide gas reacts with steam according to the equation;

$$CO(g) + H_2O(g) \rightleftharpoons H_2(g) + CO_2(g)$$

What would be the effect of increasing the pressure of the system at equilibrium? Explain. (2marks)

(b) The table below gives the volumes of oxygen gas produced at different times when hydrogen peroxide decomposed in the presence of a catalyst.

Time (Sec)	0	10	20	30	40	50	60
Volume of oxygen (cm^3)	0	66	98	110	119	120	120

(vi) Name the catalyst used for this reaction (1mark)

(vii)On the grid provided, draw the graph of volume of oxygen gas produced (vertical axis) against time. (3marks)

(viii) Using the graph, determine the rate of decomposition of hydrogen peroxide after 24 seconds. (2marks)

(ix) Give a reason why the total volume of oxygen gas produced after 50 seconds remains constant. (1mark)

- 2009 Q 27

  The following reaction is in equilibrium in a closed container.

  $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$

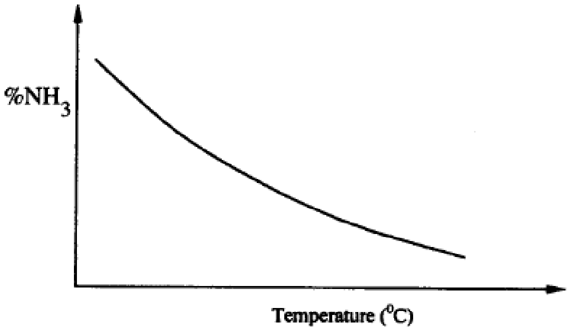
  State giving reasons how an increase in pressure would affect the amount of hydrogen. (2 marks)

- 2010 Q 22 P1

  Nitrogen and hydrogen react to form ammonia gas as shown in the following equation:

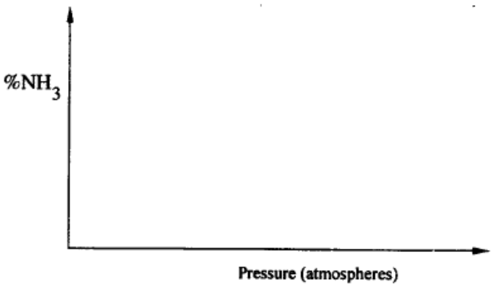
  $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$; ΔH is negative

  - The figure below shows how the percentage of ammonia gas in the equilibrium mixture change with temperature.



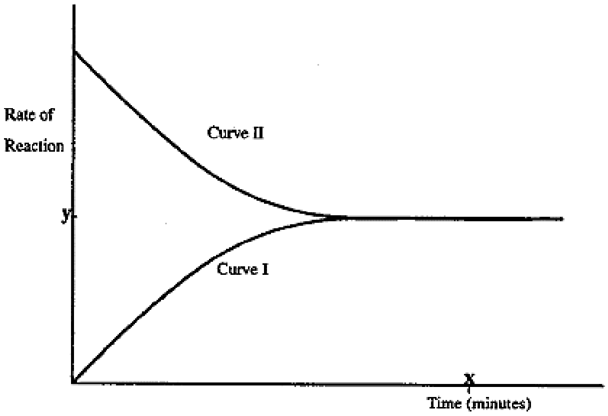
Explain why the percentage of ammonia gas change as shown in the figure.(2marks)

- On the axes below, sketch a graph showing how the percentage of ammonia gas in equilibrium mixture changes with pressure. (1 mark)



- 2010 Q 7 P2
  - The figure below shows how the rate of the following reaction varies with the time.

    $A(g) + B(g) \rightleftharpoons 2C(g) + D(g)$



(i) Which of the two curves represent the rate of the reverse reaction? Give a reason (2 marks)

(ii) What is the significance of point **X** and **Y** on the figure? (2 marks)

- State and explain the effect of an increase in pressure on the rates of the following reactions.

  (2 marks)

  (c) $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

  (d) $CH_3OH(l) + CH_3COOH(l) \longrightarrow CH_3COOCH_3(l) + H_2O(l)$

- In an experiment to study the rate of reaction between barium carbonate and dilute hydrochloric acid; 1.97g of barium carbonate were reacted with excess 2M hydrochloric acid. The equation for the reaction is

$$BaCO_3(s) + 2HCl(aq) \longrightarrow BaCl_2(aq) + CO_2(g) + H_2O(l)$$

The data in the table was obtained

Time in seconds	0	30	60	90	120	150	180	210	240
Volume of gas (cm^3)	0	80	135	175	210	230	240	240	240

11. On the grid provided, plot a graph of volume of gas produced (vertical axis) against time (3 marks)
12. From the graph, determine the rate of the reaction at:
    (I) 15 seconds (1 mark)
    (II) 120 seconds (1 mark)
    (III) Give a reason for the difference between the two values. (1 mark)

- 2011 Q 3

(v) Ethanol can be manufactured from ethane and steam as steam as shown in the equation below:

$$C_2H_4\,(g) + H_2O\,(g) \rightleftharpoons CH_3CH_2OH\,(g)$$

Temperature and pressure will affect the position of equilibrium of the above reaction. Name the other factor that will affect the position of equilibrium of the above reaction. (1 mark)

(vi) The data in the table was recorded when one mole of ethane was reacted with excess steam. The amount of ethanol in the equilibrium mixture was recorded under different conditions of temperature and pressure. Use the data to answer the questions that follow.

Temperature (°C)	**Pressure (Atm)**	**Amount of ethanol at equilibrium (moles)**
300	50	0.40
300	60	0.46
300	70	0.55
250	50	0.42
350	50	0.38

(c) State whether the reaction between ethane and steam is exothermic or endothermic. Explain your answer. (3 marks)

(d) State and explain one advantage and one disadvantage of using extremely high pressure in this reaction.

Advantage (2 marks)

Disadvantage (2 marks)

(vii) In an experiment to determine the rate of reaction between calcium carbonate and dilute hydrochloric acid, 2g of calcium carbonate were reacted with excess 2M hydrochloric acid.The volume of carbon (IV) oxide evolved was recorded at regular intervals of one minute for six minutes. The results are as shown in the table below.

Time (minutes)	1	2	3	4	5
Volume of Carbon (IV)oxide (cm^3)	170	296	405	465	480

(b) Plot a graph of time in minutes on the horizontal axis against volume of carbon (IV) oxide on the vertical axis. (3 marks)

(c) Determine the rate of reaction at 4 minutes. (2 marks)

- 2012 Q13 P1

A dynamic equilibrium is established when hydrogen and carbon (IV) oxide react as shown below:

$$H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)$$

What is the effect of adding powdered iron catalyst on the position of the equilibrium? Give a reason. (2 marks)

- 2012 Q4 P2

The factors which affect the rate of reaction between lead carbonate and dilute nitric (V) acid were investigated by carrying out three experiments:

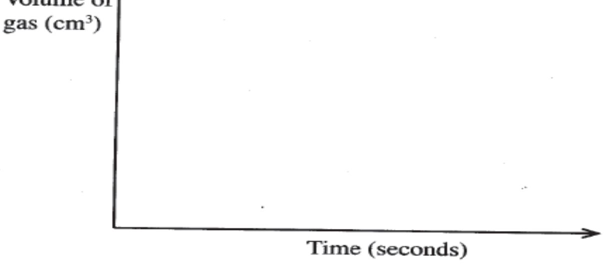
Experiment	**Lead carbonate**	**Concentration of nitric (V) acid**
1	Lumps	4M
2	Powdered	4M
3	Lumps	2M

(c) Other than concentration, name the factor that was investigated in the experiments. (1 mark)

(d) For each experiment, the same volume of acid (excess) and mass of lead carbonate were used and the volume of gas liberated measured with time.

(c) Draw a set up that can be used to investigate the rate of reaction for one of the experiments. (3 marks)

(d) On the grid provided, sketch the curves obtained when the volume of gas produced was plotted against time for each of the experiments and label each as 1, 2, or 3. (4 marks)



(e) Write an equation for the reaction that took place. (1 mark)

(e) If the experiments were carried out using dilute hydrochloric acid in place of dilute nitric (V) acid, the reaction would start, slow down and eventually stop. Explain these observations. (2 marks)

(f) A solution of bromine gas in water is an example of a chemical reaction in a state of balance. The reaction involved is represented by the equation below.

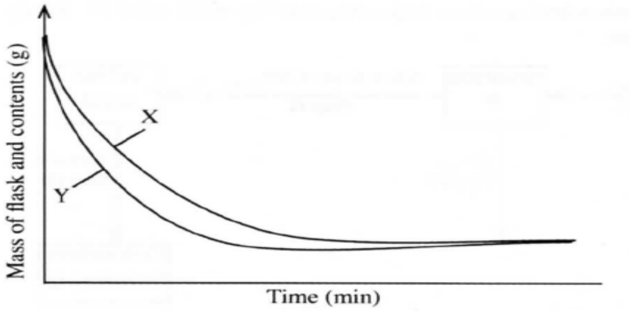
$Br_2 (g) + H_2O (l) \rightleftharpoons 2 H^+(aq) + Br^-(aq) + OBr^-(aq)$

Yellow/orange *Colourless*

State and explain the observations made when hydrochloric acid is added to the mixture.

- 2013 Q16 P1

6. The curves below represent the change in mass when equal masses of powered zinc and zinc granules were reacted with excess 2M hydrochloric acid. Study them and answer the question below.



Which curve represents the reaction with zinc granules? Explain your answer.(2 marks)

- 2014 Q 17 P1

In an experiment on rates of reaction, potassium carbonate was reacted with dilute sulphuric (VI) acid.

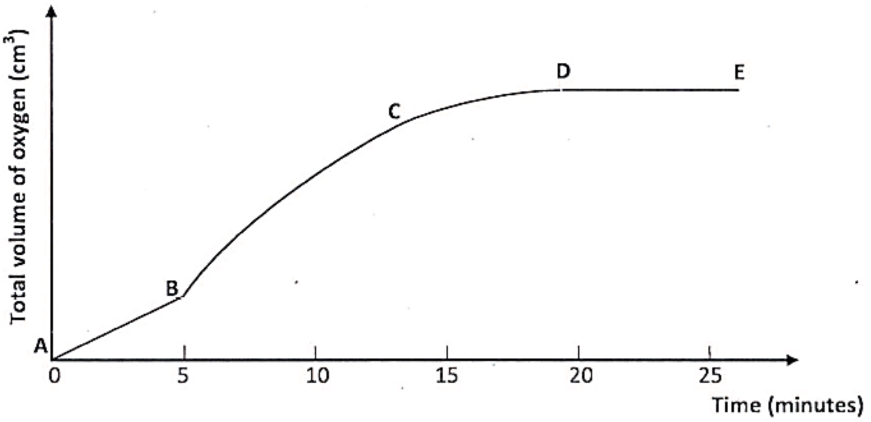
(a) What would be the effect of an increase in the concentration of the acid on the rate of the reaction? (1 mark)

(b) Explain why the rate of reaction is found to increase with temperature. (2 marks)

- 2014 Q 4 P2

III. Other than temperature, state two factors that determine the rate of a chemical reaction. (2 marks)

IV. A solution of hydrogen peroxide was allowed to decompose and the oxygen gas given off collected. After 5 minutes, substance G was added to the solution of hydrogen peroxide. The total volume of oxygen evolved was plotted against time as shown in the graph below



(i) Describe the procedure of determining the rate of the reaction at minute 12.(3 marks)

(ii) How does the production of oxygen in region **AB** compare with that in region **BC**? Explain (2 marks)

(iii) Write an equation to show the decomposition of hydrogen peroxide. (1 mark)

V. Sulphur (IV) oxide react with oxygen to form sulphur (VI) oxide as shown in the equation below

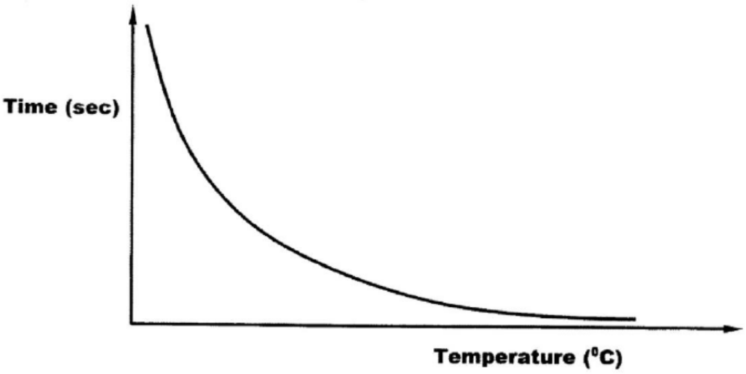
$2SO_2(g) + O_2(g) \quad 2SO_3 (g): \Delta H= -192kJ$

(III) Explain the effect on the yield of SO_3 of lowering the temperature of this reaction. (2 marks)

(IV) Name one catalyst used for the reaction. (1 mark)

- 2015 Q12 P1

  The curve shown below shows the variation of time against temperature for the reaction between sodium thiosulphate and hydrochloric acid.



(a) Write the equation for the reaction between sodium thiosulphate and dilute hydrochloric acid. (1 mark)

(b) Explain the shape of the curve (2 marks)

- 2015 Q6 P2

(iv) Other than concentration, state two factors that determine the rate of a reaction.(2 marks)

(v) In an experiment to determine the rate of reaction, excess lambs of calcium carbonate were added to 2 M hydrochloric acid. The mass of calcium carbonate left was recorded after every 30 seconds. The results are shown in the table below.

Time (seconds)	0	30	60	90	120	150	180	210
Mass of calcium carbonate left (g)	2.00	1.60	1.30	1.00	0.85	0.8	0.8	0.8

(i) Write the equation for the reaction that took place. (1 mark)

(ii) On the grid provided, plot a graph of mass of calcium carbonate vertical axis against time. (3 marks)

(iii) Determine the rate of reaction at the 105th second. (3 marks)

(vi) Why does the curve level off after some time? (1 mark)

(vii) On the same grid, sketch a curve for the same reaction using 4 M hydrochloric acid and label the curve **R**. (2 marks)

- 2016 Q19 P1

  In an experiment on rates of reaction, potassium carbonate was reacted with dilute sulphuric (VI) acid.

7. What would be the effect of an increase in the concentration of the acid on the rate of the reaction? (1 mark)
8. Explain why the rate of reaction is found to increase with temperature. (1mark)

- 2017 P1 Q8.

  a. State one characteristic of a reaction where equilibrium has been attained. (1 mark)
  b. The following equation is in a state of equilibrium: C D Use it to sketch a graphical representation of concentration against time in seconds for the equilibrium. (2 marks)

- 2018 P1 Q 13.

In the Haber process, nitrogen reacts with hydrogen according to the following equation.

$3H_2(g) + N_2(g) \longrightarrow 2NH_3(g)$; $\Delta H = -92\ kJmol^{-1}$

a. What would be the effect of adding a catalyst on the position of the equilibrium? (1 mark)

b. Explain why it is not advisable to use temperatures higher than 773 K in the Haber process. (2 marks)

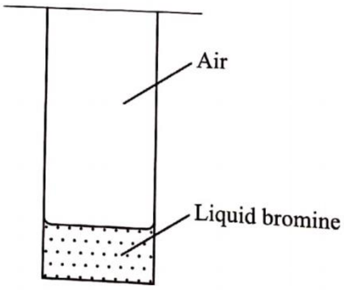
- 2019 P1 Q2.

  Describe how an increase in concentration increases the rate of a reaction. (2 mark)
- 2019 P2 Q7.

a. What is meant by rate of reaction? (1 mark)

b. In the space provided, sketch the diagram of a set-up that can be used to determine the rate of reaction between manganese (IV) oxide and hydrogen peroxide. (3 marks)

c. A student placed a small amount of liquid bromine at the bottom of a sealed gas jar of air as shown in Figure 4.



(c) Describe what will be observed: (1 mark)
I. After two minutes..............................
II. After 30 minutes ....

(d) Use the Kinetic theory to explain the observations: (2 marks)
I. After two minutes..............................
II. After 30 minutes ....

TOPIC FOUR

ELECTROCHEMISTRY

Table of Contents

Organizer

Objectives

By the end of the topic the learner should be able to:

(a) Explain reduction and oxidation in terms of gain and loss of electrons.

(b) Determine changes in oxidation number to keep track of the movement of electrons in redox reactions and write balanced redox equations.

(c) Describe the electrochemical cell and explain its working in terms of electron transfer process.

(d) Draw cell diagrams using correct cell notations.

(e) Use displacement reactions to compare reducing and oxidising power of ions.

(f) Calculate the electromotive force of a cell given standard electrode potentials.

(g) State the role of water during electrolysis and explain the factors affecting preferential discharge of ions.

(h) Relate the quantity of charge passed to the amount of substance liberated at the electrodes and explain some applications of electrolysis.

TOPIC FOUR
ELECTROCHEMISTRY

Electrochemistry is the study of how chemical reactions produce electrical energy and in turn how electrical energy causes chemical reactions.

These chemical reactions involve transfer of electrons.

Redox Reactions

A redox reaction is one in which reduction and oxidation processes occur simultaneously. Redox reactions involve electron gain and electron loss.

Gain of electrons is a reduction process. The species that gains electrons is an **oxidising agent**.

Loss of electrons is an oxidation process. The species that loses electrons is a **reducing agent.**

The reaction between iron fillings and copper(II) sulphate solution can be used to illustrate a redox reaction. When iron filings are added to a solution of copper(II) sulphate, a **brown** solid which is copper metal is formed. The colour of the solution changes from **blue** to **light-green** due to the formation of iron(II) ions in the solution.

$Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$

The reaction between iron and copper(II) ions involves transfer of electrons from the iron atoms to copper(II) ions. The reaction can be used to keep track of the transfer of electrons during the reaction as illustrated by the following ionic half equations

$Fe (s) \longrightarrow Fe^{2+} (aq) + 2e^-$

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

In the first half equation, iron atoms lose electrons (gets oxidized) and form **iron(II) ions**. Iron atoms acts as a reducing agent.

In the second half equation **copper(II) ions** gain electrons (gets reduced) to form copper atoms. The copper (II) ions is an oxidising agent.

Oxidation Numbers

An oxidation number is the **apparent charge** that an element has in a compound or the **charge on an ion.**

Oxidation number is written with the **plus (+) or minus (–) sign in front., ie +2 or – 3.**

The knowledge of oxidation numbers helps one to keep track of electron movements in redox reaction and to understand the naming of inorganic compounds.

Rules of Assigning Oxidation Numbers

1. The oxidation number of an **uncombined element is zero (0)**, e.g., in molecules, **O_2, Cl_2, H_2,** the oxidation number of all the atoms is zero.
2. The charge on an **ion containing one element** is equal to the **oxidation number of that element**. For example,

Ion	Na^+	S^{2-}	Mg^{2+}	N^{3-}	O^{2-}	H^+
Oxidation number	+1	-2	+2	−3	−2	+1

3. Oxidation number of hydrogen in all compounds that contain it is + 1 except in metal hydrides where it is –1.

Compound	HCl	H_2O	HF	NaH	MgH_2
Oxidation number	+1	+1	+1	–1	–1

4. Oxidation number of oxygen in all compounds that contain it is –2, except in peroxides where it is –1 and OF_2 where it is +2.

Compound	H_2O	CuO	H_2O_2	OF_2
Oxidation number	–2	–2	–1	+2

5. In compounds, the sum of the oxidation numbers of all constituent atoms is equal to zero, e.g.,

Compound	H_2O	Na_2O_2	NaCl
Oxidation number	2 (+1) –2=0	2 (+1) + 2(–1) = 0	(+1) + (–1) = 0

6. In ions containing more than one element, the overall charge is equal to the sum of the oxidation numbers of the constituent elements.

NH_4^+	OH^-	$SO^{4}2-$
–3 + 4 (+1) = +1	–2 + (+1) = –1	+ 6 + 4 (–2) = –2

When the oxidation number of an element in a compound or ion is not known, it is calculated from those of others by using rule 5 and 6 respectively.

Some elements have variable oxidation number. A good example is nitrogen.

Species	Oxidation number	Species	Oxidation number
$NO^{3}-$	+ 5	N_2O	+ 1
NO_2	+ 4	N_2	0
NO	+ 2	NH_3, NH_4, Mg_3N_2	– 3

Example 1

What is the oxidation number of sulphur in sulphuric acid, H_2SO_4?

The sum of oxidation numbers in $H_2SO_4 = 0$ *(Rule 5)*

2 (Oxidation number of H) + (Oxidation number of S) + 4 (Oxidation number of O) = 0

2 (+ 1) + oxidation number of S + – 8 = 0

(+ 2) + oxidation number of S + 4 (– 2) = 0

Oxidation number of S = + 6

Therefore the oxidation number of sulphur in Sulphuric(VI) acid is +6

Example 2

Determine the oxidation number of manganese in MnO_4^-

The sum of oxidation number in $MnO_4^- = -1$ *(Rule 6)*

(Oxidation number of Mn) + 4 (Oxidation number of O)=1

Oxidation number of Mn + 4 (– 2) = –1

Oxidation number of Mn + (– 8) = –1

Oxidation number of manganese = + 7

Therefore, the oxidation number of manganese in MnO_4^- *is + 7*

Use of oxidation numbers

1. *Naming compounds.*

Oxidation numbers are used for assigning names to compounds of elements which have more than one oxidation number.

For example, copper has two oxidation numbers + 1 and + 2. Compounds containing copper with an oxidation number of +1 are referred to as copper(I) compounds. Those compounds containing copper with a oxidation number of +2 are called copper(II) compounds.

The oxidation state of an element in a compound is always denoted by a Roman number written in brackets, as shown in the table below.

Substance	Oxidation number of:	IUPAC name
$CuSO_4$	**Cu (+2)**	**Copper(II) sulphate**
CuCl	**Cu (+1)**	**Copper(I) chloride**
FeS	**Fe (+2)**	**Iron(II) sulphide**
H_2SO_3	**S (+4)**	**Sulphuric(IV) acid**
H_2SO_4	**S (+6)**	**Sulphuric(VI) acid**
K_2Cr2O_7	**Cr (+6)**	**Potassium dichromate(VI)**
$KMnO_4$	**Mn (+7)**	**Potassium manganate(VII)**
CO	**C (+2)**	**Carbon(II) oxide**
CO_2	**C (+4)**	**Carbon(IV) oxide**
SO_2	**S (+4)**	**Sulphur(IV) oxide**
SO_3	**S (+6)**	**Sulphur(VI) oxide**

2. *Keeping track of electron movement in redox reactions.*

Knowledge of oxidation number helps in determining whether a reaction is a redox one or not. It also helps in determining which substance has been oxidised or reduced.

Oxidation involves increase in oxidation number while reduction involves decrease in oxidation number.

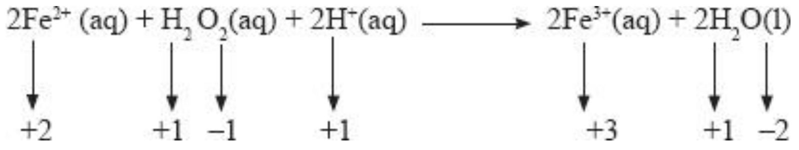
Consider the following reaction between acidified iron(II) sulphate and hydrogen peroxide.

$$2FeSO_4(aq) + H_2O_2(aq) + H_2SO_4(aq) \longrightarrow 2Fe_2(SO_4)_3(aq) + 2H_2O(l)$$

In the reaction above, iron(II) ions are converted to iron(III) ions, thus the oxidation number of iron increases from +2 to +3. Such a change is an **oxidation**.

On the other hand the oxygen from the hydrogen peroxide undergoes a **reduction** by having its oxidation number decrease from –1 to –2 on forming the water molecules.

The above is illustrated using an ionic equation thus,



$$\underset{\substack{\downarrow\\+2}}{2Fe^{2+}}(aq) + \underset{\substack{\downarrow\\+1}}{H_2}\underset{\substack{\downarrow\\-1}}{O_2}(aq) + \underset{\substack{\downarrow\\+1}}{2H^+}(aq) \longrightarrow \underset{\substack{\downarrow\\+3}}{2Fe^{3+}}(aq) + \underset{\substack{\downarrow\\+1}}{2H_2}\underset{\substack{\downarrow\\-2}}{O}(l)$$

Other Examples of Redox Reactions

1. **Reaction of a metal and water**

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

$$0 \quad +1\ -2 \qquad\qquad +1\ -2\ +1 \quad 0$$

Sodium undergoes oxidation because its oxidation number increases from 0 to + 1. Hydrogen undergoes reduction since its oxidation number decreases from + 1 to 0.

2. **Reaction of a metal and an acid**

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

$$\text{Ionic equation } Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2(aq)$$

$$0 \quad +1 \qquad\qquad +2 \quad 0$$

Magnesium undergoes oxidation and its oxidation number increases from 0 to + 2 while the hydrogen ion is reduced. The oxidation number of hydrogen decreases from +1 to 0.

Displacement Reactions

A displacement reaction takes place **when a more reactive element takes the place of another element which is less reactive in a compound.**

Reducing power of metals.

Metals higher in the reactivity series displace from solutions those metals which are lower in the series.

For example, when magnesium is reacted with copper(II) sulphate solution, a **brown** solid which is copper metal is formed. The colour of the solution changes from **blue** to colourless. This is because the blue copper (II) ions in the solution are displaced by magnesium ions which are colourless. The ionic half equations for the formation of magnesium ions and copper metal are:

$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$ (oxidation step)

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$ (reduction step)

When the two ionic half equations are combined, the following overall ionic equation is obtained:

$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$

The oxidation number of magnesium increases from 0 to +2 while that of copper decreases from +2 to 0. Magnesium is **oxidised** and Copper is **reduced**.

Displacement reactions are therefore redox reactions. The **more reactive metal** (Magnesium) is the **reducing agent** while copper(II) ions are the **oxidising agent.**

The more reactive elements such as **sodium** and **calcium** lose their electrons readily and are **strong reducing agents**.

The **less reactive** metals such as lead and copper lose electrons less readily and are **weak reducing agents.**

The order of reducing power is:

Potassium	Strongest reducing agent
Sodium	
Calcium	
Magnesium	
Aluminium	Decreasing reducing power
Zinc	
Iron	
Lead	
Copper	
Silver	Weakest reducing agent

Other Displacement Reactions

1) Aluminium displaces copper from a solution of copper(II) ions.

Oxidation step; $2Al(s) \longrightarrow 2Al^{3+}(aq) + 6e^-$

Reduction step; $3Cu^{3+}(aq) + 6e^- \longrightarrow 3Cu(s)$

$$\boxed{2Al(s) + 3Cu^{2+}(aq) \longrightarrow Al^{3+}(aq) + 3Cu(s)}$$

2) Copper metal displaced from a solution by silver ions.

Oxidation step; $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$

Reduction step; $2Ag^{+}(aq) + 2e^- \longrightarrow 2Ag(s)$

$$\boxed{Cu(s) + 2Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)}$$

Oxidizing Power of Halogens.

Halogens have a tendency to accept electrons and are therefore **strong oxidising agents.**

Among the halogens fluorine is the strongest oxidising agent. However, in most reactions, chlorine is the most common halogen used as an oxidising agent.

The more reactive halogens oxidises the less reactive halogens and vice versa.

For example, Chlorine displaces bromine and iodine from their solutions.

Oxidation step $2Br^-(aq) \longrightarrow Br\ (l) + 2\ e^-$

Reduction step $Cl_2 + 2e^- \longrightarrow 2Cl^-(aq)$

overall reaction $2Br^-(aq) + Cl_2(g) \longrightarrow Br_2(l) + 2Cl^-$

Colourless *Brown*

The oxidation number of bromine increases from –1 to 0 while that of chlorine decrease from 0 to –1. **Bromine is oxidised** while **chlorine is reduced**. **Chlorine** is the **oxidising agent** and the **bromide ion is the reducing agent**.

The ionic equation for the displacement of iodine by chlorine is:

$Cl_2(g) + 2I^-(aq) \longrightarrow 2Cl^-(aq) + I_2(s)$

The oxidation number of **iodine increases from –1 to 0** and that of **chlorine decreases from 0 to –1. Iodine is oxidised** while **chlorine is reduced**.

The above reactions show that chlorine has a greater tendency to accept electrons than both bromine and iodine. Chlorine takes electrons from the bromide and iodide ions forming bromine and iodine respectively.

Similarly, **bromine is more reactive than iodine** and **takes electrons from iodide ions**. Bromine has a **higher tendency** to accept electrons than iodine. The ionic equation for the reaction is:

$Br_2(l) + 2I^-(aq) \longrightarrow 2Br^-(aq) + I_2(s)$

The **oxidation number of bromine decreases from 0 to –1** and that of **iodine increases from –1 to 0.**

The greater the tendency of an element to accept electrons, the higher is its oxidising power.

Among the halogens considered, chlorine is the strongest oxidising agent.

The order of oxidising power is;

Chlorine
Bromine
Iodine

Decreasing oxidising power

The Electrochemical Cell

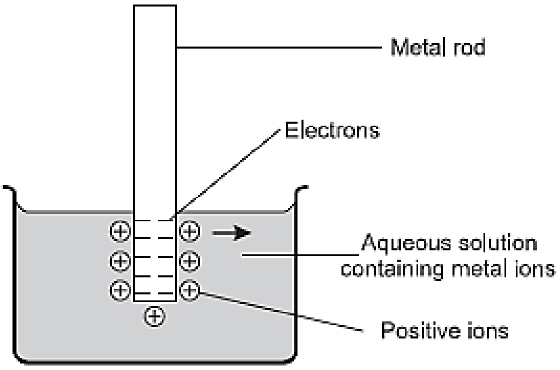
An electrochemical cell is a device that generates a potential difference between electrodes using chemical reactions.

Electrochemical Half-Cell

An **electrochemical half-cell** is formed by dipping a metal rod (electrode) in an aqueous solution of its ions, where some of its surface atoms lose electrons and go into solution as ions. The lost electrons remain on the metal surface.

Metal atoms solution $\longrightarrow$ Metal ions in solution

$M(s) \longrightarrow M^{n+}(aq) + ne^-$



The metal rod develops a negative charge and this attracts the ions back again, some of the ions accept electrons from the rod and form atoms once more.

$M^{n+}(aq) + ne^- \longrightarrow M(s)$

As the negative charge on the surface of the rod builds up, the rate at which the ions combine with electrons increases until eventually it is equal to the rate at which metal atoms lose electrons to form positive ions. At this point an equilibrium is established. The equation for the equilibrium is:

Metal atoms $\rightleftharpoons$ Metal ions

$M(s) \rightleftharpoons M^{n+}(aq) + ne^-$

A potential difference is created between the metal rod and the positively charged ions in the solution.

The concentration of electrons on the metal rod is measured by a quantity called the electrode potential.

The half-cell can be represented as: metal | metal ion. The vertical line represents the phase boundary where a potential difference develops. For example, a zinc half-cell is represented as;

Zn(s) | Zn**2+**(aq)

The tendency of metals to form ions when in contact with their ions differs from one metal to another.

The Electrochemical Cell.

An electrochemical cell is obtained when the half-cells of two different metals are connected to form a complete cell so that the difference between the potential of the half cells can be measured.

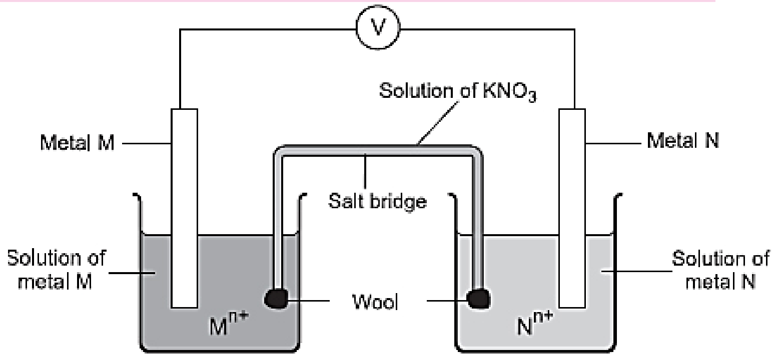
The **electrodes of the two half-cells** are connected by **metallic wires** while the solutions are connected through a **salt-bridge**.

The salt-bridge is in the form of a **filter paper soaked in a saturated solution of potassium nitrate or sodium nitrate.**

The salts chosen for a salt-bridge **must not react with either of the salt solutions in the half cells**.

Electrons flow along the wire **from** the electrode with a **higher concentration** of electrons **to** the electrode with a **lower concentration** of electrons.

The difference between the electrode potentials of the two electrodes is called the electromotive force (e.m.f.) of the cell. The e.m.f. is measured in **volts** using a **voltmeter**.



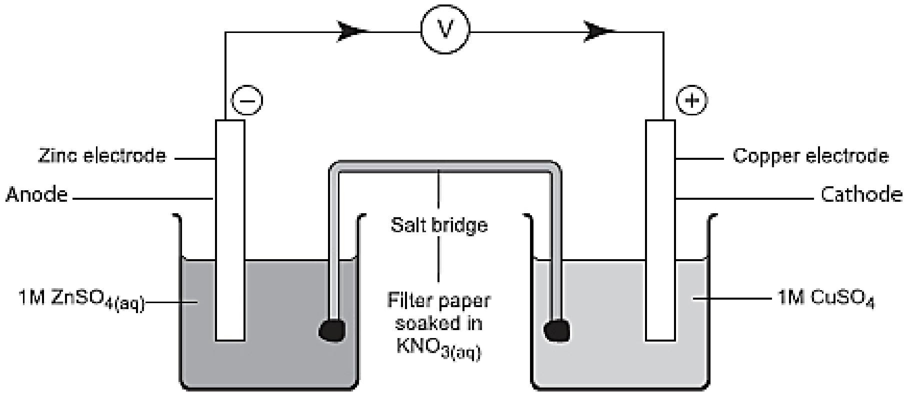
Functions of the Salt Bridge

The functions of the salt-bridge are:

- Complete the circuit by making contact between the two solutions (electrolytes).
- Maintains balance of charges in electrolytes by providing ions to replace those ions that are used up or those that are formed.

Relative tendency of metals to Ionize.

The tendency of metals to form ions when in contact with their ions differs from one metal to another. This property can be used to obtain electrochemical cells, for example the zinc-copper electrochemical cell shown below.



When a **copper-copper** ions half-cell, **Cu(s) | Cu^{2+} (aq)** is connected to a **zinc-zinc ions half-cell, Zn(s) | Zn^{2+}(aq),** the following observations are made:

(i) The zinc rod in the zinc-zinc ions half-cell wears out.
(ii) The intensity of the **blue colour of copper(II) sulphate** solution **decreases** and **red-brown** deposits appear on the copper rod in the copper-copper ions half-cell.
(iii) A voltage of **1.10V** is registered by the **voltmeter**.

In the **Zn(s) | Zn** 2+**|** **|** **Cu**2+**|** **Cu(s)** cell, the following equations represent what happens in the two half cells:

Zinc electrode (anode)

Zn (s) $\longrightarrow$ Zn**2+**(aq) + 2e**–**

Copper electrode (cathode)

Cu**2+**(aq) + 2e**–** $\longrightarrow$ Cu(s)

In the **Zn (s) | Zn**2+**(aq**) half-cell the **oxidation number of zinc increases from 0 to +2.**

In the **Cu(s) | Cu**2+ **(aq)** half-cell the **oxidation number of copper decreases from +2 to 0**.

The anode is defined as the **electrode at which oxidation takes place while the cathode is the electrode at which reduction takes place**.

Oxidation occurs at the **Zn(s) | Zn**2+**(aq) half-cell where electrons are released.**

Reduction takes place at the **Cu(s) | Cu**2+**(aq) where electrons are gained.**

These reactions show that the **zinc electrode** has a **higher tendency to form ions** than the copper electrode when the metals are placed in solutions of their ions.

The **zinc electrode** has a **higher accumulation of electrons and is more negative** compared to the **copper electrode** which has a **lower accumulation of electrons.**

Therefore, the **zinc terminal is relatively more NEGATIVE** with respect to the **copper terminal**.

When the two half cells are connected, electrons flow **FROM the zinc terminal through the connecting wire TO the copper terminal**.

Electrons lost by the zinc electrode are gained by the copper(II) ions.

When the **two ionic half equations are combined,** the ionic equation for the electrochemical cell is obtained.

Zn**2+**(s) + Cu(aq) $\longrightarrow$ Zn**2+**(aq) + Cu(s)

The ionic equation and the e.m.f. of the electrochemical cell can be summed up in what is called **a cell notation**.

Zn(s) | Zn22+**(aq) | | Cu**2+**(aq) | Cu(s)   E = + 1.10 V**

The **single vertical line** represents **phase boundaries** in the half-cells while the **two vertical parallel lines** represent the **salt-bridge**.

The half-cell in which electrons are released **(oxidation takes place)** is always on the **left– hand side** of the cell diagram, i.e., **Zn(s) | Zn**2+**(aq).**

Electrons flow FROM the left hand half-cell TO the right hand half-cell.

Standard Electrode Potentials

The standard electrode potential of any **element | element ions half-cell** is taken as the **difference between its potential and that of hydrogen | hydrogen ions half-cell.**

The chosen standard electrode is **hydrogen** and is usually referred to as the **standard hydrogen half cell**. Being the reference electrode, it is assigned an **electrode potential 0.00 V.**

The standard conditions for measuring electrode potential are:

(i) Temperature of 25°C.

(ii) All solutions have a concentration of 1 mole per litre (1 M).

(iii) Pressure of 1 atmosphere.

Platinised platinum electrode is used as the electrode when the half-cell does not include a metal, e.g., $H_2(g) \mid H^+(aq)$.

The platinised platinum electrode has three functions:

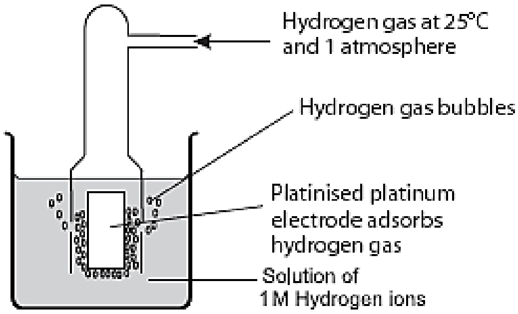
(i) It acts as an inert metal connection to the H | $H^+(aq)$ system.

(ii) It provides a surface area on which dissociation of hydrogen molecules can take place, i.e.,

$$H_2(g) \longrightarrow 2H^+(aq) + 2e^-$$

(iii) It serves as an electrical conductor to the external circuit.

The hydrogen electrode consists of an inert platinum electrode which is immersed in a 1.0 mole per litre solution of hydrogen ions, H^+ Hydrogen gas is bubbled on the platinum electrode which is dipped in a solution containing 1 M hydrogen ions.



$$\tfrac{1}{2} H_2(g) \rightleftharpoons H^+(aq) + e^-$$

The standard hydrogen half-cell is presented as:

$$Pt(s), H_2(g) \mid H^+(aq).$$

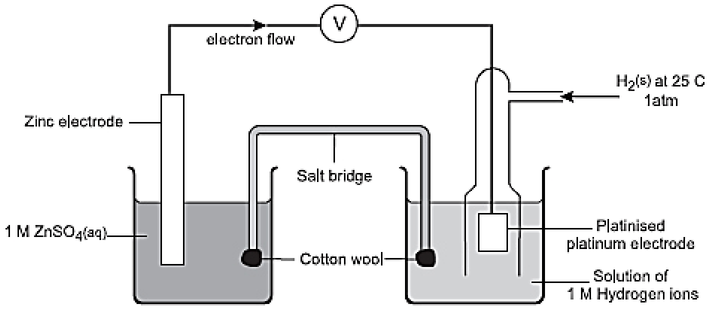
While the standard electrode potential, E for this reference half-cell which is zero can be represented as:

$$\tfrac{1}{2} H_2(g) \longrightarrow H^+(aq) + e^- \quad E = 0.00\text{ V}$$

The **standard electrode potential difference is** the potential difference for a cell comprising a particular element in contact with one molar solution of its ions and the standard hydrogen electrode. It is denoted by the symbol, E^{θ}.

If an element has a greater tendency to lose electrons than hydrogen, the electrode potential of its half-cell is negative with respect to the hydrogen half-cell. (e.g., zinc)

The ionic half equations for the reactions occurring at the electrode are:



$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^- \quad E\theta = 0.76\text{ V}$$

$$2H^+(aq) + 2e^- \longrightarrow H_2(g) \quad E\theta = 0.00\text{ V}$$

The overall ionic equation is:

$$Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g) \quad E\theta = -0.76\text{ V}$$

Similarly, if a F | $F^-(aq)$ half-cell is connected with the Pt(s)H(g) | $H^+(aq)$ half-cell, the e.m.f. registered for the cell is + 2.87 V. The half-cell reaction is as follows:

$$F_2(g) + 2e^- \longrightarrow 2F^-(aq)$$

$$H_2(g) \longrightarrow 2H^+(aq) + 2e^-$$

The overall ionic equation being:

$$H_2(g) + F_2(g) \longrightarrow 2F^-(aq) + 2H^+(aq) \quad E\theta = +2.87\text{ V}$$

On the other hand, if the tendency of an electrode to lose electrons is lower than the hydrogen electrode, the electrode is positive with respect to hydrogen electrode and its potential is positive, e.g., copper.

Standard electrode potentials are sometimes referred to as standard reduction potential because they relate to the reduction reactions.

The table of standard electrode potentials for some elements is arranged so that the strongest oxidising agent, fluorine, which has the most positive value for E° is at the top of the list. The weakest oxidising agent, lithium ions, Li**+**, with the most negative value of E° is at the bottom.

Similarly, fluoride ions is the weakest reducing agent while lithium is the strongest reducing agent.

	Half reaction			
Strongest oxidising agents	$F_2(g) + 2e^- \longrightarrow$	$2F^-(aq)$	+2.87 V	Weakest reducing agents
	$Cl_2(g) + 2e^- \longrightarrow$	$2Cl^-(aq)$	+1.36 V	
	Br_2 aq) + $2e^- \longrightarrow$	$2Br^-(aq)$	+1.09 V	
	$Ag^+(aq) + e^- \longrightarrow$	$Ag(s)$	+0.80 V	
	I_2 (aq) +$2e^- \longrightarrow$	$2I^-(aq)$	+0.54 V	
	$Cu^{2+}(aq) + 2e^- \longrightarrow$	$Cu(s)$	+0.34 V	
	$2H^+(aq) + 2e^- \longrightarrow$	$H_2(g)$	0.00 V	
	$Pb^{2+}(aq) + 2e^- \longrightarrow$	$Pb(s)$	−0.13 V	
	$Fe^{2+}(aq) + 2e^- \longrightarrow$	$Fe(s)$	−0.44 V	
	$Zn^{2+}(aq) + 2e^- \longrightarrow$	$Zn(s)$	−0.76 V	
	$Al^{3+}(aq) + 2e^- \longrightarrow$	$Al(s)$	−1.66 V	
	$Mg^{2+}(aq) +2e \longrightarrow$	$Mg(s)$	−2.37 V	
	$Na^+(aq)+ e^- \longrightarrow$	$Na(s)$	−2.71 V	
Weakest oxidising agents	$Ca^{2+}(aq) + 2e^- \longrightarrow$	$Ca(s)$	−2.87 V	Strongest reducing agents
	$K^+(aq) + e^- \longrightarrow$	$K(s)$	−2.92 V	
	$Li^+(aq) +e^- \longrightarrow$	$Li(s)$	−3.04 V	

Uses of Standard Electrode Potentials

Standard electrode potentials are used in:

(i) Comparing the oxidising and reducing powers of substances.

(ii) Determining the e.m.f. of a cell.

(iii) Predicting whether or not a reaction will take place.

Comparing Oxidising and Reducing Power

Lithium with E**θ** = – 3.04 V has the highest tendency to lose electrons and therefore it is the **strongest reducing agent.** The more negative the E**θ** value the greater the reducing power. Lithium has the least tendency to accept electrons hence it is the weakest oxidising agent.

Fluorine with E**θ** of + 2.87 V has the highest tendency to accept electrons and therefore the **strongest oxidising agent.** The more positive the E**θ** value, the greater the oxidising power. Conversely, fluorine is the weakest reducing agent since it has the least tendency to lose electrons.

Using Standard Electrode Potentials to calculate the e.m.f of a Cell

The e.m.f of a cell is obtained by changing the sign of the electrode potential of the half-cell that undergoes oxidation and then adding to the electrode potential of the half-cell that undergoes reduction.

E**θcell** = E**reduction** – E**oxidation**

E**reduction** is the electrode potential of the half-cell that undergoes reduction.

E**oxidation** is the electrode potential of the half-cell that undergoes oxidation.

Example 1

The reduction potentials of Mg(s) | Mg2+(aq) and Zn(s) | Zn2+(aq) half-cells are:

$$Mg^{2+}(aq) + 2e^- \longrightarrow Mg(s) \quad E^\theta = -2.37V$$

$$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s) \quad E^\theta = -0.76 V$$

Using the electrode potentials, answer the following questions.

(i) Write an ionic equation for a cell made by combining the two half cells.

(ii) Calculate the e.m.f. of the cell formed in part (i).

(iii) Write the cell notation for the cell.

Solution

(i) *Since the electrode potential of magnesium is more negative, magnesium is the stronger reducing agent. Zinc undergoes reduction while magnesium is oxidised. The sign of the* E^θ *value of Mg changes from negative to positive.*

$$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^- \qquad +2.37 V$$

$$Zn^{2+}(aq) + 2e \longrightarrow Zn(s) \qquad -0.76 V$$

$$Mg(s) + Zn^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Zn(s) \qquad +1.61 V$$

(ii) *The e.m.f. of the cell is obtained by changing the sign of the electrode potential of magnesium and adding them, i.e. + 2.37 V + (– 0.76) = + 1.61 V.*

(iii) $Mg(s) \mid Mg^{2+}(aq) \mid\mid Zn^{2+}(aq) \mid Zn(s) \quad E^\theta cell = +1.61V$

Example 2

Calculate the e.m.f. for the electrochemical cell represented below:

$Al(s) \mid Al^{3+}(aq) \mid\mid Zn^{2+}(aq) \mid Zn(s)$

Given that:

$$Al^{3+}(aq) + 3e^- \longrightarrow Al(s) \qquad E° = -1.66 V$$

$$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s) \qquad E° = -0.76 V$$

Solution

$$Al(s) \longrightarrow Al^{3+}(aq) + 3e^- \qquad +0.66 V$$

$$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s) \qquad -0.76 V$$

To combine the two half equations, the number of electrons should be equal. To do this, we multiply the first equation by 2 and the second equation by 3. The electrode potentials remain the same. This is because the voltage does not depend on the number of electrons flowing.

$$2A(s) \longrightarrow 2Al^{3+}(aq) + 6e^- \qquad +1.66 V$$

$$3Zn^{2+}(aq) + 6e^- \longrightarrow 3Zn(s) \qquad -0.76 V$$

$2Al(s) + 3Zn2+(aq) \longrightarrow 2Al3+(aq) + 3Zn\ (s)$ *+ 0.90 V*

Or

$E\theta cell = E\theta R.H.S - E\theta L.H.S$

$= - 0.76\ V - (- 1.66\ V)$

$= - 0.76\ V + 1.66\ V$

$= + 0.90\ V$

Example 3

Use the cell representation below to answer the question that follows

Cu(s) | Cu2+(aq) | | Ag+(aq) | | Ag(s) Eθcell = + 0.46 V

Given that the E value for Ag+(aq) | Ag(s) is + 0.80 V, calculate the Eθ value for Cu(s) | Cu2+(aq).

Solution

$E\theta Cell = E\theta R.H.S - E\theta L.H.S$

Substituting

$0.46\ V = 0.180\ V - E L.H.S$

$0.46\ V - 0.80\ V = -E L.H.S$

$- 0.34 = - E L.H.S$

$E\theta L.H.S = + 0.34\ V$

Example 4

Use the standard electrode potentials for elements A, B, C, D and E given below to answer the questions that follow. The letters do not represent actual symbols of elements.

	Eθ (volts)
$A2+(aq) + 2e- \longrightarrow A(s)$	**–2.37**
$B2+(aq) + 2e- \longrightarrow B(s)$	**–0.76**
$C+(aq) + e- \longrightarrow \frac{1}{2}\ C2(g)$	**0.00**
$D2+(aq) + 2e \longrightarrow D2(s)$	**+0.34**
$\frac{1}{2}\ E2(g) + e- \longrightarrow E-(aq)$	**+ 1.36**

(i) What is the Eθ value of the strongest oxidising agent? Explain.

Answer: ***+ 1.36 most positive.***

(ii) Which two of the above elements would produce the largest e.m.f or potential difference in an electrochemical cell Explain.

Answer: ***A and E, the elements with the most positive Eθ and the most negative Eθ.***

(iii) What would be the initial potential difference of the cell chosen in(ii) above?

Answer:

$E.m.f\ of\ cell = E\ \text{reduction} - R\ \text{oxidation}$

$= 1.36 - (-2.37)$

$= 1.36\ V + 2.37\ V$

$= + 3.73\ V$

(vi) Write the cell representation for the electrochemical cell formed.

Answer: $A(s)\ |\ A2+(aq)\ |\ |\ E2(g)\ |\ E-(aq), Pt\ E\theta = 3.73\ V$

Using standard Electrode potential to predict it a Reduction will take place

Previously it was established that Zn reduces Cu^{2+}

$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e$ + 0.76 V

$Cu^{2+} + 2e^- \longrightarrow Cu(s)$ + 0.34 V

$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$ + 1.10 V

The cell potential or e.m.f., + 1.10 V is positive showing the reaction takes place. The e.m.f. for the reverse reaction:

$Cu(s) + Zn^{2+}(aq) \longrightarrow Cu^{2+}(aq) + Zn(s)$ is $E^\theta = -1.10$ V

The negative value implies that the reaction is unlikely to occur. **In general reactions with an overall positive e.m.f can take place; while those with negative: e.m.f. cannot.**

Example

Predict whether a reaction will occur between iodine and chloride ions.

$I_2(aq) + 2e^- \longrightarrow 2I^-(aq)$ $E\theta = +1.36$ V

$Cl_2(aq)$ $2e^- \longrightarrow 2Cl^-(aq)$ $E\theta = +1.51$ V

Solution

Adding the ionic half equations;

$I_2(aq) + 2e^- \longrightarrow 2I^-(aq)$ $E^\theta = +1.36$ V

$2Cl(aq) \longrightarrow Cl_2(aq) + 2e$ $E^\theta = -1.51$ V

$I_2(aq) + 2Cl^-(aq) \longrightarrow 2I^-(aq) + Cl_2(g)$ $E^\theta = -0.15$ V

The overall E^θ of the cell is negative. Therefore, iodine cannot displace chlorine from a chloride solution.

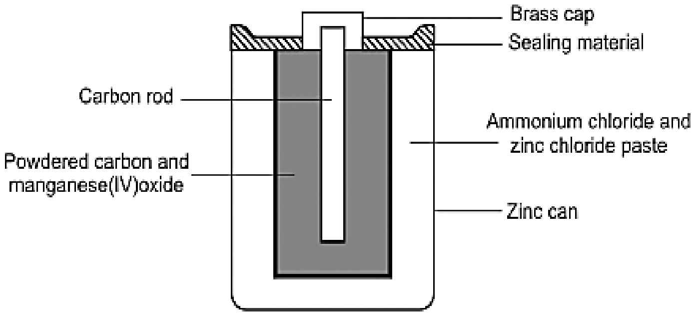
Uses of Electrochemical Cells

Electrochemical cells are used as a source of energy.

Dry Cells

Dry cells are used in a wide range of electrical appliances such as radios, watches, clocks, flashlights and electric bells. The dry cells are cheap and convenient to use because they contain the electrolyte in form of a paste rather than a liquid therefore cannot spill or leak.

An example of a dry cell is the **Le' Clanche cell** shown below.



It consists of a **zinc can which forms the negative terminal** and a **graphite rod** which is the **positive terminal.**

The graphite rod is surrounded by a **paste of ammonium chloride and zinc chloride,** and **powdered manganese(IV) oxide** mixed with **carbon.** The **powder increases the surface area of the positive terminal**.

The **function** of **manganese(IV) oxide** is to **oxidise the hydrogen** produced at the electrode **to water** thus **preventing any bubbles from coating the carbon terminal which would reduce its efficiency.**

At the negative terminal:

$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

At the positive terminal, **ammonium ions** are converted to **ammonia and hydrogen** gases

$2NH_4^+(aq) + 2e^- \longrightarrow 2NH_3(g) + H_2(g)$

The gases produced **do not escape** but are **immediately used up in other reactions**. The **hydrogen is oxidised to water** by the **manganese(IV) oxide. Ammonia** forms a **complex** with the **zinc chloride** in the paste.

A simple dry cell can produce a potential of **1.5 V.**

Dry ammonium chloride does not conduct an electric current hence a **paste**, which is **an electrolyte is used.**

Such a cell is called a **primary cell** because **once the cell is used to supply energy the chemicals are used up and therefore the cell has to be discarded (it cannot be recharged).**

Some cells can be reused after being recharged. These are called secondary cells. The lead acid accumulator is the most common secondary cells.

Accumulators

The main features of a lead-acid accumulator are the **lead plate** which is the **negative terminal** and the **lead (VI) oxide plate** which is the **positive terminal**. Both of these electrodes dip into an aqueous solution of sulphuric(VI) acid.

At the **negative terminal**, lead atoms **lose electrons to form lead(II) ions.**

$Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^-$

At the **positive terminal**, **lead(IV) oxide reacts with hydrogen ions in sulphuric (VI) acid** forming **lead(II) Ions**:

$PbO_2(s) + 4H^+(aq) + 2e^- \longrightarrow Pb^{2+}(aq) + 2H_2O(l)$

The **lead (II) ions** formed **react instantly with the sulphate ions** to form **lead(II) sulphate** which is insoluble and **adheres** to the electrodes.

$Pb^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s)$

The net reaction that takes place is:

$$Pb(s) + PbO(s) + 4H^+(aq) + 2SO_4^{2-}(aq) \xrightarrow{\textit{Discharging}} 2PbSO_4(s) + 2H_2O(l)$$

As the battery **discharges**, **lead and lead(IV) oxide** are **depleted** and the **concentration of sulphuric(VI) acid decreases.**

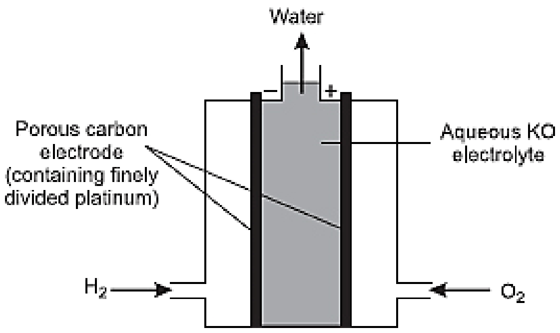
Since the **density of the aqueous solution depends on the concentration of sulphuric(VI) acid, measurement of its density can be used as means of telling how far the battery is discharged.**

During **recharging** of the battery, the electrode reactions shown are **reversed** so as to restore its original reactants:

$$2PbSO_4(s) + 2H_2O(l) \xrightarrow{\textit{Recharging}} Pb(s) + PbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq)$$

Fuel Cells

Fuel cells are **electrochemical** cells which **convert the chemical energy of a fuel directly to electrical energy** such as the hydrogen oxygen cell shown below.



At the **negative terminal**, hydrogen reacts with hydroxide ions to form water and electrons are released.

$$2H2(g) + 4OH^{-}(aq) \longrightarrow 4H2O(l) + 4e^{-}$$

At the **positive terminal**, oxygen and water acquire electrons to form hydroxide ions.

$$O2(g) + H2O(l) + 4\ e^{-}(aq) \longrightarrow 4OH(aq)$$

The overall reaction in the hydrogen | oxygen cell is

$$2H2(g) + O2(g) \longrightarrow 2H2O(l)$$

The cell goes on producing electricity as long as hydrogen and oxygen are fed into it, so that it **does not become exhausted like a primary cell.**

The fuel cell unlike a secondary cell **does not store energy.**

The electrode, **other than completing the circuit** also **catalyses** the reactions which **increase the output of the cell.**

Electrolysis

Electrolysis is the process in which electrical energy is used to cause non-spontaneous chemical reactions to occur.

In these reactions, the substance undergoes chemical decomposition.

Preferential Discharge of Ions During Electrolysis

In the aqueous solution, there are more than two ions since water also ionises. During electrolysis, only one of the anions and one of the cations can be discharged. Preferential discharge therefore takes place, according to the factors discussed below.

Factors Affecting Preferential Discharge During Electrolysis

1. **The concentration of the electrolyte**

   A cation or anion whose concentration is high is preferentially discharged if the ions are close in the electrochemical series.

2. The product obtained at the electrode depends on the **nature of electrode used.**

3. **Position in the Electrochemical Series**

   The ease of reduction of cations and oxidation anions depends on their position in the electrochemical series. The cations high in the series require more energy to be reduced. Anions high in the series require more energy to be oxidised.

Cation	E^θ(Volts)
K^+	–2.92
Ca^{2+}	–2.87
Na^+	– 2.71
Mg^{2+}	–2.38
Al^{3+}	–1.66
Zn^{2+}	–0.76
Fe^{2+}	–0.44
Pb^{2+}	–0.13
H^+	–0.00
Cu^{2+}	+ 0.034
Hg^+	+0.79
Ag^+	+0.80

Least easily discharged

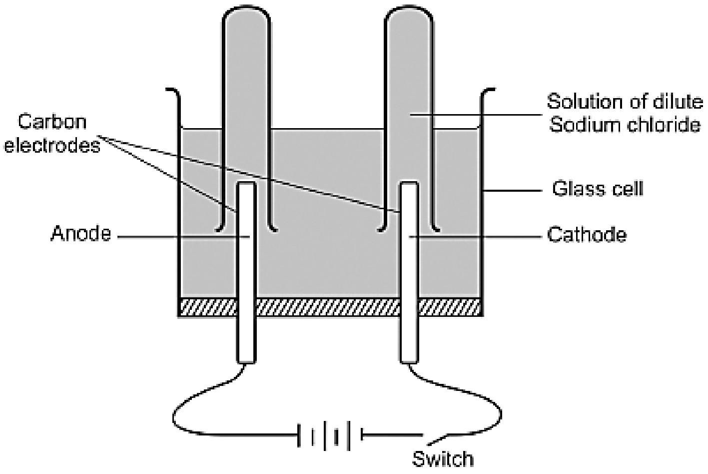
Increased ease of discharge

Most readily discharged

Anion	E^θ(Volts)
F^-	+ 2.87
SO^{2-}_4	+2.01
Cl^-	+1.36
Br	+1.09
NO^-_3	+0.80
I^-	+0.54
OH^-	+0.40

1. Electrolysis of dilute sodium chloride

Sodium chloride solution contains **sodium ions (Na^+)** and **chloride ions (Cl^-)** from **sodium chloride, hydrogen ions (H^+)** and **hydroxide ions (OH^-)** from **water.**



When an electric current is passed through the solution, **chloride ions (Cl^-) and hydroxide ions (OH^-)** migrate to the **anode.**

Hydroxide ions (E**θ**= + 0.04 V) **have a greater tendency to lose electrons compared to the chloride ions.** (E**θ**= +1.36 V).

Reaction at Anode

$4OH^-(aq) \longrightarrow 4OH(aq) + 4e^-$

$4OH(aq) \longrightarrow 2H_2O(l) + O_2(g)$

Overall;

$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$

The anion with a low Eθ value is preferentially discharged.

Sodium ions (Na^+) and hydrogen ions (H^+) migrate to the cathode but **hydrogen ions (Eθ= 0.00** volts) are **preferentially discharged** because **they have a greater tendency to gain electrons than sodium ions (Eθ = –2.71 volts).**

Reactions at Cathode

$4H^+(aq) + 4e \longrightarrow 4H(aq)$

$4H(aq) \longrightarrow 2H_2(g)$

The cations with a higher Eθ has a higher tendency to gain electrons.

Overall;

$$4H^{+}(aq) + 4e^{-} \longrightarrow 2H_{2}(g)$$

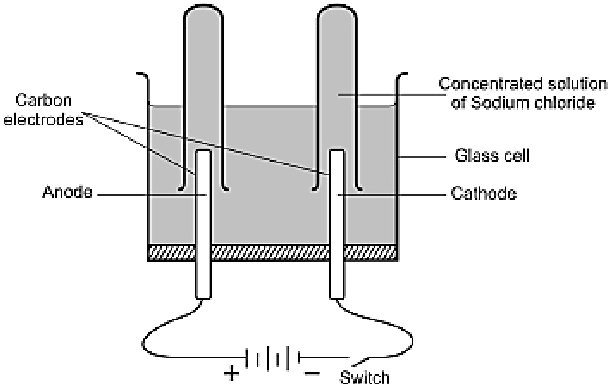
Electrolysis of dilute sodium chloride evolves **oxygen gas at the anode** and **hydrogen gas at the cathode**. This is essentially the electrolysis of water since oxygen and hydrogen are the constituents of water.

2. Electrolysis of concentrated sodium chloride solution (Brine)

Brine is concentrated sodium chloride solution. The solution therefore contains the same ions as dilute sodium chloride.

When an electric current is passed through brine, **chloride ions (Cl^-) and hydroxide ions**

(OH^-) migrate to the anode. The **chloride ions (Eθ = + 1.36 volts) are preferentially discharged because of their relatively high concentrations.**



Reaction at Anode

$$2Cl^{-}(aq) \longrightarrow 2Cl(g) + 2e^{-}$$

$$2Cl(g) \longrightarrow Cl_{2}(g)$$

Overall;

$$2Cl^{-}(aq) \longrightarrow Cl_{2}(g) + 2e^{-}$$

Sodium ions (Na^+) and hydrogen ions (H^+) migrate to the cathode.

At the cathodes, sodium ions **are not preferentially discharged** in spite of their high concentration. This is because the **tendency of hydrogen ions (Eθ = 0.000 volts) to gain electrons is much higher than that of sodium ions(Eθ = –2.71 volts).**

Therefore, **hydrogen ions are preferentially discharged.**

Reactions at Cathode

$$2H^{+}(aq) + 2e^{-} \longrightarrow 2H(g)$$

$$2H(g) \longrightarrow H_{2}(g)$$

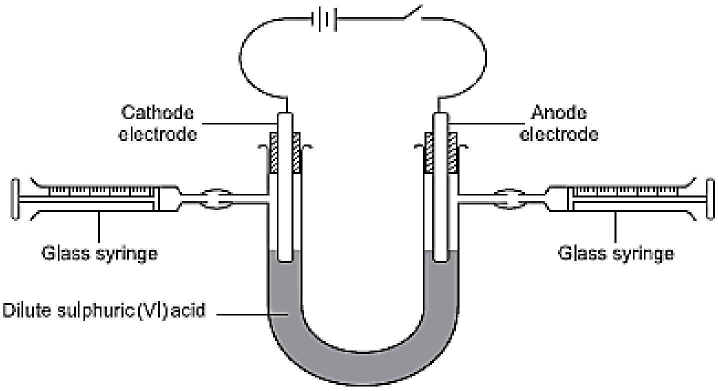
Overall;

$$2H^{+}(aq) + 2e^{-} \longrightarrow H_{2}(g)$$

The discharge of hydrogen ions leads to an **increase in hydroxide (OH^-) ions concentration and the solution becomes alkaline.**

As the electrolysis process **continues**, the **concentration of chloride ions decreases and eventually hydroxide ions are oxidised to water and oxygen gas.**

3. Electrolysis of dilute sulphuric (VI) acid.



Dilute sulphuric(VI) acid contains sulphate ($SO\mathbf{42–}$), hydroxide ($OH–$) and hydrogen ($H\mathbf{+}$) ions.

When an electric current is passed through the dilute acid, sulphate and hydroxide ions migrate to the anode while the hydrogen ions migrate to the cathode.

Reaction at anode

The **hydroxide ions(Eθ = + 0.40 volts)** are **preferentially discharged** because they have a **greater tendency to lose electron**s than sulphate ions (E**θ** = + 2.01 volts).

$4OH–(aq) \longrightarrow 4OH(aq) + e–$

$4OH(aq) \longrightarrow 2H2O(l) + O2(g)$

Overall;

$4OH–(aq) \longrightarrow 2H2O(l) + O2(g) + 4e–$

Reaction at Cathode

$4H+(aq) + 4e– \longrightarrow 4H(g)$

$4H(g) \longrightarrow 2H2(g)$

Overall;

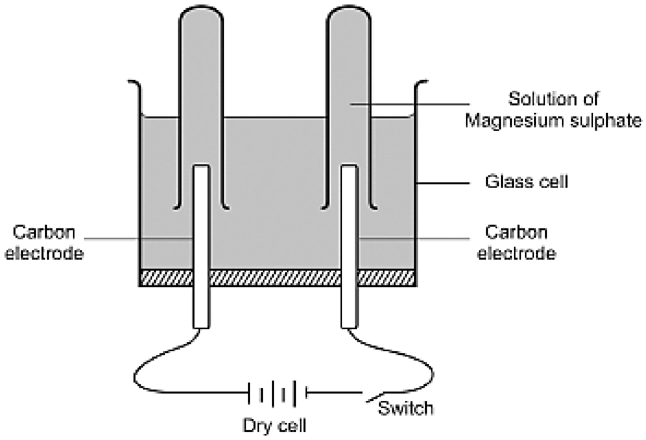
$4H+(aq) + 4e– \longrightarrow 2H2(g)$

The **four electron**s lost by hydroxide ions to form one mole of oxygen molecules are gained by the **four hydrogen ions to form two moles of hydrogen molecules.** For every mole of oxygen gas produced at the anode two moles of hydrogen are formed at the cathode.

The **volume of hydrogen** is therefore **twice** that of **oxygen**.

The amount of **water in the electrolyte decreases** as the electrolysis process continues. This causes an **increase in the concentration of the acid**.

4. Electrolysis of aqueous magnesium sulphate



The ions present in magnesium sulphate solution are magnesium ($Mg\mathbf{2+}$), sulphate ($SO\mathbf{42–}$), hydrogen ($H+$) and hydroxide ($OH\mathbf{–}$).

When an electric current is passed through the solution, hydroxide ions (E**θ** = + 0.40 volts) and sulphate ions (E**θ** = + 2.01 volts) migrate to the anode.

Hydroxide ions are preferentially discharged because of their greater tendency to lose electrons.

Reaction at Anode

$4OH–(aq) \longrightarrow 4OH(aq) +4e–$

$4OH(aq) \longrightarrow 2H2O(l) + O2(g)$

Overall;

$4OH^-(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e$

Both magnesium (E**θ** = –2.38 volts) and hydrogen ions (E**θ** = 0.00 volts) migrate to the cathode.

Hydrogen ions are preferentially discharged because of their greater tendency to gain electrons.

Reactions at Cathode

$4H^+(aq) + 4e^- \longrightarrow 4H(g)$

$4H(g) \longrightarrow 2H_2(g)$

Overall;

$4H^+(aq) + 4e^- \longrightarrow 2H_2(g)$

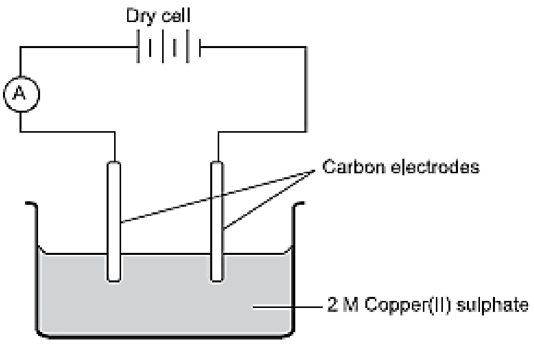
It is observed that the **volume of oxygen gas produced at the anode and hydrogen gas at the cathode are in the ratios of 1:2 respectively.**

5. Electrolysis of copper(II) sulphate solution using different electrodes.

(a) Using inert electrodes (carbon or platinum)

Copper(II) sulphate solution contains copper(II) (Cu^{2+}), Sulphate (SO_4^{2-}), hydrogen ions (H^+), and hydroxide (OH^-) ions.

When the solution is electrolysed using carbon or platinum electrodes, sulphate ions (E**θ** = + 2.01 volts) and hydroxide ions (E**θ**= 0.20 volts) migrate to the anode while the copper(II) (E**θ**= +0.34 volts) and hydrogen ions (E**θ** = 0.00 volts) migrate to the cathode.



Reactions at Anode

The **hydroxide ions have a greater tendency to lose electrons** and therefore are **preferentially discharged**.

$4OH^-(aq) \longrightarrow 4OH(aq) + 4e^-$

$4OH(aq) \longrightarrow 2H_2O(l) + O_2(g)$

Overall;

$4OH(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^-$

Reactions at Cathode

The **copper(II) ions have greater tendency to gain electrons that the hydrogen ions** and is therefore **preferentially discharged.** The atoms are deposited on the cathode as **red-brown** coating.

As a result, the mass of the cathode increases while that of the anode remains the same.

$Cu^{2+}(aq) + 2e \longrightarrow Cu(s)$

The **concentration of copper(II) ions in solution decreases** and the **blue colour** of the **copper(II) sulphate** solution becomes pale and finally colourless.

Hydrogen ions **accumulate in the solutions** and therefore the **solution becomes acidic.**

(b) Using copper electrodes

When copper electrodes are used in the electrolysis of copper(II) sulphate solution**, the mass of the anode decreases,** while that of the **cathode increases**.

Sulphate and hydroxide ions migrate to the anode, but **none of them is discharged;** instead the **copper anode** is gradually oxidised and goes into solution.

Cu(s) $\longrightarrow$ Cu2+(aq) + 2e

This explains the loss in mass of the anode. Less energy is needed for the copper anode to lose electrons than hydroxide ions.

Hydrogen and copper(II) ions migrate to the **cathode** where **copper(II) ions are preferentially discharged** because they have **greater tendency to accept electrons.**

Cu2+(aq) + 2e $\longrightarrow$ Cu(s)

The cathode is thus **coated with a red-brown deposit of copper metal.**

The amount of **copper oxidised** at the anode is **equal** to the **amount of copper deposited on the cathode** and therefore the concentration of **copper(II) ions** in the solution **remains the same. The colour of the blue solution does not fade.**

Carbon or platinum electrodes **allow passage of an electric current into and out of the electrolyte without wearing out such electrodes.**

Summary of Electrolysis

1. **Electrolyte:** dilute sodium chloride

Ions Present	
Cations	**Anions**
Sodium, $Na^+(aq)$ and hydrogen $H^+(aq)$ ions. Hydrogen, $H^+(aq)$ ions discharged at the cathode, because sodium is above hydrogen in electro-chemical series. Hydrogen gas evolved.	Chloride $Cl^-(aq)$ and hydroxide $OH^-(aq)$ ions. Hydroxide $(OH^-)(aq)$ discharged at the anode, because hydrogen ions required less energy to discharge than chloride $Cl^-(aq)$ ions. Oxygen gas evolved.

2. **Electrolyte:** Brine (Concentrated Sodium chloride)

Ions Present	
Cations	**Anions**
Sodium, $Na^+(aq)$ and hydrogen $H^+(aq)$ ions. Hydrogen, $H^+(aq)$ ion discharged at the cathode, because sodium is above hydrogen in electro-chemical series. Hydrogen gas evolved.	Chloride $Cl^-(aq)$ and hydroxide $OH^-(aq)$ ions Hydroxide $(OH^-)(aq)$ discharged at the anode, because hydrogen ions required less energy. To discharge than chloride $Cl^-(aq)$ ions. Chlorine gas evolved.

3. **Electrolyte:** dilute sulphuric acid

Ions Present	
Cations	**Anions**
Hydrogen, $H^+(aq)$	Sulphate, $SO_4^{2-}(aq)$ and hydroxide, $OH^-(aq)$

Hydrogen, H^+(aq) ions discharged at the cathode, because no other cations are present. Hydrogen gas evolved.	Hydroxide (OH^-)(aq) discharged at the anode, because hydroxyl ions require less energy to discharge than sulphate ions. Oxygen gas evolved.

4. **Electrolyte:** Magnesium sulphate

Ions Present	
Cations	**Anions**
Magnesium, Mg^{2+}(aq) and hydrogen, H_+ (aq) ions. Hydrogen ions discharges at the cathode, because magnesium is above hydrogen in the electro-chemical series. Hydrogen gas evolved.	Sulphate, SO_4^{2-} (aq) and Hydroxide OH^-(aq) ions. Hydroxide ions discharged at the anode because hydroxide ions require less energy to discharge than sulphate ions. Oxygen gas evolved.

5. **Electrolyte:** Copper(II) sulphate **Electrode:** Carbon rods

Ions Present	
Cations	**Anions**
Copper, Cu^{2+}(aq) and Hydrogen H^+(aq) ions. Copper ions discharged at the cathode, because copper is below hydrogen in the electro-chemical series. Colour of solution eventually fades.	Sulphate, SO_4^{2-}(aq) and Hydroxide OH–(aq) ions. Hydroxide ions discharged at the anode because hydroxide ions require less energy to discharged than sulphate ions. Oxygen gas evolved.

6. **Electrolyte:** copper(II) sulphate **Electrode:** Carbon

Ions Present	
Cations	**Anions**
Copper, Cu^{2+} (aq) and Hydrogen H^+ (aq) ions. Copper ions discharged at the cathode, because copper is below hydrogen in the electro-chemical series. Colour of solution does not fades.	Sulphate, SO_4^{2-} (aq) and Hydroxide OH^- (aq) ions. No ion discharged at the anode instead the electrode dissolves, and goes into solution as copper (II) ions.

Quantitative Treatment of Electrolysis

An electric current is measured in amperes. The quantity of electric charge(Q) is measured in coulombs(C).

- **A coulomb is the quantity of electricity passed when a current (I) of one ampere flows for a time(t) of one second.**

The relationship between the mass of substance produced and the quantity of electricity passed is the basis of **Faraday's Law of electrolysis** which states that **the mass of a substance produced during electrolysis is directly proportional to the quantity of electricity passed**.

The quantity of electricity carried by one mole of electrons is a constant called a **Faraday (F)** and is equivalent to **96,487 coulombs**.

- **1 Faraday is equivalent to one mole of electrons. The number of electrons required to deposit a given ion is equivalent to the charge on the ion.**

Worked Examples

1. **What mass of copper would be deposited on the cathode when a steady current of one ampere flows for 30 minutes through copper(II) sulphate solution?**
   (Cu = 63.5) Faraday constant = 96,487 C mol^{-1})

   Solution

   Reaction at the cathode

   $Cu^{2+}(aq) + 2e \longrightarrow Cu(s)$

   One mole of Cu ions required 2 moles of electrons.

   Quantity of electricity (Q) = 1 × 30 × 60 coulombs.

   1 mole of electrons carries a charge of 96,487 coulombs.

   2 moles of electrons will carry 2 × 96,487 coulombs.

   2 × 96,487 coulombs deposit 63.5 g of copper at the cathode.

   Therefore 1 × 30 × 60 coulombs deposits:

   $$\frac{63.5 \times 130 \times 60}{2 \times 96{,}487} = 0.59231 \text{ g}$$

2. **What volume of oxygen will be liberated at the anode when a current of 3 amperes is passed through magnesium sulphate solution for 45 minutes and 30 seconds?**
   (Molar gas volume at r.t.p. = 24.0 litres, Faraday constant = 96,500 coulombs).

   Solution

   Reaction equation at the anode

   $4OH^{-}(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e$

   1 mole of electrons carry 96,500 coulombs.

   4 moles of electrons carry 4 × 96,500 coulombs.

   4 × 96,500 coulombs liberate 24 litres of oxygen.

   3 × (45 × 60) +30) coulombs will liberate $\dfrac{24 \times 32{,}730}{4 \times 96{,}500}$

   = 0.50922 litres.

3. In an experiment to electrolyse copper(II) sulphate solution using copper electrodes, 0.2 amperes were passed through the solution for 1,930 seconds. The mass of copper cathode increased from 6.35 to 6.478 g. Find the charge on a copper ion. (1 Faraday = 96,500 coulombs, Cu = 64).

   Solution

   Mass of copper deposited = (6.478 – 6.350) g = 0.128 g

   Quantity of electricity passed = 0.2 × 1,930 = 386 C

   0.128 g copper was deposited by 386 C

   64 g of copper would be deposited by $\dfrac{(386 \times 64)}{0.128}$ *C = 193,000 C*

   1 mole of copper atoms (Mass 64 g) require 193,000 C

   Number of Faradays required to deposit 64 g of copper at cathode is therefore

   $$\frac{193000}{96500} = 2 \text{ F}$$

Applications of Electrolysis

Extraction of Reactive Elements

Electrolysis is used in the extraction of reactive elements such as **sodium, magnesium, aluminum and chlorine.**

Electroplating

This is the process of using electricity to coat one metal with another. This is done to protect some metals from corrosion. Electroplating is also done to make an article look attractive. Gold plated watches, silver utensils are common items.

Sacrificial metal (cathodic protection)

Iron or steel structures are protected from corrosion through sacrificial protection either by galvanising or cathodic protection.

Cathodic protection

Corrosion involves loss of electrons by an element to form ions. If it is a less reactive metal it is connected to a more reactive metal by a conductor when the conditions for causing corrosion are present, the more reactive metal ionises at the expenses of the less reactive. The more reactive metal is sacrificed and the method is sacrificial protection.

Galvanising

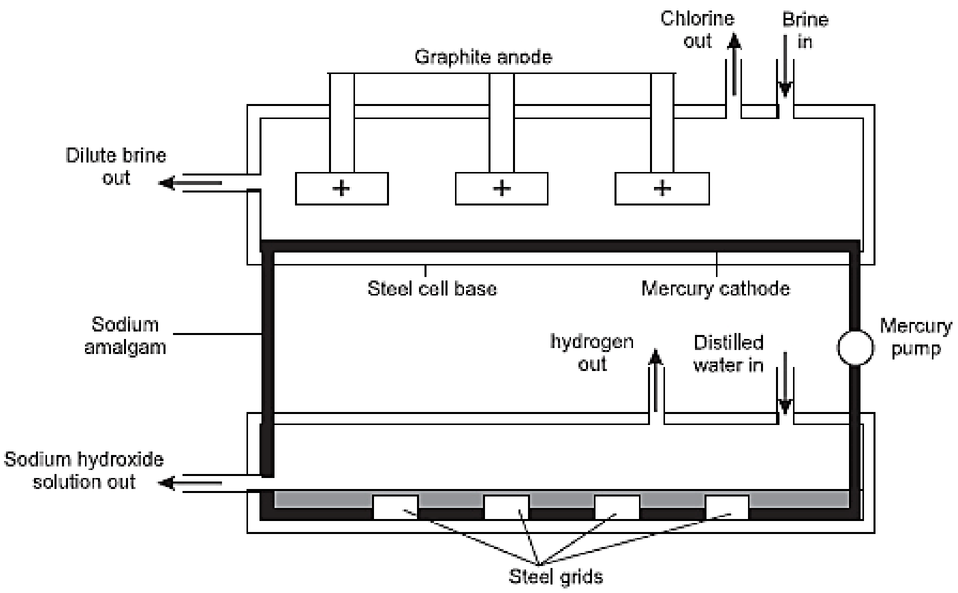
When the galvanised surface is scratched and iron is exposed, zinc passes into solution as zinc ions rather than Fe^{2+} ions. This is possible because zinc is easily oxidised than iron.

Purification of Metals

Copper and other less reactive metals are purified by electrolysis. The impure metal is made the anode and pure metal the cathode. The electrolyte contains the cation of the metals being purified.

Manufacture of Sodium Hydroxide and Chlorine from Electrolysis of Concentrated Sodium Chloride (Brine)

Sodium hydroxide and chlorine are manufactured by the electrolysis of brine by use of **mercury cell, shown alongside.**



The electrolyte in the mercury cell is a **concentrated solution of sodium chloride (Brine)**. The **anode** in the cell is made of **carbon or titanium** because **they do not react with chlorine gas.** While the cathode is a moving film of mercury.

When an electric current is passed through concentrated sodium chloride solution, chloride (Cl^-) ions and hydroxide (OH^-) ions migrate to the **anode**. **Chloride ions are preferentially discharged because of their relatively high concentration.**

$$2Cl^-(aq) \longrightarrow 2Cl(g) + 2e^-$$

$$2Cl(g) \longrightarrow Cl_2(g)$$

Both sodium (Na**+**) ions and hydrogen (H**+**) ions migrate to the **cathode (moving film of mercury). Hydrogen ions are not discharged because of the over-potential (excess power) required to discharge it. Sodium ions are preferentially discharged instead.**

$$2Na^+(aq) + 2e^- \longrightarrow 2Na(l)$$

The sodium atoms formed dissolve in the hot mercury to form **sodium amalgam (Na Hg)**

$$Na(l) + Hg(l) \longrightarrow NaHg(l)$$

The sodium amalgam is passed through a trough in the cell that contains **distilled water.** The sodium in the amalgam then **reacts with water** to form a solution of **sodium hydroxide** and **hydrogen gas.** Hydrogen is pumped out to the required place, while **mercury is regenerated and recycled.**

$$2NaHg(l) + 2H_2O(l) \longrightarrow 2NaOH(aq) + 2Hg(l) + H_2(g)$$

The sodium hydroxide obtained this way is about **fifty per cent pure.** Pure sodium hydroxide is obtained by **evaporating the water in the aqueous sodium hydroxide solution to get pellets or flakes.**

The process is **expensive** due to the high cost of mercury and the safety measures applied since mercury is poisonous.

Review Exercises

1. 2006 Q 15 P1

Study the standard reduction potential given and answer the questions that follow. (The letters are not the actual symbols of the elements).

	E^θ (volts)
$M^{2+}(aq) + 2e^- \longrightarrow M(s)$	-0.76
$N^{2+}(aq) + 2e^- \longrightarrow N(s)$	-2.37
$P^+(aq) + e^- \longrightarrow P(s)$	+0.80
$Q^{2+}(aq) + 2e^- \longrightarrow Q(s)$:	-0.14

(a) The standard reduction potential for $Fe^{2+}(aq)$ is -0.44 volts. Select the element which would best protect iron from rusting. (1 mark)

(b) Calculate the E^θ value for the cell represented as $M(s)/M^{2+}(aq)//P^+(aq)/P(s)$.

(2 marks)

2. 2006 Q 1 P2

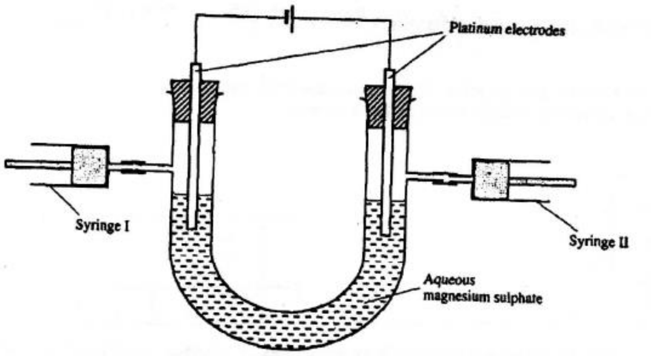
(a) What is an electrolyte? (1 mark)

(b) State how the following substances conduct electricity.

(i) Molten calcium chloride. (1 mark)

(ii) Graphite. (1 mark)

(c) The diagram below shows a set up that was used to electrolyse aqueous magnesium sulphate.



(i) On the diagram above, using an arrow, show the direction of flow of electrons. (1 mark)

(ii) Identify the syringe in which hydrogen gas would be collected. Explain (1 mark)

(d) Explain why the concentration of magnesium sulphate was found to have increased at the end of the experiment. (2 marks)

(e) During the electrolysis, a current of 0.72 A was passed through the electrolyte for 15 minutes. Calculate the volume of gas produced at the anode. (1 Faraday = 96 500 coulombs; molar gas volume is 24000 cm^3 at room temperature). (4 marks)

3. 2006 Q 2b P2

Use the reduction potentials given below to explain why a solution containing copper ions should not be stored in a container made of zinc.

$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$; $E^\theta = -0.76$ V

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$; $E^\theta = +0.34$ V

(2 marks)

4. 2007 Q 7 P1

(a) Use the information given below to draw a labelled diagram of an electrochemical cell that can be constructed to measure the electromotive force between G and J. (2 marks)

$G^{2+}(aq) + 2e^- \longrightarrow G(s)$; $E^\theta = -0.74$ V

$J^{2+}(aq) + 2e^- \longrightarrow J(s)$; $E^\theta = -0.14$ V

(b) Calculate the E^θ value for the cell constructed in (a) above. (1 mark)

5. 2007 Q 21 P1

(a) When brine is electrolyzed using inert electrodes, chlorine gas is liberated at the anode instead of oxygen. Explain this observation. (2 marks)

(b) Name the product formed at the cathode. (1 mark)

6. 2007 Q 28 P1

During the electrolysis of aqueous silver nitrate, a current of 5.0 A was passed through the electrolysis for 3 hours.

(a) Write the equation for reaction which took place at the anode. (1 mark)

(b) Calculate the mass of silver deposited (Ag = 108; 1 F=96500 C) (2 marks)

7. 2008 Q 19 P1

Select a letter which represents a mono atomic gas.

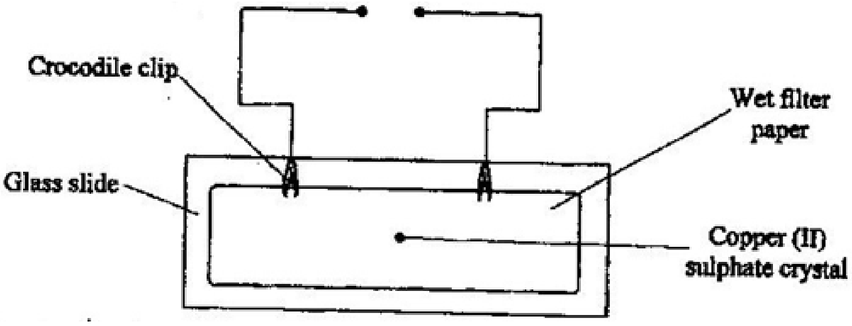
	E^θ (Volts)
$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$	-0.76
$Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$	-0.13
$Ag^+(aq) + 2e^- \longrightarrow Ag(s)$	+0.80
$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	+0.30

(a) Write the cell representation for the electrochemical cell that would give the highest E (1 mark)

(b) State and explain the observations made when a copper rod is placed in a beaker containing silver nitrate solution. (2 marks)

8. 2008 Q 21

The diagram below represents an experiment that was set up to investigate movement of ions during electrolysis.



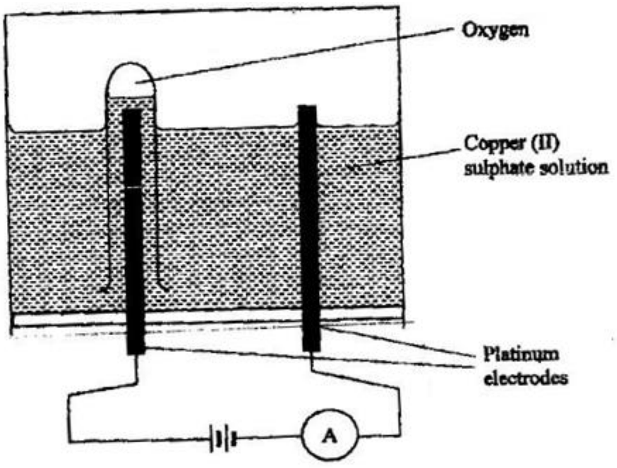
When the circuit was completed, it was noticed that a blue colour spread towards the right.

(a) Explain this observation (2 marks)

(b) Write the equation for the reaction that occurred at the anode. (1 mark)

9. 2008 Q 6 P2

The diagram below represents a set up that can be used to electrolyze aqueous copper (II) sulphate.



(a) (i) Describe how oxygen gas is produced during the electrolysis.

(2 marks)

(ii) Explain why copper electrodes are not suitable for this electrolysis. (2 marks)

(b) Impure copper is purified by an electrolytic process
- (i) Name one ore from which copper is obtained (1 mark)
- (ii) Write the equation for the reaction that occur at the cathode during the purification of copper. (1 mark)
- (iii) In an experiment to electroplate a copper spoon with silver, a current of 0.5 A was passed for 18 minutes. Calculate the amount of silver deposited on the spoon (n = 96500 coulombs, Ag = 108) (3 marks)
- (iv) Give two reasons why some metals are electroplated (2 marks)

10. 2009 Q 7 P1

When aluminium oxide was electrolysed, 1800kg of aluminium metal was obtained.
- (a) Write an equation for the formation of aluminium metal. (1 mark)
- (b) Calculate the quantity of electricity in faradays used. (Al = 27) (2 marks)

11. 2009 Q 10 P1

Hydrogen and oxygen can be obtained by electrolysis of acidified water. Using equations for the reactions at the electrodes, explain why the volume of hydrogen obtained is twice that of oxygen. (2 marks)

12. 2009 Q 12 P1

The standard reduction potentials of two half cells are:

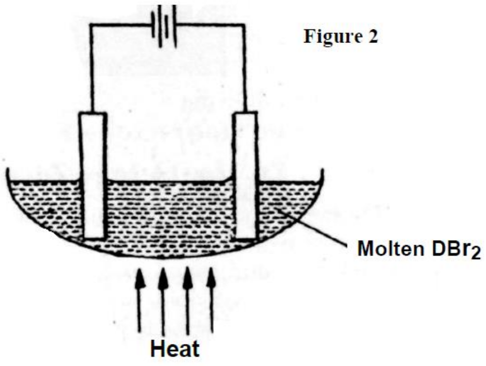
$Ag^+ (aq) + e^- \longrightarrow Ag(s)$ $E^\theta = 0.80V$

$2H_2O (l) + 2e^- \longrightarrow H_2 (g) + 2OH^- (aq);$ $E^\theta = 0.83V$

Draw a labelled diagram of an electrochemical cell that can be constructed using the two half cells. (3 marks)

13. 2009 Q 3 P2

The set-up below (figure 2) was used to electrolyse a bromide of metal D, DBr_2



(a) Write the equation for the reaction at the:
- (i) Cathode (1 mark)
- (ii) Anode (1 mark)

(b) The electrodes used in the experiment were made of carbon and metal D which of the two electrodes was used as the anode? Give a reason. (2 marks)

(c) Give a reason why this experiment is carried in a fume cupboard. (1 mark)

(d) When a current of 0.4A is passed for 90 min, 2.3g of metal D were deposited.

(i) Describe how the amount D deposited was determined. (3marks)

(ii) Calculate the relative atomic mass of metal D. (1 Faraday=96500 coulombs) (3marks)

14. 2010 Q 7 P1

Complete the table below by writing the product formed at the electrodes during the electrolysis of the electrolytes given in the table. (3 marks)

Electrolyte	**Product at anode**	**Product at cathode**
Aqueous sodium sulphate using inert electrodes	(½mark)	(½mark)
Aqueous copper (II) sulphate using copper electrodes.	(1 mark)	(1 mark)

15. 2010 Q 19 P1

The half equations involved in a cell are:

$2H_2O\,(l) + 2e^- \longrightarrow H_2\,(g) + 2OH\text{-}(aq)$: $E^\theta = -0.83V$

$O_2\,(g) + 2H_2O\,(l) + 4e^- \longrightarrow 4OH^-(aq)$: $E^\theta = +0.40V$

(a) Write the overall equation for the electrochemical cell. (1 mark)

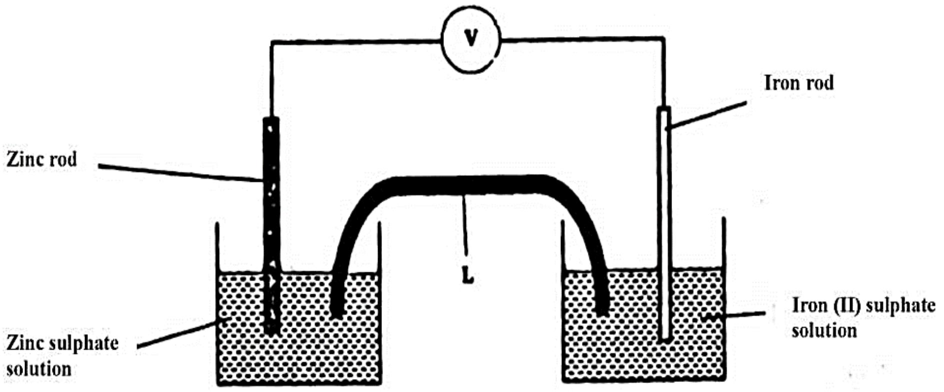
(b) Calculate the e.m.f. generated by a battery consisting of ten cells. (1 mark)

(c) State one environment advantage of using these cells in spacecrafts. (1 mark)

16. 2010 Q 1 P2

(a) Which one of the following compounds; urea, ammonia, sugar and copper (II) chloride will conduct an electric current when dissolved in water? Give reasons. (2 marks)

(b) The diagram below shows an electrochemical cell. Study it and answer the questions that follows.



Given the following;

$Fe^{2+}\,(aq) + 2e \longrightarrow Fe\,(s);\ E^\theta = -0.44V$

$Zn^{2+}\,(aq) + 2e \longrightarrow Zn\,(s);\ E^\theta = -0.76\ V$

(i) Show on the diagram using an arrow, the direction of flow of electrons (1 mark)

(ii) Name two substances that are used to fill the part labelled L (2 marks)

(c) In an experiment to electroplate iron with silver, a current of 0.5 amperes was passed through a solution of silver nitrate for one hour.

(i) Give two reasons why it is necessary to electroplate iron with silver. (2 marks)

(ii) Calculate the mass of silver that was deposited on iron (Ag = 108, 1 Faraday= 96,500 coulombs) (3 marks)

17. 2011 Q 12 P1, 2016 Q15 P1

Sodium hydroxide can be prepared by the following methods; I and II.

I. ***Sodium metal*** $\xrightarrow{\textit{cold water}}$ ***Sodium hydroxide + Hydrogen***

II. ***Concentrated sodium chloride*** $\xrightarrow{\textit{Process A}}$ ***Sodium hydroxide + Chlorine + Hydrogen***

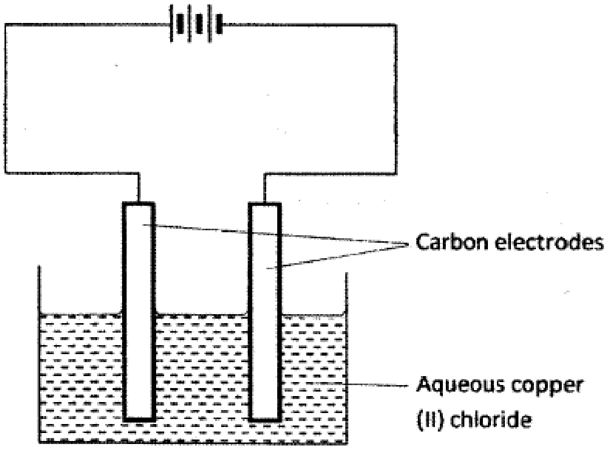
(a) Name one precaution that needs to be taken in method **I**. (1 mark)

(b) Give the name of process **A**. (1 mark)

(c) Give one use of sodium hydroxide. (1 mark)

18. 2011 Q 2 P2

The set-up below was used by a student to investigate the products formed when aqueous copper (II) chloride was electrolysed using carbon electrodes.



(a) (i) Write the equation for the reaction that takes place at the cathode. (1 mark)

(ii) Name and describe a chemical test for the product initially formed at the anode when a highly concentrated solution of copper (II) chloride is electrolysed. (3 marks)

(iii) How would the mass of the anode change if the carbon anode was replaced with copper metal? Explain. (2 marks)

(b) 0.6g of metal **B** were deposited when a current of 0.45A was passed through an electrolyte for 72 minutes. Determine the charge on the ion of metal B. (Relative atomic mass of **B** = 59, 1 Faraday = 96 500 coulombs) (3 marks)

(c) The electrode potentials for cadmium and zinc are given below:

$Cd^{2+}(aq) + 2e^- \rightleftharpoons Cd(s)$; $E^\theta = -0.4V$

$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$; $E^\theta = -0.76V$

Explain why it is not advisable to store a solution of cadmium nitrate in a container made of zinc. (2 marks)

19. 2012 Q15 P1

Below is a representation of an electrochemical cell.

$Pb(s) / Pb^{2+}(aq) // Ag^+(aq) / Ag(s)$

(a) What does **//** represent? (1 mark)

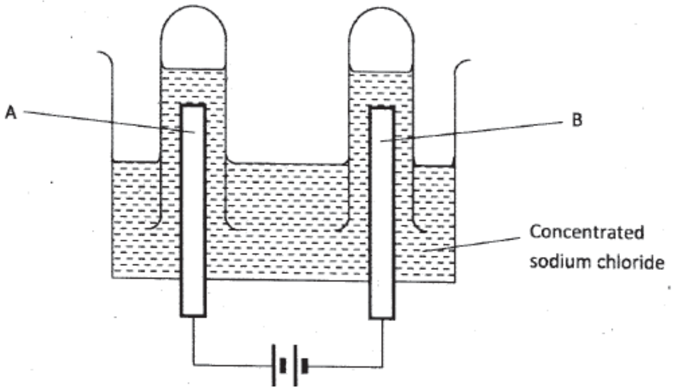
(b) Given the following:

	E^θ V
$Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$;	- 0.13
$Ag^+(aq) + e^- \longrightarrow Ag(s)$;	- 0.80

Calculate the E.M.F of the electrochemical cell. (2 marks)

20. 2012 Q28 P1

The apparatus shown in the diagram below were used to investigate the products formed when concentrated sodium chloride was electrolysed using inert electrodes.

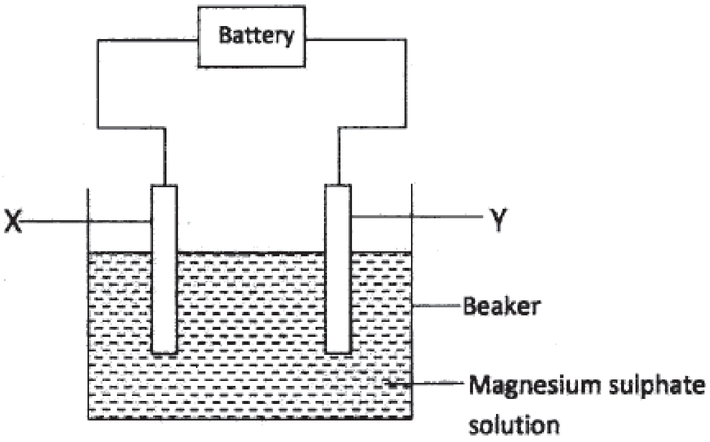


(a) Write the equation for the reaction that takes place at electrode A. (1 mark)

(b) If the concentrated sodium chloride was replaced with dilute sodium chloride, what product would be formed at electrode A? Explain. (2 marks)

21. 2012 Q5 P2

(a) The set up below was used to investigate the products formed at the electrodes during electrolysis of aqueous magnesium sulphate using inert electrodes. Use it to answer the questions that follow.



(i) During the electrolysis, hydrogen gas was formed at electrode Y. Identify the anode. Give a reason for your answer. (2 marks)

(ii) Write the equation for the reaction which takes place at electrode X (1 mark)

(iii) Why is the concentration of magnesium sulphate expected to increase during electrolysis? (2 marks)

(iv) What will be observed if red and blue litmus papers were dipped into the solution after electrolysis? (2 marks)

(b) During electrolysis of magnesium sulphate, a current of 0.3a was passed for 30 minutes. Calculate the volume of gas produced at the anode

(Molar gas volume = 24dm3; 1 faraday = 96,500C) (3 marks)

(c) State two applications of electrolysis (1 mark)

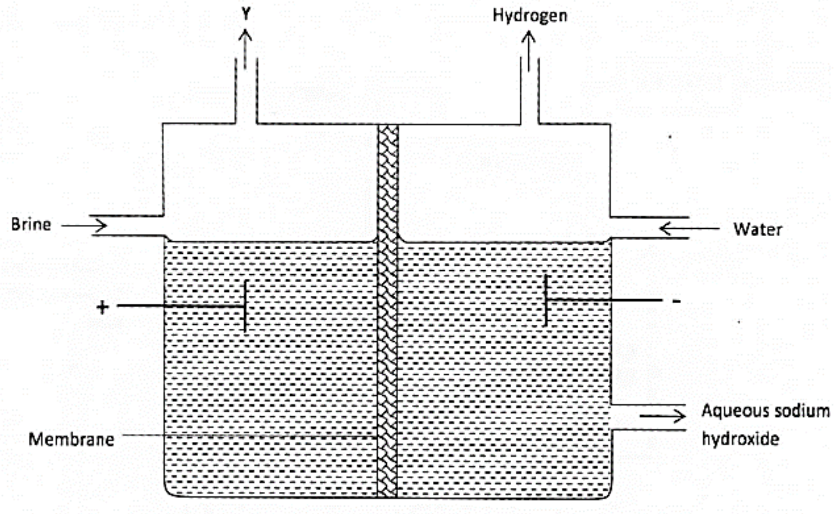
22. 2013 Q6 P1

(a) A student electroplated a spoon with copper metal. Write an equation for the process that took place at the cathode. (1 mark)

(b) Calculate the time in minutes required to deposit 1.184g of copper if a current of 2 amperes was used. (1 Faraday = 96500 coulombs, Cu=63.5). (2 marks)

23. 2013 Q4 P2, 2016 Q4 P2.

(a) The set below can be used to produce sodium hydroxide by electrolyzing brine



(i) Identify gas Y. (1 mark)

(ii) Describe how aqueous sodium hydroxide is formed in the above set-up. (2 marks)

(iii) One of the uses of sodium hydroxide is in the manufacturing of soaps. State one other use of sodium hydroxide. (1 mark)

(b) Study the information given in the table below and answer the question that follows

Half reaction	**Electrode potential E^θ (V)**
D^{2+} (aq)+2e$^-$ $\longrightarrow$ D (s)	-0.13
E^+ (aq)+e$^-$ $\longrightarrow$ E (s)	+0.80
F (aq) + e$^-$ $\longrightarrow$ F^- aq	+0.68
G^{2+}(aq) +2e $\longrightarrow$ G(s)	-2.87
H^{2+}(aq) +2e$^-$ $\longrightarrow$H (s)	+0.34
J^+(aq) + e$^-$ $\longrightarrow$J(s)	-2.71

(i) Construct an electrochemical cell that will produce the largest e.m.f (3 marks)

(ii) Calculate the emf of the cell constructed in(i) above (2 marks)

(iii) Why is it **not** advisable to store a solution containing E^+ ions in a container made of H? (2 marks)

24. 2014 Q24 P1

(a) A student electrolyzed dilute sodium chloride solution using inert carbon electrodes. Name the products at:

i) Anode:

ii) Cathode: (2 marks)

(b) If the experiment was repeated using concentrated sodium chloride instead of dilute sodium chloride solution, write the half equation at the anode. (1 mark)

25. 2015 Q11 P1

Dilute sulphuric (VI) acid was electrolysed using platinum electrodes.
Name the product formed at the anode and give a reason for your answer. (2 marks)

26. 2015 Q18 P1

Study the standard electrode potentials in the table below and answer the questions that follow.

E^θ (volts)

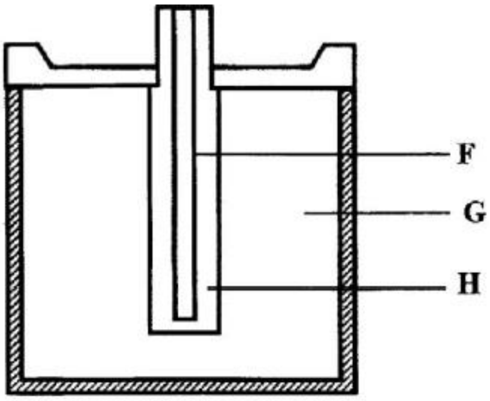
Cu^{2+} (aq) + 2e$^-$ $\longrightarrow$ Cu (s) + 0.34
Mg^{2+} (aq) + 2e$^-$ $\longrightarrow$Mg (s); - 2.38
Ag^+ (aq) + e$^-$ $\longrightarrow$Ag (s); + 0-80
Ca^{2+} (aq) + 2e$^-$ $\longrightarrow$Ca (s); -2-87

(a) Which of the metals is the strongest reducing agent? (1 mark)

(b) What observations will be made if a silver coin was dropped into an aqueous solution of copper (II) sulphate? Explain. (2 marks)

27. 2015 Q4 P2

(a) The diagram below represents a dry cell. Use it to answer the questions that follows.



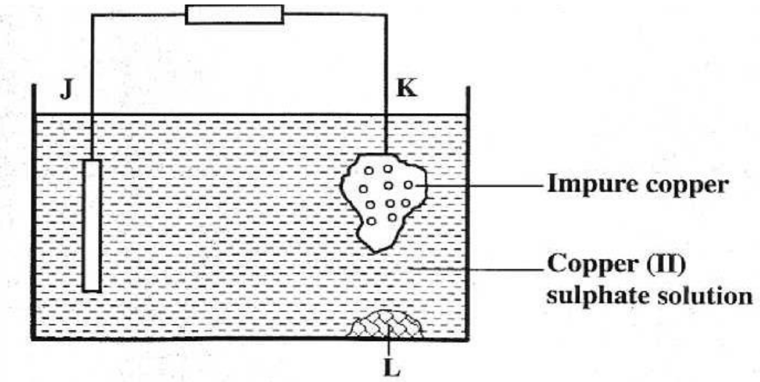
(i) Which of the letters represent;

I. Carbon electrode? (1 mark)

II. The electrolyte? (1 mark)

(ii) One of the substances used in a dry cell is manganese (IV) oxide.
State two roles of manganese (IV) oxide in the dry cells. (2 marks)

(b) Below is simplified electrolytic cell used for purification of copper. Study it and answer the questions that follows.



(i) Identify the cathode. (1 mark)

(ii) Write the equation for the reaction at the anode. (1 mark)

(iii) What name is given to **L**? (1 mark)

(iv) A current of 0.6 A was passed Through the electrolyte for 2 hours. Determine the amount of copper deposited.
(Cu=63.5; 1 Faraday = 96,500 coulombs) (3 marks)

(v) State two uses of copper metal (1 mark)

28. 2017 P1 Q3.

The diagram in **Figure 1** shows a section of a dry cell. Study it and answer the questions that follow.

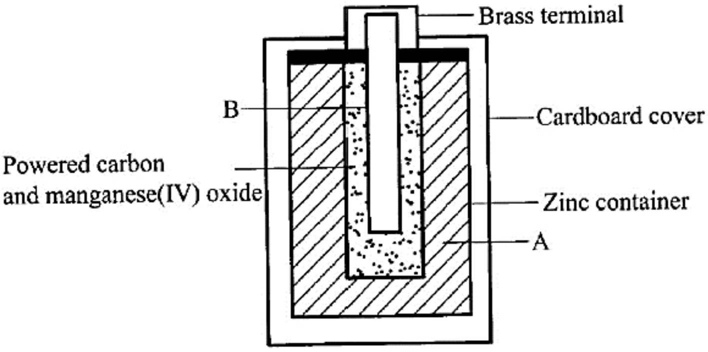


Figure 1

(a) Name the part labelled **B.** (1 mark)

(b) The part labelled **A** is a paste. Give a reason why it is not used in dry form. (1 mark)
(c) What is the purpose of the zinc container? (1 mark)

29. 2017 P1 Q22.
(a) What is an inert electrode? (1 mark)
(b) State the products formed when brine is electrolysed using inert electrodes.
Anode: ……………. (1 mark)
Cathode: ………….. (1 mark)

30. 2017 P2 Q2(b)
(b) Copper (II) sulphate solution was electrolysed using the set up in **Figure 1**.

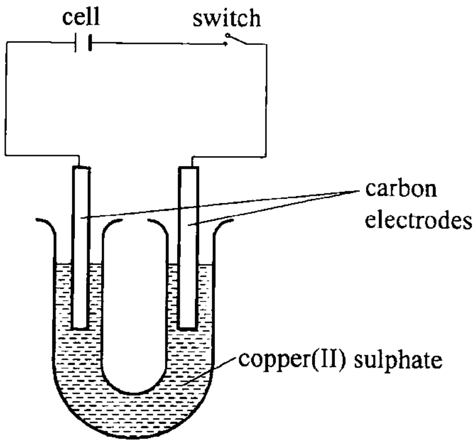


Figure 1

(i) State the observations made during electrolysis. (1½ marks)
(ii) Write the equation for the reaction that occurs at the anode. (1 mark)
(iii) State the expected change in pH of the electrolyte after electrolysis. (½mark)

(c) The experiment was repeated using copper electrodes instead of carbon electrodes. Describe the observations made at each electrode. (1 mark)

(d) Electroplating is an important industrial process.
(i) What is meant by electroplating. (1 mark)
(ii) State the purpose of electroplating. (1 mark)
(iii) During electroplating of an iron spoon, a current of 0.6 amperes was passed through aqueous silver nitrate solution for 11/2 hours. Calculate the mass of silver that was deposited on the spoon. (3 marks)
(Ag = 108.0; 1 F = 96,500 C mol^{-1})

31. 2018 P1 Q 16.
Metals **X** and **Y** have standard electrode potentials of -0.13 V and -0.76V respectively. The metals were connected to form a cell as shown in **Figure 4**.

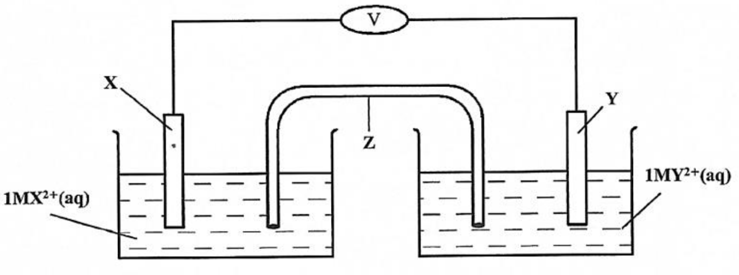


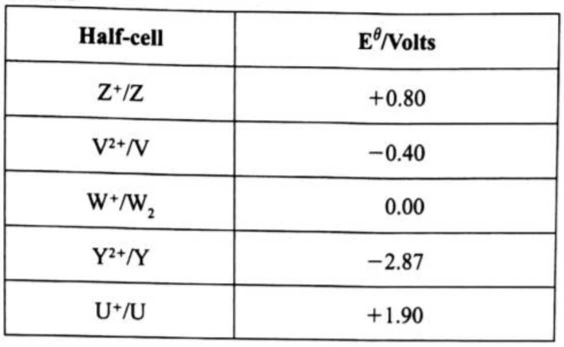
Figure 4

(a) Name the part labelled **Z.** (1 mark)
(b) State one function of the part labelled **Z**. (1 mark)
(c) Calculate the e.m.f. of the cell. (1 mark)

32. 2018 P2 Q3(c)

Use the standard electrode potentials in **Table 2** to answer the questions that follow.

Table 2



Half-cell	E^θ/Volts
Z^+/Z	+0.80
V^{2+}/V	−0.40
W^+/W_2	0.00
Y^{2+}/Y	−2.87
U^+/U	+1.90

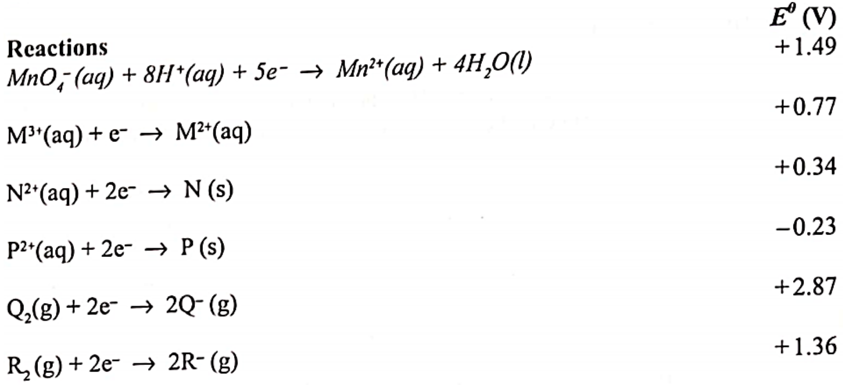
(i) Write the half-cell representation for the element whose electrode potential is for hydrogen. (1 mark)
(ii) Arrange the elements in order of reducing power, starting with the weakest reducing agent. (1 mark)
(iii) I. Select two half cells which combine to give a cell with the least e.m.f. (1 mark)
II. Calculate the e.m.f of the half cells identified in (iii) I. (1 mark)

33. 2019 P1 Q 19.

Given that the E^θ of $Cu(s)/Cu^{2+}(aq)$ is + 0.34V and that 0f $Zn(s)/Zn^{2+}(aq)$ is - 0.76V, draw a labelled diagram of zinc and copper electrochemical cell. (3 marks)

34. 2019 P2 Q6.

(a) What is meant by standard electrode potential of an element? (1 mark)
(b) Use the standard electrode potentials given below to answer the questions that follow.



Reactions	E^θ (V)
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+1.49
$M^{3+}(aq) + e^- \rightarrow M^{2+}(aq)$	+0.77
$N^{2+}(aq) + 2e^- \rightarrow N(s)$	+0.34
$P^{2+}(aq) + 2e^- \rightarrow P(s)$	−0.23
$Q_2(g) + 2e^- \rightarrow 2Q^-(g)$	+2.87
$R_2(g) + 2e^- \rightarrow 2R^-(g)$	+1.36

(i) State whether acidified MnO^{-4} can oxidise M^{2+}. Give a reason. (2 marks)
(ii) Select two half-cells which when combined will give the highest e.m.f. (1 mark)
(iii) Write the cell representation for the cell formed in b (ii). (1 mark)
(iv) Calculate the E^θ value for the cell formed in b (iii). (1 mark)

(c) A mass of 1.24g of a divalent metal was deposited when a current of 6A was passed through a solution of a metal sulphate for 12 minutes. Determine the relative atomic mass of the metal.
(1 Faraday = 96,500 C mol $^{-1}$) (3 marks)

(d) State two applications of electrolysis. (1 mark)

TOPIC FIVE

METALS

Table of Contents

Organizer

Objectives

By the end of this topic the learner should be able to:

(a) Name the chief ores of certain metals.
(b) Describe and explain the general methods used in the extraction of metals from their ores.
(c) Select and describe suitable methods of extraction of certain metals from their ores.
(d) Describe and explain physical and chemical properties of some metals.
(e) State and explain various uses of these metals and their alloys.
(f) Describe the effects of the industrial production processes of metals on the environment.

METALS

Some metals occur naturally in a free uncombined state while others are found combined with other elements. Compounds in which metals can be extracted are called **ores**.

Chief Ores of Metals

An ore from which a metal can be obtained on a commercial scale is referred to as the **chief ore**.

The table below gives the chief ores of some common metals.

Metal	Ores of the metal	Formula of ores
Sodium	Sodium chloride Rock salt*	$NaCl$
Aluminium	Bauxite* Mica Corundum	$Al_2O_3.\ 2H_2O$ $K_2Al_2.Si_6O_6$ Al_2O_3
Zinc	Zinc blende* Calamine* Zinc oxide	ZnS $ZnCO_3$ ZnO
Iron	Haematite* Magnetite Siderite Pyrite	Fe_2O_3 Fe_3O_4 $FeCO_3$ FeS_2
Copper	Copper pyrites* Malachite Cuprite Azurite	$CuFeS_2$ $CuCO_3Cu(OH)_2$ Cu_2O $Cu_3(OH)_2(CO_3)_2$
Lead	Galena* Cerussite Anglesite	PbS $PbCO_3$ $PbSO_4$

* Chief ore

Before any extraction process is carried out, qualitative and quantitative analysis is done to determine the metal present and the quantity of the metal in the ore.

General Methods of Extraction

Extraction of metals involves several stages. The stages involved include:

(a) Mining the metal ores from the earth's crust.

(b) Concentrating the ores to remove most of the impurities. The process of concentrating an ore may include:

- (i) **Removing** any **magnetic** materials with a magnet.
- (ii) **Dissolving** the mineral in a suitable solvent.
- (iii) **Washing** with water to remove earthy matter.
- (iv) **Froth flotation**. The ore is ground into a fine powder. It is then mixed with water containing special oils such as pine oil as frothing agents. A froth rich in minerals is formed at the top while the

impurities sink to the bottom. The froth is skimmed and dried. Copper pyrites, zinc blende and galena one concentrated this way.

(c) Roasting the ore to obtain metal oxides.
(d) Reducing the oxide using suitable reducing agents to obtain the desired metals.
(e) In case of more reactive metals, electrolytic extraction is applied.

The method of extraction of a metal **depends on the metal's reactivity.** The most reactive metals such as **sodium** are obtained by **electrolysis**. The less reactive metals such as iron are extracted by **reduction** of their oxides using suitable **reducing agents.**

This can be summarized in the table below.

Metal	Main ore	Main constituent compound in ore	Methods of extraction
Sodium	Rock salt	Sodium chloride, NaCl	Electrolysis of molten sodium chloride.
Aluminium	Bauxite	Hydrated aluminium oxide, $Al_2O_3.2H_2O$	Electrolysis of aluminium oxide in molten cryolite.
Zinc	Zinc blende	Zinc sulphide, ZnS	Roasting in air followed by reduction of zinc oxide by carbon
Iron	Haematite	Iron (III) oxide, Fe_2O_3	Reduction of iron (III) oxide by carbon (II) oxide.
Copper	Copper pyrites	Copper (II) sulphide, $CuFeS_2$	Heating copper sulphide in regulated amounts of air to form copper (II) oxide, Reduction of copper (II) oxide by the copper sulphide.
	Malachite	Basic copper (II) carbonate, $CuCO_3.Cu(OH)_3$	Heating in air followed by reduction by carbon.

Sodium

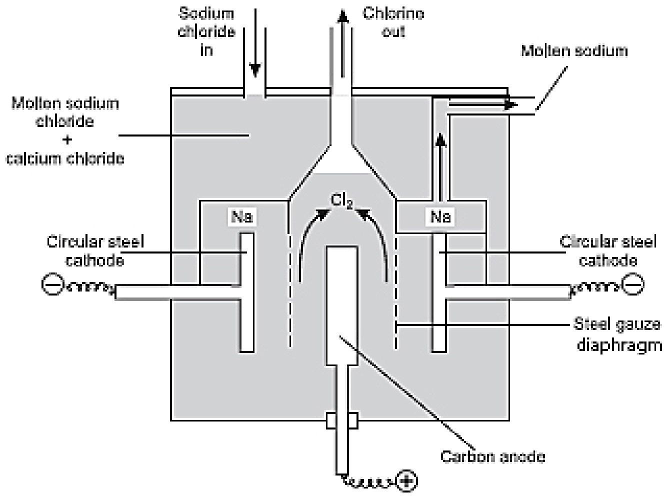
Occurrence

Sodium occurs as **dissolved chloride** in sea water and **salt lakes**. It also occurs as a double salt, **$NaHCO_3.Na_2CO_3.2H_2O$ (trona**) in salty lakes in the Rift valley such as Lake Magadi. Sodium also occurs as **rock salt (solid sodium chloride)** at various places and as **saltpetre** (solid sodium nitrate).

Saltpetre is mainly found in Chile, hence its common name, **Chile saltpetre**. The **chief ore** from which sodium is extracted is **rock salt.**

Extraction

Sodium metal is extracted by the Down's process where molten sodium chloride is electrolysed.



- The Down's cell consists of an **iron shell** lined with **heat bricks** on the outside to **maintain the high temperature** so that the **electrolyte does not crystallise**.

At the centre of the cell is a carbon anode surrounded by a steel cathode.

- **Calcium chloride** is **added** to the sodium chloride to **lower the melting point of sodium chloride from about 800°C to 600°C**. This is economical because it saves on electricity used in heating.

During electrolysis, sodium metal forms at the cathode whereas chlorine gas forms at the anode.

At the Anode

$2Cl^{-}(aq) \longrightarrow Cl_2(g) + 2e^{-}$

At the cathode

$2Na^{+}(aq) + 2e^{-} \longrightarrow 2Na(l)$

- A **steel diaphragm** is suspended between the electrodes to **prevent sodium and chlorine from recombining.**

Molten sodium is **less dense** than molten sodium chloride, hence it **rises to the top of the cathode** from where it is periodically removed. However, chlorine is not allowed into the atmosphere because it is a poisonous gas and hazardous to the environment.

Liquid calcium metal may also be produced at the cathode. However, calcium liquid **does not mix with sodium liquid as it is much denser.** In addition, calcium has a **higher melting point** compared to sodium. During cooling, **calcium crystallises first** leaving liquid sodium which is trapped.

Uses of Sodium

1. Manufacture of sodium compounds such as sodium cyanide (NaCN) and sodium peroxide (Na_2O_2). Sodium cyanide is used in the extraction of gold.
2. An alloy of sodium with lead is used in the manufacture of tetraethyl lead ($(Pb(C_2H_5)_2$ used as an anti-knock additive in petrol. This has been discontinued as use of leaded fuel has been phased out.
3. Making an alloy of sodium and potassium which is used as a coolant in nuclear reactors because the alloy is a liquid over a wide range of temperatures.
4. Sodium vapour is used in street lamps which give yellow orange light.
5. Sodium is used as a reducing agent in some reactions such as reduction of titanium(IV) chloride to form titanium metal.
6. Sodium hydroxide, a compound of sodium, is used in the manufacture of detergents, paper glass and artificial silk.

Aluminium

Occurrence

Chief ore is **bauxite ($Al_2O_3.2H_2O$)** found in France, South America, Jamaica and Ghana.

Bauxite ore has **impurities of iron(III) oxide and silica (SiO_2).** Other ores are **mica($K_2Al_2Si_6O_6$)** and **corundum (Al_2O_3).**

Extraction

Aluminum is extracted by the **electrolytic** method because it is a **reactive metal.**

The ore is concentrated before it is electrolysed.

Bauxite is **ground** into a fine powder and then dissolved in **hot concentrated sodium hydroxide under pressure**.

The **amphoteric** aluminium oxide and **acidic** silicon (IV) oxide **dissolve in the** base while iron(III) oxide which is **insoluble** in the base is filtered out as **red mud.**

$$Al_2O_3(s) + 2OH^-(aq) + 3H_2O(s) \longrightarrow 2[Al(OH)_4]^-(aq)$$

$$SiO_2(s) + 2OH^-(aq) \longrightarrow SiO_3^{2-}(aq) + H_2O(g)$$

Carbon(IV) oxide gas is bubbled through the filtrate to **precipitate** the aluminium hydroxide.

$$2[Al(OH)_4]^-(aq) + CO_2(g) \longrightarrow 2Al(OH)_3(s) + CO_3^{2-}(aq) + H_2O(l)$$

Alternatively, aluminium hydroxide may be precipitated by **seeding** process using **pure aluminium hydroxide** crystals

$$Al(OH)_4^-(aq) \xrightarrow[Seed]{Al(OH)_3(s)} Al(OH)_3(s) + OH^-(aq)$$

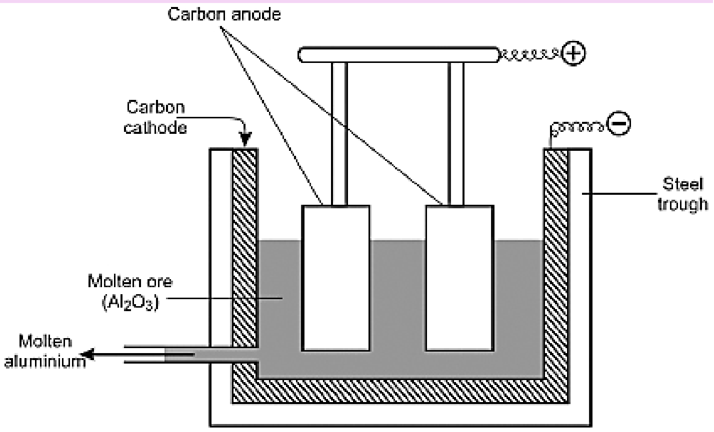
The Aluminium hydroxide is then **heated** to obtain aluminium oxide, (Al_2O_3).

$$2Al(OH)_3(s) \longrightarrow Al_2O_3(s) + 3H_2O(l)$$

The Aluminium oxide is **dissolved** in molten **cryolite (Na_3,AlF_6)** to **lower** its melting point from **2015°C to around 800°C** to save on the amount of heat needed to melt it**.**

The molten mixture is then **electrolysed** in a steel tank lined with **graphite** which acts as the **cathode**. **Graphite rods** dipping into the electrolyte act as the **anode**.

During electrolysis, aluminium is deposited at the cathode and oxygen is liberated at the anode.



At the anode

$$6O^{2-} \longrightarrow 6O(g) + 12e^-$$

$$6O(g) \longrightarrow 3O_2(g)$$

At the cathode

$$4Al^{3+}(l) + 12e^- \longrightarrow 4Al(s)$$

At the high temperature of about 800°C, the oxygen evolved **reacts** with the carbon electrode to form carbon(IV) oxide. This **corrodes** the carbon anode which should be replaced from time to time.

Uses of Aluminium

1. An alloy of aluminum and magnesium is used in making parts of aeroplanes, railway trucks, trains, buses, tankers, furniture and cars because of its low density. Aluminium **can easily be stretched** due to its **low tensile strength;** therefore, its alloys such as **duralumin** are used as they are **light, hard and strong.** Duralumin is used in the construction of aircraft and car window frames.
2. For cooking vessels such as *sufurias,* because it is a **good conductor of heat**. It is **NOT easily corroded** by cooking liquids because of the **unreative coating of aluminium oxide.**
3. For making **overhead cables,** because it is **light and is a good conductor of electricity**.
4. As a **reducing agent** in the **thermite process** in the extraction of some elements such as chromium, iron, cobalt, manganese and titanium.

$$Cr_2O_3(s) + 2Al(s) \longrightarrow 2Cr(s) + Al_2O_3(s)$$

5. **Corundum (emery)** is a natural oxide of aluminium which is useful as an **abrasive**.

Iron

Occurrence

Iron is the second most abundant metal after aluminium. The chief ore is **haematite (Fe_2O_3).** The other ores are **Magnetite (Fe_3O_4)** and **siderite ($FeCO_3$).**

The ores of iron contain **silica (SiO_2)** and **aluminium oxide as impurities.**

Extraction

Iron is usually extracted from its oxides or siderite. When extracting iron from siderite, the ore is **first roasted** in air to convert it to **iron(II) oxide** which is the stable oxide of iron.

The siderite (**carbonate**) is first decomposed by heat to form iron(II) oxide and carbon(IV) oxide.

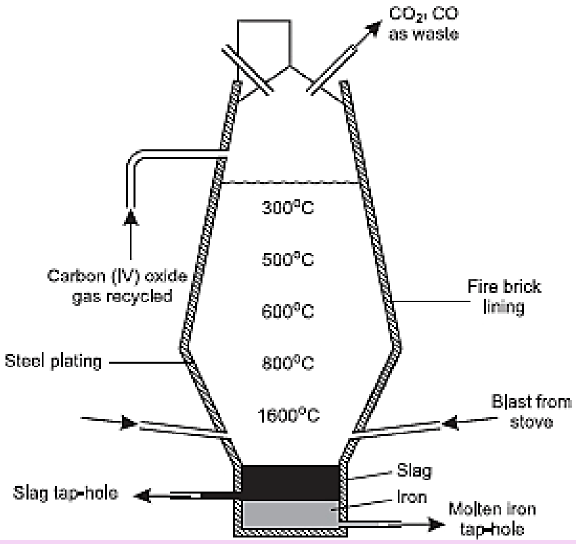
$$FeCO_3(s) \longrightarrow FeO(s) + CO_2(g)$$

Iron(II) oxide is then **oxidised** by oxygen in the atmosphere to form **iron(III) oxide.** (haematite).

$$4FeO(s) + O_2(s) \longrightarrow 2Fe_2O_3(g)$$

Iron(III) oxide is also mined as the ore. The iron(III) oxide obtained by either method is ground **into a powder** then **mixed with limestone** and **coke** then fed into a furnace from the top.

The mixture is **heated by blasts of hot air** at temperatures of between **800°C – 1000°C** from the bottom of the furnace.



The blast furnace is about **30 m high. It is made of steel** and the inner side is lined with **bricks made from magnesium oxide** which **conserve heat energy** in order to maintain the **optimum temperature** needed for the reduction of iron(III) oxide.

In the blast furnace three important reactions take place.

1. At the **bottom** of the furnace, **coke (carbon) is oxidised to carbon(IV) oxide**. The reaction **raises the temperature of the furnace to 1600 °C since it is exothermic.**

$$C(s) + O_2(g) \longrightarrow CO_2(g)\ \Delta H^{\theta} = -393\ kJ\ mol^{-1}$$

2. In the **middle** part of the furnace, **carbon (IV) oxide is reduced by coke to carbon(II) oxide** and the **temperature drops to about 1000°C since the reaction is endothermic**.

$$CO_2(g) + C(s) \longrightarrow 2CO(g)\ \Delta H^{\theta} = -283\ kJ\ mol^{-1}$$

3. At the **upper** part of the furnace where the temperature has fallen to about 700 °C, **iron(III) oxide is reduced to iron metal.** Both **carbon** and **carbon(II) oxide** act as **reducing agents.**

$$2Fe_2O_3(s) + 3C(s) \longrightarrow 4Fe(l) + 3CO_2(g)$$

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$$

Carbon(II) oxide is the **main reducing agent** because of the **large surface** that is in contact with the iron(III) oxide.

The iron produced falls to the lower part of the furnace where the temperatures are high enough to keep it molten. The **carbon(IV) oxide** produced is **recycled**.

The limestone fed into the furnace alongside coke and iron(II) oxide is decomposed by heat in the furnace to form **calcium oxide and carbon(IV) oxide.**

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

Calcium oxide being basic reacts with **acidic** and **amphoteric** oxide in the furnace to form **slag**.

$$CaO(s) + SiO_2(s) \longrightarrow CaSiO_3(l)$$

$$CaO(s) + Al_2O_3(s) \longrightarrow CaAl_2O_4(l)$$

The slag formed is **tapped off** at a higher level of the furnace because of its **lower density.**

The iron obtained from the blast furnace is **90–95%** pure and it is called "**pig Iron**". The main impurities in pig iron are **carbon, silicon, manganese, sulphur and phosphorus**, depending on the composition of the original ore.

These impurities considerably affect the properties of iron by **making it less hard and brittle** as well as **lowering the melting point of the iron.**

Uses of Iron

Cast Iron (Pig Iron)

This is the name given to the iron after it has been produced in the blast furnace. It contains about 3–5% carbon 1% silicon and 2% phosphorus. Although this type of iron has the disadvantages of being very brittle, it is extremely hard and is used in making furnaces, gates drainage pipes, engine blocks, iron boxes, etc. An important use of cast iron is the manufacture of wrought iron and steel railings for balconies.

Wrought Iron

Contains about 0.1% carbon. It is malleable and thus can be easily forged (moulded) and welded. It is used to make iron nails, iron sheets, horse shoes and agricultural implements, wrought iron is becoming less important due to increased use of mild steel.

Steel

The name is given to many different alloys whose main component is iron. The other substances may be carbon, vanadium, manganese, tungsten, nickel and chromium. Mild steel contains about 0.3% carbon. Special steel contains a small percentage of carbon together with other substances.

Mild steel is used to make nails, car bodies, railway lines, ship bodies, gliders, rods for reinforced concrete, pipes. Mild steel contains 99.75% iron and 0.75% carbon. It is easily worked on.

Stainless steel

Stainless steel contains 74% iron, 18% chromium and 8% nickel. Stainless steel containing 10–12% chromium and some nickel is used to make cutlery, sinks and vats. Steel containing 5–18% tungsten is used for making high speed cutting and drilling tools because it is tough and hard.

Cobalt Steel

This contains about 97.5% iron and 2.5% cobalt. It is tough and hard. It is highly magnetic and so it is used to make electromagnets.

Zinc

Occurrence

It occurs in many parts of the world as calamine (zinc carbonate), zinc blende (zinc sulphide). Zinc blende is often found mixed with galena (PbS). The **chief ores** of zinc are **calamine** and **zinc blende**.

Extraction

The ore is concentrated by **froth floatation**. The concentrated ore is then **roasted** to form the metal oxide.

In the case of **calamine**, carbonate decomposes to zinc oxide and carbon(IV) oxide.

$$ZnCO_3(s) \longrightarrow ZnO(s) + CO_2(g)$$

In the case of **zinc blende**, two reactions occur:

(i) Zinc sulphide is roasted in air to produce zinc oxide and sulphur(IV) oxide gas.

$$2ZnS(s) + 3O_2(g) \longrightarrow 2ZnO(s) + 2SO_2(g)$$

(ii) The impurity, lead(II) sulphide in the ore produces lead(II) oxide and sulphur(IV) oxide.

$$2PbS(s) + 3O_2(g) \longrightarrow 2PbO(s) + 2SO_2(g)$$

Zinc metal may be obtained from the oxide either by **reduction** using **carbon** or carbon monoxide or it may be **converted to zinc sulphate and electrolysed.**

(a) Reduction using carbon and carbon(II) oxide

Zinc oxide from the roaster is mixed with coke and limestone and heated in a blast furnace where it is reduced to zinc.

$$ZnO(s) + C(s) \longrightarrow Zn(g) + CO(g)$$

$$ZnO(s) + CO(g) \longrightarrow Zn(g) + CO_2(g)$$

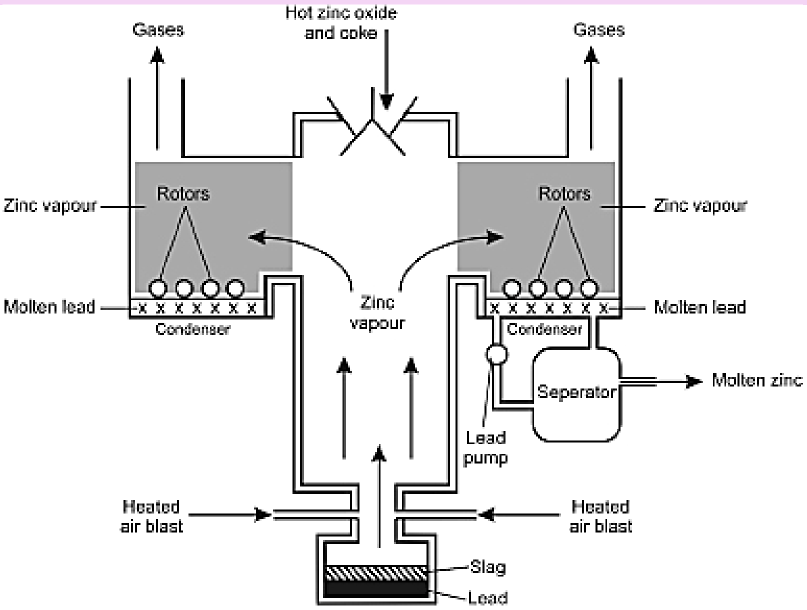
The limestone decomposes into calcium oxide and carbon(IV) oxide.

$$CaCO_3(s) \xrightarrow{Heat} CaO(s) + CO_2(g)$$

The carbon(IV) oxide is reduced by coke to carbon(II) oxide.

$$CO_2(g) + C(s) \longrightarrow 2CO(g)$$

The **carbon(II) oxide** and the **coke** are the **reducing agents.**



Zinc has a boiling point of 913°C. At the furnace temperatures which are maintained above 1,000°C, zinc exists in **vapour form**. The **zinc vapour** leaves at the top of the furnace with the hot gases.

It is cooled very rapidly to 600°C by mixing it with a **spray of molten lead**. The lead spray **condenses** the zinc and **prevents it from being re-oxidised.**

At this temperature liquid zinc separates and settles above the molten lead since it is less dense and is run off.

The zinc can be **purified by distillation.**

The lead produced during the extraction is a liquid at the furnace temperatures and it trickles to the bottom of the furnace from where it is taped off. **Calcium oxide** combines with **silica** and is removed as **slag**.

(b) Electrolytic Extraction of Zinc

The zinc oxide obtained from the roaster is converted to zinc sulphate.

$$ZnO(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2O(l)$$

Any lead oxide present reacts with the acid to form lead(II) sulphate which is **insoluble** and is therefore **precipitated**.

$$PbO(s) + H_2O_4(aq) \longrightarrow PbSO_4(s) + H_2O(l)$$

The zinc sulphate is then **dissolved in water** and the solution electrolysed. The **cathode** is made of **lead containing 1% silver** and the **anode** is made of **aluminium sheets**. The electrode reactions are:

Cathode

Zinc ions are discharged.

$$2Zn^{2+}(aq) + 4e^{-} \longrightarrow Zn(s)$$

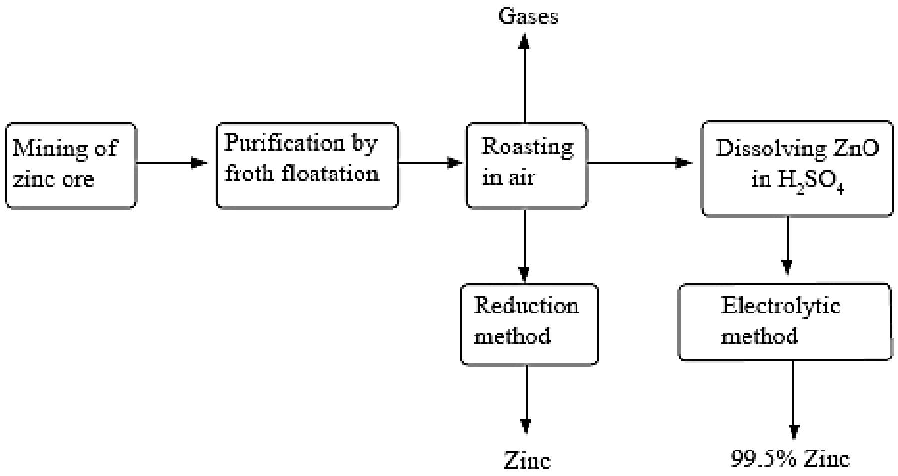
If graphite electrode were used, **hydrogen gas would have been evolved instead.** Zinc is stripped off the cathode regularly. The metal is about **99.5% pure**.

Anode

Hydroxide ions are discharged in preference to sulphate ions.

$$4OH^{-}(aq) \longrightarrow 2H_2O(l) + O_2(g) + 4e^{-}$$

Over **80%** of zinc is extracted by the electrolytic methods.



Flow-chart to summarise the extraction of zinc

Uses of Zinc

1. Zinc is used to galvanise iron to prevent it from rusting.
2. To make brass, an alloy of copper and zinc.
3. Making of outer casing in dry batteries.

Lead

Its chief ore is **galena(PbS).** Other ores of less industrial importance are **cerussite ($PbCO_3$)** and **anglesite ($PbSO_4$)**.

Extraction

The ore is first **ground** into a fine powder and then **concentrated through froth floatation.** The concentrated ore is then **roasted** in air to obtain **lead(II) oxide** and **sulphur(IV) oxide**.

$$2PbS(s) + 3O_2(g) \longrightarrow 2PbO(s) + 2SO_2(g)$$

The lead(II) oxide from the roaster is **mixed with coke** and **calcium carbonate (limestone).** The mixture is then heated in a blast furnace.

The lead(II) oxide is reduced by the coke to lead.

$$2PbO(s) + C(s) \longrightarrow Pb(s) + CO_2(g)$$

Iron is added to the blast furnace **to reduce any remaining lead sulphide** to lead.

$$Fe(s) + PbS(s) \longrightarrow FeS(l) + Pb(s)$$

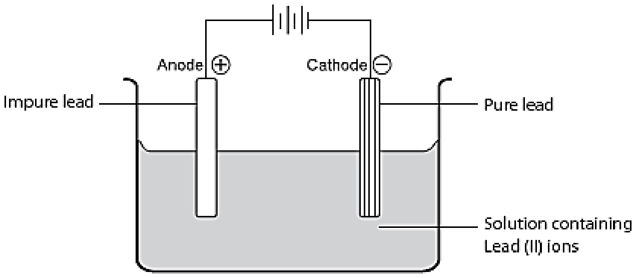
Calcium oxide combine with silica to form calcium silicate.

$$CaO(s) + SiO_2(s) \longrightarrow CaSiO_3(l)$$

The **iron(II) sulphide** and **calcium silicate** form **slag** which is tapped off separately from the lead. The lead produced by the method is not pure.

Pure lead is obtained by **blowing a blast of air through the molten impure lead.** Oxygen in the air **oxidises** the impurities into **compounds which are less dense than lead.** These compounds **float** on the molten lead and are **skimmed off**.

More **pure lead** can be obtained by electrolysis.The **cathode** is made of a **pure strip of lead** while the **impure lead is made the anode**. At the cathode, lead is deposited while at the anode lead dissolves.



At the **cathode:** $\mathbf{Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)}$

At the **anode:** $\mathbf{Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^-}$

Uses of Lead

1. Manufacture of storage batteries (lead acid accumulators) .
2. It is used in ammunition (shot and bullets) and as a constituent of solder, type metal, bearing alloys, fusible alloys, and pewter.
3. In heavy and industrial machinery, sheets and other parts made from lead compounds may be used to dampen noise and vibration.
4. **Lead Pipes-**Lead pipes due to its corrosion resistant properties are used for carriage of corrosive chemicals at chemical plants.
5. Lead Sheet is used in the building industry for flashings or weathering to prevent water penetration & for roofing and cladding. By virtue of its resistance to chemical corrosion, Lead Sheet also finds use for the lining of chemical treatment baths, acid plants and storage vessels.
6. Because lead effectively absorbs electromagnetic radiation of short wavelengths, it is used as a protective shielding around nuclear reactors, particle accelerators, X-ray equipment, and containers used for transporting and storing radioactive materials.

Copper

Occurrence

Copper ores include, **pyrites ($CuFeS_2$), cuprite (Cu_2O_2), chalcocite (Cu_2S) and malachite ($CuCO_3.Cu(OH)_2$.** It also occurs in uncombined state in various parts of the world such as Canada, USA, Zambia, Tanzania and the Democratic Republic of Congo (DRC).

Impurities in copper ores may include **traces of gold and silver**.

The chief ore of copper is **copper pyrites.**

Extraction

Copper is mostly extracted from **copper pyrites**. The ore is first **crushed** into a fine powder and **concentrated by froth flotation.** The concentrated ore is then **roasted in a limited supply of air** to obtain copper(I) sulphide and iron(II) oxide

$$2CuFeS_2(s) + 4O_2(g) \longrightarrow Cu_2S(s) + 2FeO(s) + 3SO_2(g)$$

Silica (SiO_2) is then added and the mixture is heated in the **absence of air.** The silica **reacts with iron(II) oxide to form iron(II) silicate** which separates out as a **slag** leaving behind the copper(I) sulphide.

$$FeO(l) + SiO_2(l) \longrightarrow FeSiO_3(l)$$

The copper(I) sulphide is then **heated in a regulated supply of air** where some of it is converted to copper(I) oxide.

$$2Cu_2S(l)) + 3O_2(g) \longrightarrow 2Cu_2O(l) + 2SO_2(g)$$

The copper(I) oxide then **reacts** with the **remaining copper(I) sulphide** to form copper metal and sulphur(VI) oxide.

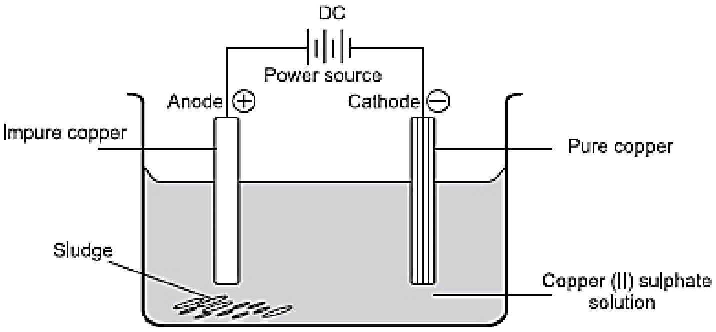
$$Cu_2\,S(l) + 2Cu_2O(l) \longrightarrow 6Cu(l) + SO_2(g)$$

The sulphur(IV) oxide produced in the process is **either fed into adjacent sulphuric(IV) acid plant or scrubbed using calcium hydroxide.**

$$SO_2(g) + Ca(OH)_2(aq) \longrightarrow CaSO_3(s) + H_2O(l)$$

The copper obtained in this process is about 97.5% pure. This is called **blister copper**.

It is refined by electrolysis to obtain 99.8% pure copper. During the refining, stripes of pure copper are used as the cathode whereas the anode is made of the impure copper. Copper(II) sulphate solution is used as the electrolyte.



During electrolysis, the **impure copper anode goes into solution as copper ions** while **copper metal is deposited on the pure copper cathodes.**

At the anode: $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$

At the cathode $Cu^{2+}(aq) + 2e– \longrightarrow Cu(s)$

Traces of **gold** and **silver** collect as a **sludge** at the bottom of the electrolytic cell.

Where **copper carbonate is the chief ore, it is roasted first to obtain copper(II) oxide**.

$$CuCO_3\,(s) \longrightarrow CuO(s) + CO_2(g)$$

The copper(II) oxide is then reduced to copper metal using coke and carbon(II) oxide as reducing agents.

$$2CuO(s) + C(s) \longrightarrow Cu(s) + CO_2(g)$$

$$CuO(s) + CO(g) \longrightarrow Cu(s) + CO_2(g)$$

Uses of Copper

1. Making electrical wires and contacts in switches, plugs and sockets because copper is a good conductor of electricity. Pure copper is necessary for this use because impurities increase electrical resistance.
2. Making soldering instruments due to its high thermal conductivity.
3. Making alloys such as brass (Cu and Zn), bronze (Cu and Sn), German silver (Cu, Zn and Ni), etc.
4. Making coins and ornaments.

Properties of Some Metals

Physical Properties of Metals

Physical properties of metals depend on the size of the atoms, their electron arrangement and the crystal lattice.

- Metals generally have **high melting and boiling points due to strong metallic bonds**.
- They are **good conductors of both heat and electricity** due to the presence of **delocalised valence electrons** in the metallic lattice. The number of delocalised electrons and their ease of movement within the lattice account for the **difference** in electrical conductivity.

Metal	Melting point (°C)	Boiling point (°C)	Thermal & Electrical conductivity	Density (gcm^{-3})	Malleability	Ductility
Sodium	97.8	890	Good	0.97	-	-
Aluminium	660	2 470	Very good	2.70	Malleable	Ductile
Zinc	1 535	3 000	Good	2.86	Malleable	Ductile
Iron	420	907	Good	2.14	Malleable	Ductile
Copper	1083	2 395	Very good	8.90	Malleable	Ductile

- The metals generally have **high density**. Differences in density in metals are mainly due to different atomic masses, packing of the atoms in the metallic lattice and the size of the atoms.
- Metals can be **pressed into sheets and also drawn into wires.** These properties are referred to as **malleability** and **ductility** respectively.

Chemical Properties of Metals

Reactions with air

Freshly cut or polished surfaces of metals have a shiny appearance. Sodium rapidly tarnishes in dry air forming sodium oxide.

$$4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$$

In moist air, both sodium metal and sodium oxide react with water vapour to form sodium hydroxide.

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

$$Na_2O(s) + 2H_2O(l) \longrightarrow 2NaOH(aq)$$

The resulting alkali absorbs acidic carbon(IV) oxide in the air to form sodium carbonate.

$$2NaOH(aq) + CO_2\ (g) \longrightarrow Na_2CO_3\ (aq) + H_2O(l)$$

When heated in air sodium readily burn with a yellow flame to form mainly sodium peroxide.

$$2Na(s) + O_2(g) \longrightarrow Na_2O_2(s)$$

Polished aluminium is immediately coated by aluminium oxide. If heated to 800 °C, aluminium burns to form aluminium oxide and aluminium nitride.

$$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$$

$$2Al(s) + N_2(g) \longrightarrow 2AlN(s)$$

Zinc tarnishes very slowly in air due to the formation of zinc oxide. When heated, zinc burns to form zinc oxide

$$2Zn(s) + O_2(g) \longrightarrow 2ZnO(s)$$

Iron rusts in the presence of moist air to form hydrated iron(III) oxide, $Fe_2O_3.H_2O(s)$.

$$3Fe(s) + 2H_2O(l) + 3O_2(g) \longrightarrow 2Fe_2O_3.H_2O(s)$$

When heated, iron reacts with oxygen to form tri-iron tetraoxide, Fe_3O_4.

$$3Fe(s) + 2O_2(g) \longrightarrow Fe_3O_4(s)$$

Copper forms a black coating of copper(II) oxide when heated in air. Finely divided copper burns with a blue flame to form copper(II) oxide.

$$\underset{\textit{Brown}}{2Cu(s)} + O_2(g) \longrightarrow \underset{\textit{black}}{2CuO(s)}$$

Reactions with cold water and steam

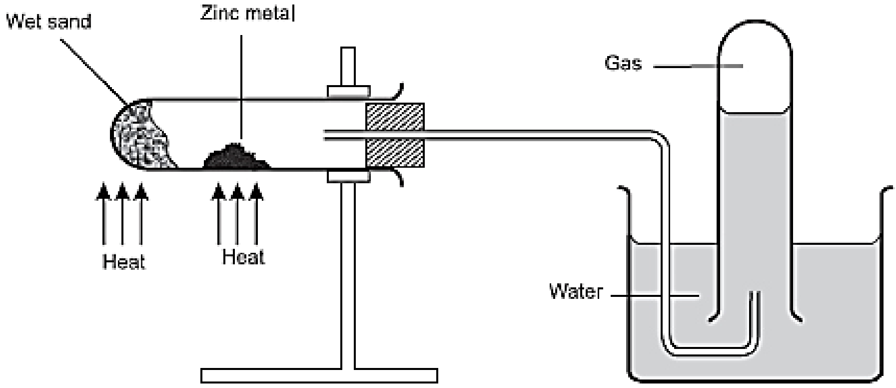
Sodium reacts **vigorously** with cold water liberating hydrogen gas.The resulting solution is alkaline(basic). The alkaline solution is **sodium hydroxide.**

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH + H_2\,(g)$$

Aluminium, zinc and iron do not readily react with cold water.

Aluminium does not react with cold water **because of a thin layer of aluminium oxide on its surface.** If the thin layer of aluminium oxide is removed, aluminium reacts with cold water **very slowl**y liberating hydrogen gas.

Copper does not react with cold water.



Aluminium, zinc and iron react with steam liberating hydrogen gas and forming metals oxides.

$$2Al(s) + 3H_2O(g) \longrightarrow Al_2O_3(s) + 3H_2(g)$$
$$Zn(s) + H_2O(g) \longrightarrow ZnO(s) + H_2(g)$$
$$3Fe(s) + 4H_2O(g) \longrightarrow Fe_3O_4(s) + 4H_2(g)$$

Copper **does not react with steam at all.**

Among the metals discussed, sodium is the most reactive while copper is the least reactive.

Sodium	Most reactive
Aluminium	
Zinc	Decreasing order of reactivity
Iron	
Copper	Least Reactive

The position of **aluminium** is **not easy to determine** in this experiment because its reaction with cold water and steam are **impaired due to the presence of an oxide coating.**

During the reaction between the metal and water, **the reactive metals displace hydrogen to form the hydroxides of the metals.**

Copper **does not** react with water because it is **low** in the reactivity series.
If hydrogen is put together with metals in order of reactivity, its position would be higher than that of copper.

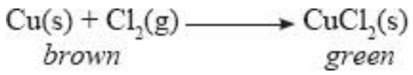
Sodium	Most reactive
Aluminium	
Zinc	Decreasing order of reactivity
Iron	
Hydrogen	
Copper	Least Reactive

Reaction with chlorine

Hot sodium metal reacts with chlorine to form white fumes of sodium chloride.

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

Hot copper glows red in chlorine gas forming brown copper(II) chloride which turns green in the presence of moisture.



$$\underset{brown}{Cu(s) + Cl_2(g)} \longrightarrow \underset{green}{CuCl_2(s)}$$

Hot zinc reacts with chlorine gas to form white zinc chloride.

$$Zn(s) + Cl_2(g) \longrightarrow ZnCl_2(s)$$

Hot iron **glows red** in chlorine gas. This is because the **reaction is exothermic.** During the reaction **brown** fumes are observed. The fumes solidify on the cooler parts of the tube to form black crystals of iron(III) chloride.

$$Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$$

Hot aluminium burns in chlorine to form a white solid which sublimes and condenses on the cooler parts of the apparatus.

$$2Al(s) + 3Cl_2(g) \longrightarrow 2AlCl_3(s)$$

Both iron(III) chloride and aluminium chloride are observed to fume when exposed in damp air. This is because both chlorides are readily hydrolysed by water vapour to produce hydrogen chloride gas.

$$FeCl_3(s) + 3H_2O(l) \longrightarrow Fe(OH)_3(s) + 3HCl(g)$$

$$AlCl_3(s) +3H_2O(l) \longrightarrow Al(OH)_3(s) + 3HCl(g)$$

Reactions with acids

Polished aluminium reacts **very slowly** with dilute hydrochloric acid to liberate hydrogen gas and form aluminium chloride.

$$2Al(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2(g)$$

There is **no apparent reaction between aluminium and dilute sulphuric(VI) acid.**

There is also **no apparent** reaction between aluminium and nitric(V) acid at any concentration. This is because **nitric(V) acid is a strong oxidising agent.** A **thin layer of aluminium oxide** forms on the metal surface **immediately** it comes into contact with the acid.

Aluminium **reduces** hot concentrated **sulphuric(VI) acid** to sulphur(IV) oxide and is itself oxidized to aluminium sulphate.

$$2Al(s) + 6H_2SO_4(l) \longrightarrow Al_2(SO_4)_3(aq) + 6H_2O(l) + 3SO_2(g)$$

Iron reacts with both dilute hydrochloric acid or sulphuric (VI) acid to liberate hydrogen gas.

$$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$$

$$Fe(s) + H_2SO_4(aq) \longrightarrow FeSO_4(aq) + H_2(g)$$

Iron reduces hot concentrated sulphuric(VI) acid to sulphur(IV) oxide and is itself oxidised to iron(II) sulphate.

$$2Fe(s) + 6H_2SO_4(s) \longrightarrow Fe_2(SO_4)_3 + 6H_2O(l) + 3SO_2(g)$$

Dilute nitric(V) acid reacts with iron to form a mixture of nitrogen(I) oxide and nitrogen(II) oxide.

There **is no apparent reaction between iron and concentrated nitric(V) acid. The concentrated acid oxidises iron and forms a thin layer of impervious tri-iron tetraoxide, Fe_3O_4, which prevents further reaction.**

Zinc reacts with both dilute hydrochloric acid and sulphuric(VI) acid to liberate hydrogen gas.

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

Zinc reduces concentrated sulphuric(VI) acid to sulphur(IV) oxide gas and is itself oxidised to zinc sulphate.

$$Zn(s) + 2H_2SO_4(l) \longrightarrow ZnSO_4(aq) + 2H_2O(l) + SO_2(g)$$

Zinc reacts with 50% concentrated nitric(V) acid to liberate nitrogen (II) oxide gas.

$$3Zn(s) + 8HNO_3(aq) \longrightarrow 3Zn(NO_3)_2(aq) + 4H_2O(l) + 2NO(g)$$

Copper **does not react with dilute hydrochloric acid, dilute sulphuric(VI) acid or very dilute (less than 50%) nitric(V) acid**. It however reacts with nitric(V) acid at 50% concentration to liberate nitrogen(II) oxide.

$$3Cu(s) + 8HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 4H_2O(l) + 2NO(g)$$

With concentrated nitric(V) acid and sulphuric(VI) acid, nitrogen(IV) oxide and sulphur(IV) oxides are produced respectively.

$$Cu(s) + 4HNO_3(l) \longrightarrow Cu(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$$

$$Cu(s) + 2H_2SO_4(l) \longrightarrow CuSO_4(aq) + 2H_2O(l) + 2SO_2(g)$$

Effects of Extraction of Metals on the Environment

Extraction of metals leads to land pollution, air pollution and water pollution.

Mining of the ores from the ground may lead to gaping holes being left in the ground if not refilled with earth. If undesired earthy material accompanying the ores are carelessly disposed of, it may lead to serious land pollution.

Roasting of the ores, reduction of the oxide and electrolysis of chlorides leads to evolution of gaseous by-products such as sulphur(IV) oxides, carbon(IV) oxide gas and chlorine gas. These gaseous products could lead to air pollution if allowed to escape into the atmosphere This could lead to acid rain effects.

The gaseous by-products such as chlorine gas and sulphur(IV) oxide gas are fed into hydrochloric acid and sulphuric(VI) acid plants respectively.

Solid by-products such as slag could lead to land pollution if not disposed off safely. Conversely, solid by products can be disposed off carefully by either burning them or making other uses of them. For example, slag may be used in carpeting roads.

Review Exercises

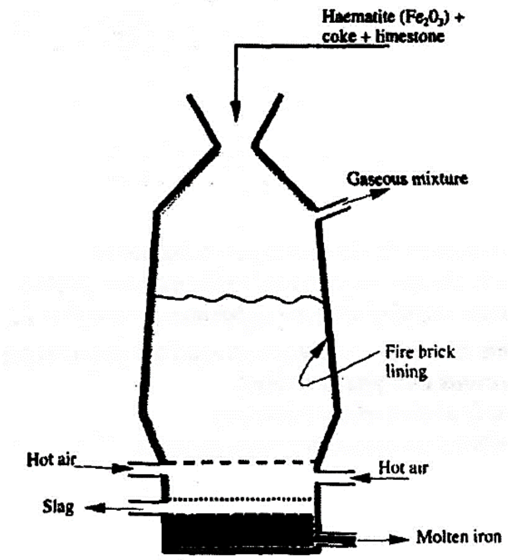
1. 2006 Q 21

(a) Explain why the metals magnesium and aluminium are good conductors of electricity. (1 mark)

(b) Other than cost, give two reasons why aluminium is used for making electric cables while magnesium is not (2 marks)

2. 2006 Q 6 P2

The extraction of iron from its ores takes place in the blast furnace. Study it and answer the questions that follow.



(a) Name:
  (i) One of the substances in the slag; (1 mark)
  (ii) Another iron ore material used in the blast furnace; (1 mark)
  (iii) One gas which is recycled. (1 mark)

(b) Describe the process which leads to the formation of iron in the blast furnace. (3 marks)

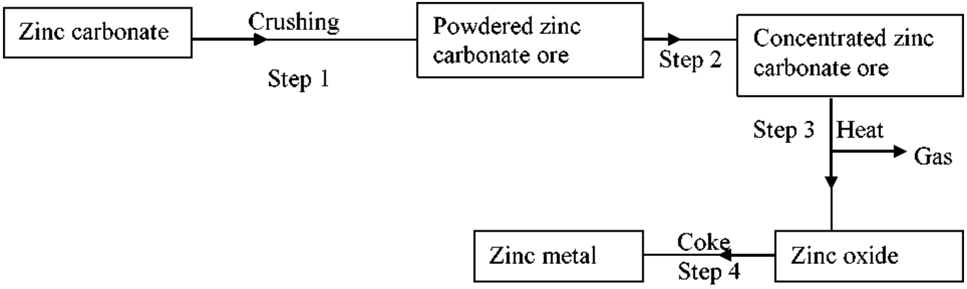
(c) State the purpose of limestone in the blast furnace. (2 marks)

(d) Give a reason why the melting point of the iron obtained from the blast furnace is 1200 °C while that of pure iron is 1535 °C (1 mark)

(e) State two uses of steel (2 marks)

3. 2007 Q 19 P1

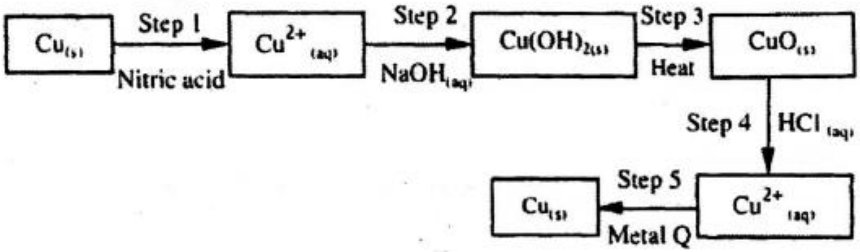
The flow chart below shows steps used in the extraction of zinc from one of its ores.



(a) Name the process that is used in step 2 to concentrate the ore. (1 mark)
(b) Write an equation for the reaction which takes place in step 3. (1 mark)
(c) Name one use of zinc other than galvanizing. (1 mark)

4. 2007 Q 3 P2

The flow chart below shows a sequence of chemical reactions starting with copper study it and answer the questions that follow.



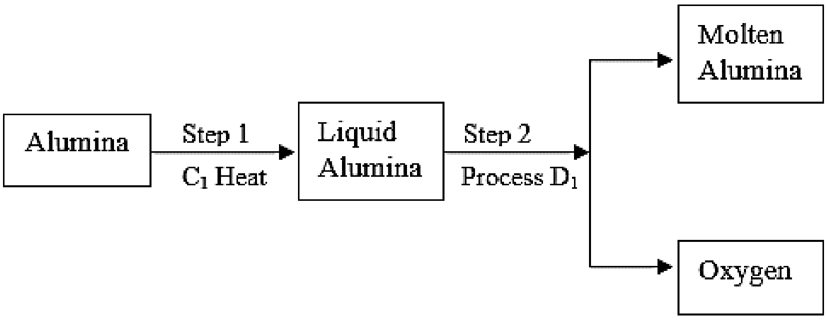
(a) (a) In step 1, excess 3M nitric acid was added to 0.5g of copper powder.
 (i) State two observations which were made when the reaction was in progress (2marks)
 (ii) Explain why dilute hydrochloric acid cannot be used in step 1 (1mark)
 (iii) I. Write the equation for the reaction that took place in step 1 (1mark)
  II. Calculate the volume of 3M nitric that was needed to react completely with 0.5g of copper powder. (Cu = 63.5) (3 marks)

(b) Give the names of the types of reactions that took place in steps 4 and 5. (1 mark)

(c) Apart from the good conductivity of electricity, state two other properties that make it possible for copper to be extensively used in the electrical industry. (2marks)

5. 2008 Q 28 P1

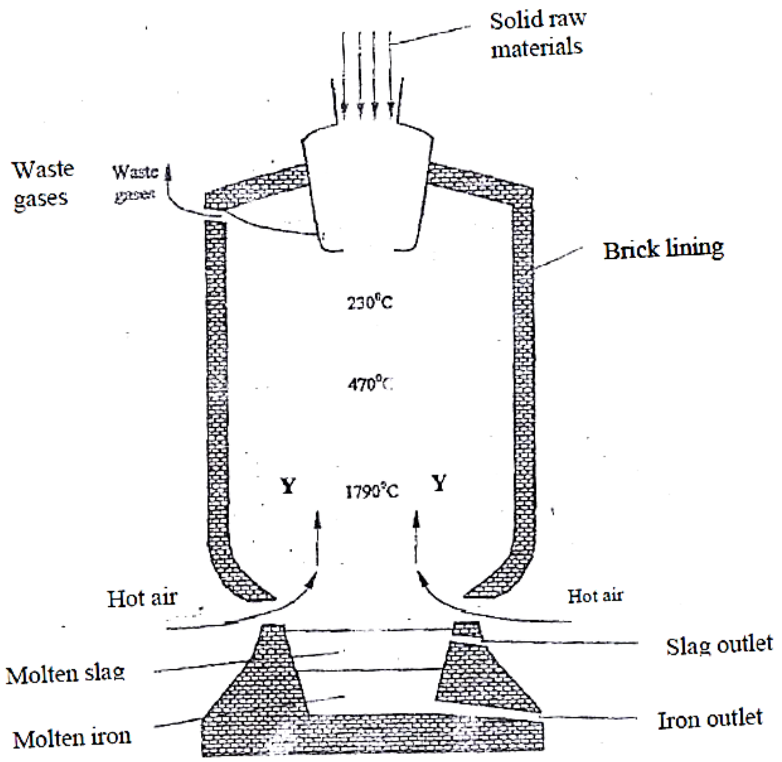
During the extraction of aluminium from its ores; the ore is first purified to obtain alumina.



(a) Name
 (i) Substance C_1 (1 mark)
 (ii) Process D_1 (1 mark)
(b) Give two reasons why aluminium is used extensively in the making of cooking pans. (1 mark)

6. 2009 Q 7 P2

Iron is obtained from hematite using a blast furnace shown in **figure 5** below.



(a) Four raw materials are required for the production of iron. Three of these are iron oxide, hot air and limestone. Give the name of the fourth raw material.(1 mark)

(b) Write an equation for the reaction in which carbon (IV) oxide is converted into carbon (II) oxide. (1 mark)

(c) Explain why the temperature in the region marked Y is higher than of the incoming hot air. (2 marks)

(d) State one physical property of molten slag other than density that allows it to be separated from molten iron as shown in the figure 5. (1 mark)

(e) One of the components of the waste gases is Nitrogen (IV) oxide. Describe the adverse effect it has on the environment. (2 marks)

(f) Iron from the blast furnace contains about 5% carbon

(i) Describe how the carbon content is reduced (2 marks)
(ii) Why is it necessary to reduce the carbon content? (1 mark)

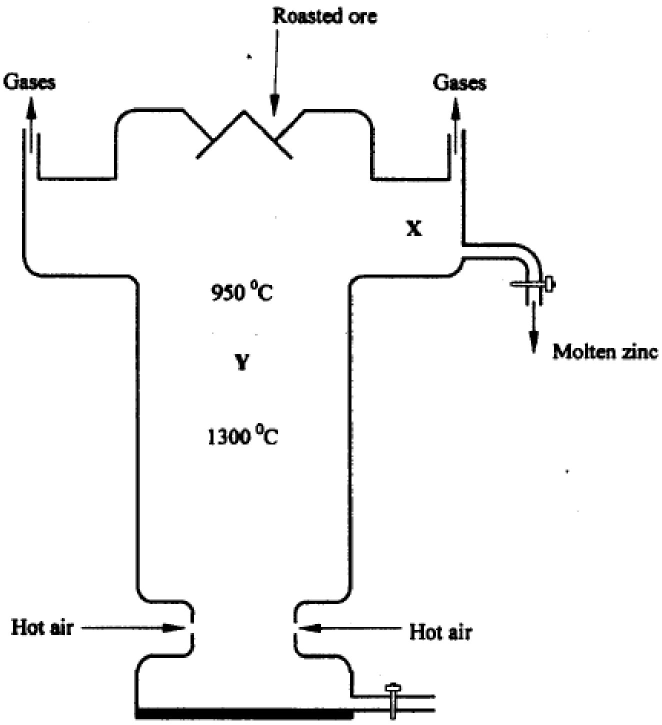
7. 2010 Q 6 P2

The melting and boiling points of zinc are 419 °C and 907 °C respectively.

One of the ores of zinc is zinc blende. To extract zinc, the ore is first roasted in air before feeding it into a furnace.

(a) (i) Write the formula of the main zinc compound in zinc blende. (1 mark)
(ii) Explain using an equation why it is necessary to roast the ore in air before introducing it into the furnace. (2 marks)

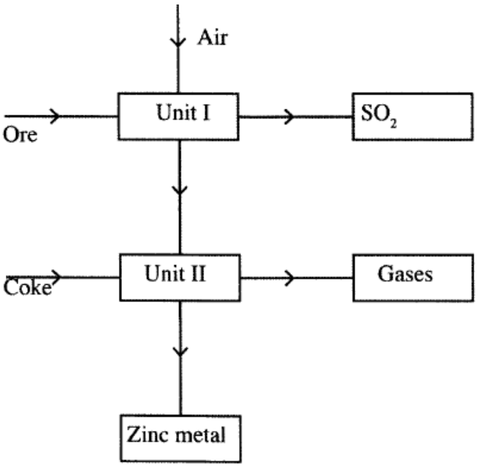
(b) The diagram below shows a simplified furnace used in the extraction of zinc. Study it and answer the questions that follows:



(i) Name two other substances that are also introduced into the furnace together with roasted ore. (1 mark)
(ii) The main reducing agent in the furnace is carbon (II) oxide. Write two equations showing how it is formed. (2 marks)
(iii) In which physical state is zinc at point Y in the furnace? Give a reason. (1 mark)
(iv) Suggest a value for the temperature at point X in the furnace. Give a reason. (1 mark)
(v) State and explain one environmental effect that may arise from the extraction of zinc from zinc blende (2 marks)
(vi) Give two industrial uses of zinc. (1 mark)

8. 2011 Q 27 P1

The flow chart below shows some processes involved in the industrial extraction of zinc metal.



(a) Name one ore from which zinc is extracted. (1 mark)
(b) Write the equation of the reaction taking place in unit II. (1 mark)
(c) Name two uses of zinc metal. (1 mark)

9. 2012 Q22 P1

Aluminium is both malleable and ductile.

(a) What is meant by?
  (i) Malleable: (1 mark)
  (ii) Ductile (1 mark)
(b) State one use of aluminium based on:
  (i) malleability (½ mark)
  (ii) ductility (½ mark)

10. 2013 Q22 P1

(a) Name two ores from which copper is extracted. (1 mark)
(b) During extraction of copper metal, the ore is subjected to froth flotation. Give a reason why this process is necessary. (1 mark)
(c) Name one alloy of copper and state its use. (1 mark)

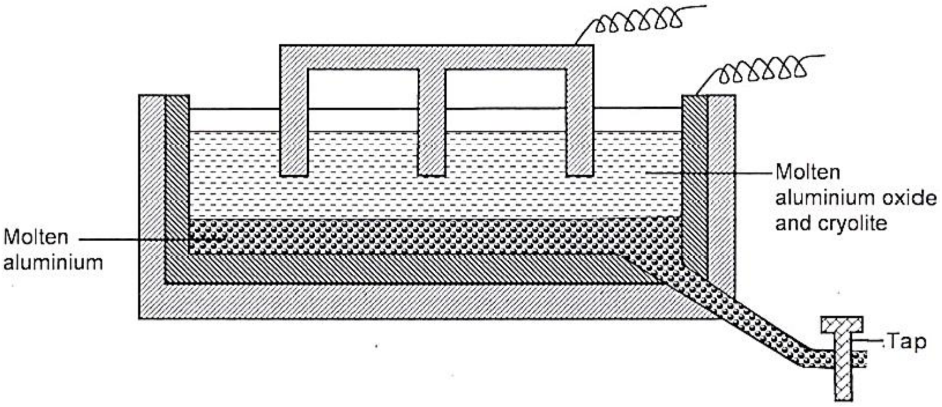
11. 2014 Q10 P1

One of the ores of copper has formula, $CuFeS_2$.

(a) Describe how iron in the ore is removed during concentration of copper metal. (1 mark)
(b) State two environmental problems associated with extraction of copper metal. (2 marks)

12. 2014 Q6 P2, 2016 P2 Q12.

The diagram below represents a set-up of an electrolytic cell that can be used in the production of aluminium



(a) One the diagram, label the anode. (1 mark)
(b) Write the equation for the reaction at the anode. (1 mark)

(c) Give a reason why the electrolytic process is not carried out below 950 °C. (1 mark)

(d) Give a reason why the production of aluminium is not carried out using reduction process. (1 mark)

(e) Give two reasons why only the aluminium ions are discharged. (2 marks)

(f) State two properties of duralumin that makes it suitable for use in aircraft industry. (1 mark)

(g) Name two environmental effects caused by extraction of aluminium. (2 marks)

13. 2015 Q2 P1

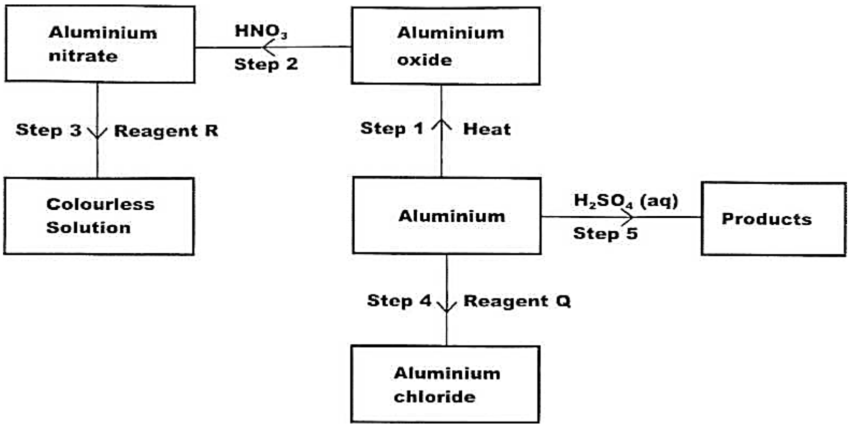
(a) Name the raw material from which sodium is extracted. (1 mark)

(b) Give a reason why sodium is extracted using electrolysis. (1 mark)

(c) Give two uses of sodium metal. (1 mark)

14. 2015 Q16 P1

The flow chart below shows various reactions of aluminium metal. Study it and answer the questions that follow.



(a) (i) Other than water, name another reagent that could be R. (1 mark)

(ii) Write the formula of reagent Q. (1 mark)

(b) Write an equation or the reaction in step 5. (1 mark)

15. 2017 P1 Q27.

(a) Name two ores in which sodium occurs. (1 mark)

(b) During extraction of sodium using the down's process, calcium chloride is added to the ore. Give a reason for the addition of calcium chloride. (1 mark)

(c) State two uses of sodium. (1 mark)

16. 2017 P2 Q6.

The following steps were used to analyse a metal ore.

(i) An ore of a metal was roasted in a stream of oxygen. A gas with a pungent smell was formed which turned acidified potassium dichromate (VI) green.

(ii) The residue left after roasting was dissolved in hot dilute nitric(V) acid. Crystals were obtained from the solution.

(iii) Some crystals were dried and heated. A brown acidic gas and a colourless gas were evolved and a yellow solid remained.

(iv) The solid was yellow when cold.

(v) The yellow solid was heated with powered charcoal. Shiny beads were formed.

(a) Name the:

(i) Gas formed when the ore was roasted in air; (1 mark)
(ii) Gases evolved when crystals in step (iii) were heated; (2 marks)
(iii) Yellow solid formed in step (iii); (1 mark)
(iv) Shiny beads in step (iv). (1 mark)

(b) The yellow solid from procedure (iii) was separated, dried, melted and the melt electrolysed using graphite electrodes.
(i) Describe the observations made at each electrode. (2 marks)
(ii) Write the equation for the reaction that took place at the anode. (1 mark)

(c) Some crystals formed in step (ii) were dissolved in water, and a portion of it reacted with potassium iodide solution. A yellow precipitate was formed. Write an ionic equation for this reaction. (1 mark)

(d) To another portion of the solution from (f), sodium hydroxide solution was added drop by drop until there was no further change. Describe the observation made. (1 mark)

(e) To a further portion of the solution from (f), a piece of zinc foil was added.
(i) Name the type of reaction taking place. (1 mark)
(ii) Write an ionic equation for the above reaction. (1 mark)

17. 2018 P1 Q 20.

(a) Zinc reacts with hydrochloric acid according to the following equation.

$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

Identify the reducing agent. Give a reason for the answer. (2 marks)

(b) Iron sheets are dipped in molten zinc to prevent rusting. Name this process. (1 mark)

18. 2018 P1 Q24.

(a) Name two ores of iron. (1 mark)
(b) Describe how the amount of iron in a sample of iron(III) oxide can be determined. (2 marks)

19. 2019 P1 Q3.

The flow chart in **Figure 1** represents some stages in the extraction of copper metal. Study it and answer the questions that follow.

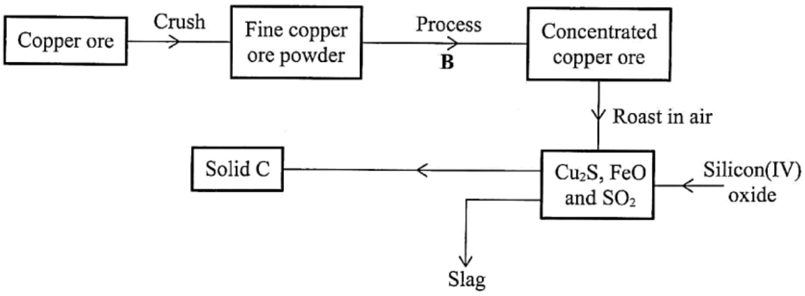


Figure 1

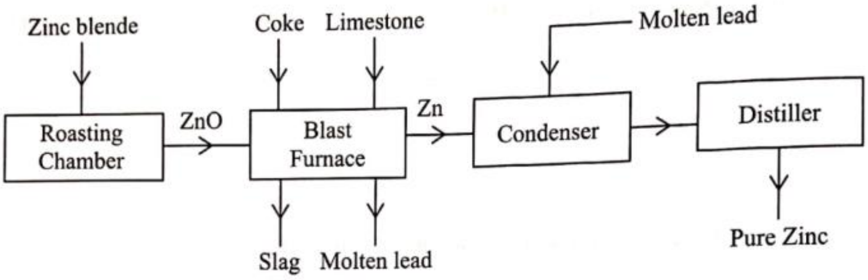
(a) Identify:
(i) The copper ore; (1 mark)
(ii) Process B; (½ mark)
(iii) Solid C. (½ mark)

(b) Write an equation for the reaction that forms the slag. (1 mark)

20. 2019 P2 Q2.

(a) Zinc occurs mainly as zinc blende. Name one other ore from which zinc can be extracted. (1 mark)

(b) The flow chart in **Figure 2** shows the various stages in the extraction of zinc metal. Study it and answer the questions that follow.



(i) Write an equation for the reaction which occurs in the roasting chamber. (1 mark)

(ii) Describe the process that takes place in the blast furnace. (3 marks)

(iii) Explain why molten lead is added to the condenser. (1 mark)

(iv) State two uses of zinc. (1 mark)

(v) Give one reason why the extraction of zinc causes pollution to the environment. (1 mark)

(b) Explain the observations made when zinc metal is added to hot sodium hydroxide. (2 marks)

TOPIC SIX

ORGANIC CHEMISTRY II

(ALKANOLS AND ALKANOIC ACIDS)

Table of Contents

Objectives

By the end of the topic, the learner should be able to:

(a) Name and draw the structures of simple alkanols and alkanoic acids.
(b) Describe the preparation and explain the physical and chemical properties of alkanols and alkanoic acids.
(c) State the main features of a homologous series.

(d) State and explain the uses of some alkanols and alkanoic acids.

(e) Describe the preparation, properties and uses of detergents and explain their effect on hard water.

(f) List some natural and synthetic polymers and fibres and state their uses.

(g) Describe the preparation, properties and uses of some synthetic polymers.

(h) Identify the structure of a polymer given the monomer.

(i) State the advantages and disadvantages of synthetic materials compared to those of natural polymers.

ORGANIC CHEMISTRY II

Alkanols (Alcohols)

Alkanols belong to a class of organic compounds which contain **carbon, hydrogen and oxygen**. Alkanols have a **hydroxyl group (OH–)** which is the **functional group** of the series.

Alkanols have a **general formula** $C_nH_{2n+1}OH$ where n = 1, 2, 3, 4...

Alkanols may be considered as **derivatives of water** in which **one of the hydrogen atoms in the water molecule is substituted by an alkyl group**. For example, methanol (CH_3OH) is obtained by replacing one hydrogen atom by a methyl ($-CH_3$) group. Ethanol is obtained by replacing one hydrogen atom in the water molecule by an ethyl ($-CH_2CH_3$) group.

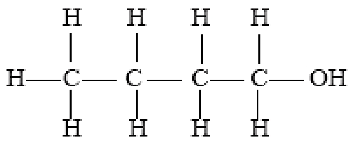
Nomenclature

Alkanols are named by **replacing the 'e'** of the **corresponding alkane** with the **suffix -ol,** for example:

Name of Alkane	*Name of corresponding alkanol*
Methane	Methanol
Ethane	Ethanol
Propane	Propanol
Butane	Butanol
Pentane	Pentanol

When naming alkanols, the following **rules** are used:

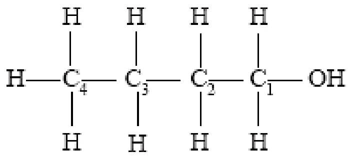
(i) Identify the **longest carbon chain containing the hydroxyl group (-OH)** which gives the **parent** name:



$$H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-OH$$

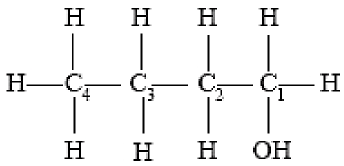
Butan–

(ii) **Number** the longest carbon chain such that the **carbon to which the hydroxyl group is attached has the lowest number possible.**

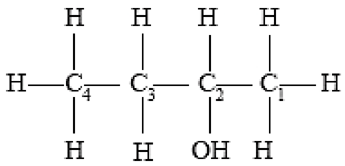


$$H-\underset{H}{\overset{H}{C_4}}-\underset{H}{\overset{H}{C_3}}-\underset{H}{\overset{H}{C_2}}-\underset{H}{\overset{H}{C_1}}-OH$$

(iii) Indicate the **position** of the hydroxyl group in the name, e.g.,

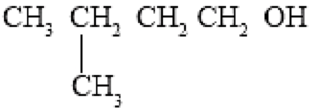


Butan –1– ol



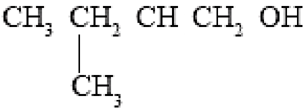
Butan – 2– ol

(iv) Locate the position of the **other substituent groups u**sing numbers that correspond to their position along the carbon chain.



$CH_3\ \underset{\displaystyle CH_3}{\underset{|}{CH_2}}\ CH_2\ CH_2\ OH$

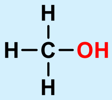
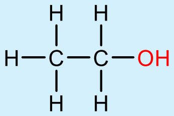
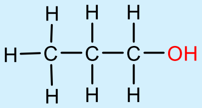
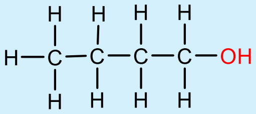
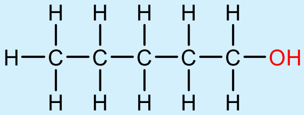
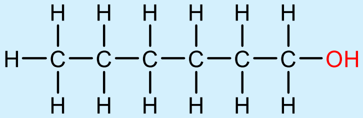
2-methyl butan-1-ol



$CH_3\ \underset{\displaystyle CH_3}{\underset{|}{CH_2}}\ CH\ CH_2\ OH$

3-methyl butan-1-ol

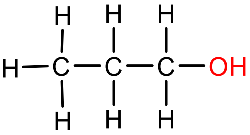
The first 5 alkanols

Alkanol	**Molecular formula**	**Structure**	**Condensed structural formula**
Methanol	CH_3OH		CH_3OH
Ethanol	C_2H_5OH		CH_3CH_2OH
Propanol	C_3H_7OH		$CH_3CH_2CH_2OH$
Butanol	C_4H_9OH		$CH_3CH_2CH_2CH_2OH$
Pentanol	$C_5H_{11}OH$		$CH_3CH_2CH_2CH_2CH_2OH$
Hexanol	$C_6H_{13}OH$		$CH_3CH_2CH_2CH_2CH_2CH_2OH$

Isomerism

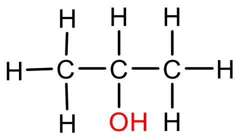
Alkanols exhibit two types of isomerism, **positional and branching isomerism**.

In **positional isomerism** the **position of the functional group (-OH) varies** within the carbon chain. For example: When the -OH group is attached to the first carbon atoms, the molecular structure can be represented as:



propan-**1**-ol

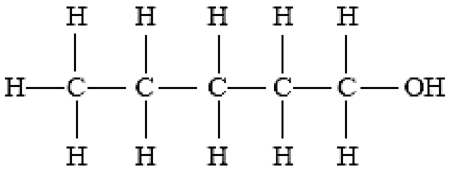
When the OH group is attached to the second carbon atom the molecular structure will be:



propan-2-ol

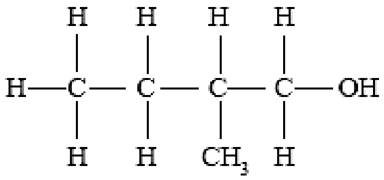
In **branching isomerism,** the **molecular formula of the compound remains the same**. However, there is **rearrangement of carbon atom**s such that one or more carbon atoms from the molecule form **alkyl groups attached to the longest carbon chain.** For example:

(i)



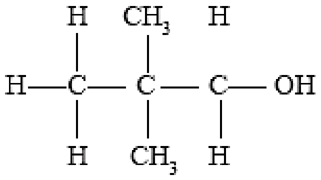
Pentan-1-ol

(ii)



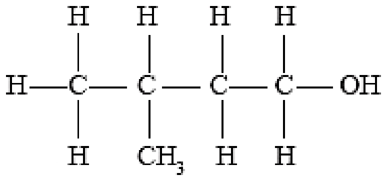
2- methylbutan-1-ol

(iii)



2, 2 - dimethylpropan-1-ol

(iv)



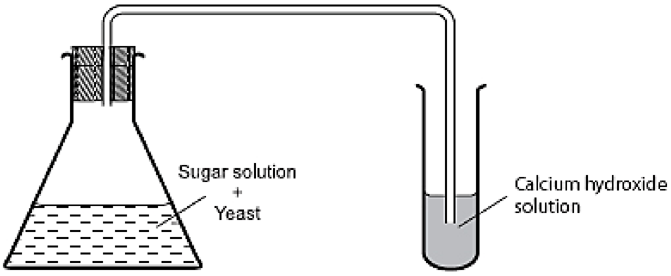
3- methylbutan-1-o1

In condensed structural formula of isomers, **constituents** may be **shown in brackets after the carbon atom to which they are attached.**

The condensed formula of 2-methylbutan-1-ol may be written as CH3CH(CH3) CH2CH2OH

Preparation of Alkanols

Ethanol can be prepared by the **decomposition of glucose molecules in the presence of enzymes.**



The sugar (C6H12O6) molecules are broken down into ethanol and carbon(IV) oxide by the enzymes in the yeast. Calcium hydroxide is used to test the presence of carbon(IV) oxide.

$$\text{Glucose} + \xrightarrow{yeast} \text{Ethanol} + \text{carbon(IV) oxide}$$

$$C_6H_{12}O_6(aq) \xrightarrow{yeast} 2C_2H_5OH(aq) + 2CO_2(aq)$$

This process is referred to as fermentation. Fermentation is the decomposition of an organic substance by micro-organisms to produce alcohol, carbon(IV) oxide and heat.

However, a very small amount of alcohol is usually produced by fermentation (about 10% ethanol by volume). The alcohol content can be **increased by fractional distillation of the crude solution.**

The ethanol obtained contains 5% water. The water can be removed by using a suitable drying agent such as calcium oxide, to obtain absolute ethanol.

Alternatively, ethene may be hydrolysed using concentrated sulphuric(VI) acid.

$\mathrm{CH2H4(g) + H2SO4(aq) \longrightarrow C2H5OSO3H(l)}$

When water is added to the mixture, the ethylhydrogen sulphate is hydrolysed to ethanol.

$\mathrm{CH3CH2OSO3H(g) + H2O(l) \longrightarrow CH3CH2OH(l) + H2SO4(aq)}$

The mixture of ethanol and the acid is separated by distillation because of the difference in boiling points.

On a large scale, ethanol is manufactured by **reacting steam and ethene** at a temperature of 300°C and a pressure of about 60–70 atmospheres over phosphoric(V) acid. The acid is used as a catalyst.

The ethene for this reaction is obtained from **cracking of large alkanes**.

Properties of Alkanols

Physical Properties of Alkanols.

- Ethanol is a colourless liquid with a characteristic odour. It has a **melting point of –114°C** and **boiling point of 78°C.**

  These values are **high when compared with those of alkane**s having corresponding molecular masses. This is due to the **presence of hydrogen bonds in addition to van der Waals forces.**
- Ethanol is **highly soluble in water.** The high solubility of ethanol in water is because the **ethanol molecules are polar** like those of water. The ethanol molecules are therefore able to f**orm hydrogen bonds with the water molecules.**

Name	**Molecular mass**	**Molecular formula**	**Boiling point (°C)**	**Melting point (°C)**	**Solubility g/100 g water**
Methanol	32	CH_4O	65	-97.5	Highly soluble
Ethanol	46	C_2H_6O	78	-114	Highly soluble
Propan-1-ol	60	C_3H_8O	97	-126	Highly soluble
Butan-1-ol	74	$C_4H_{10}O$	117	-90	8
Pentan-1-ol	88	$C_5H_{12}O$	138	-79	2.7
Hexan-1-ol	102	$C_6H_{14}O$	157	-52	0.6

- Alkanols are soluble in water but their **solubility decreases gradually as the molecular mass increases**. For example, methanol is more soluble than hexan-1-ol.
- The **melting points and boiling points** of alkanols **increase down** the series due to the **increases in the strength of intermolecular forces of attraction.** Alkanols have higher melting and boiling points than their corresponding alkanes with the same molecular mass.

  This is **because of the hydrogen bonds between alkanol molecules, caused by the functional group (-OH). Hydrogen bonds are stronger than van der Waals' forces.**
- The melting and boiling points **increase with increase in molar mass.** This is attributed to the **increase in the strength of inter-molecular forces.**

Chemical Properties of Alkanols

In solution, ethanol has a **pH slightly below 7**. This is because in solution it behaves as a **weak acid.**

Combustion

Ethanol burns readily in air with a **pale blue flame** to produce carbon(IV) oxide and water. This is because **ethanol is saturated and undergoes complete combustion**.

$C_2H_5OH(l) + 2O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$

Reactions with metals

Ethanol **reacts with sodium liberating hydrogen gas and a salt, sodium ethoxide.** The sodium ethoxide **hydrolyses** in water to generate -OH which makes the solution **alkaline**.

$2CH_3CH_2OH\ (l) + 2Na\ (s) \longrightarrow 2CH_2CH_2ONa\ (aq) + H_2\ (g)$

Sodium metal reacts with any other alkanol to liberate a salt and hydrogen gas, e.g.,

$2ROH(l) + 2Na(s) \longrightarrow 2RONa(s) + H_2(g)$

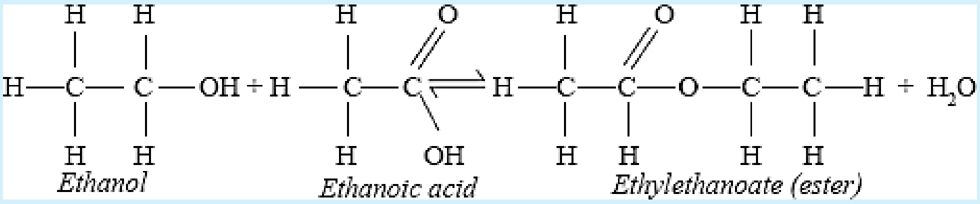
Esterification

Ethanol **reacts** with **ethanoic acid** in the presence of a **few drops of concentrated sulphuric(VI) acid** to form a substance with a **pleasant smell.** The product formed is known as **ethylethanoate** which belongs to a group of compounds known as **esters**. The process of ester formation is known as esterification.

$$C_2H_5OH(l) + CH_3COOH(aq) \underset{}{\overset{\text{Conc } H_2SO_4}{\rightleftharpoons}} CH_3COOC_2H_5\ (l) + H_2O(l)$$

Ethanol *Ethanoic acid* *Ethylethanoate (ester)*

Below is the equation showing the structural formulae of the reactants and products.



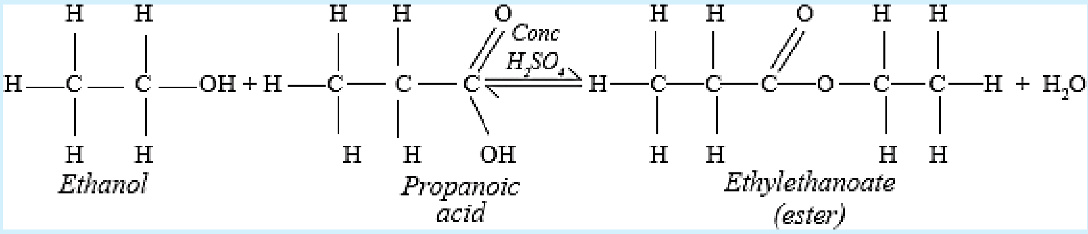
The **alkyl** part of the **ester** is derived from the **alkanol**, while the **alkanoate** part is derived from the **acid**. The **alkyl** group from the alkanol **attaches itself to the carboxylic acid** thereby **displacing a hydrogen atom.**

Ethanol also reacts with **propanoic** acid to form an ester known as **ethylpropanaote**.

$$C_2H_5OH + CH_3CH_2COOH(aq) \overset{\text{Conc } H_2SO_4}{\rightleftharpoons} CH_3CH_2COOC_2H_5(l) + H_2O(l)$$

Ethanol *Propanoic acid* *Ethyl propanoate (ester)*

Below is the equation showing the structural formulae of the reactants and products.



Under ordinary conditions the reaction takes place slowly. Therefore, concentrated sulphuric(VI) acid is added to **catalyse** the reaction.

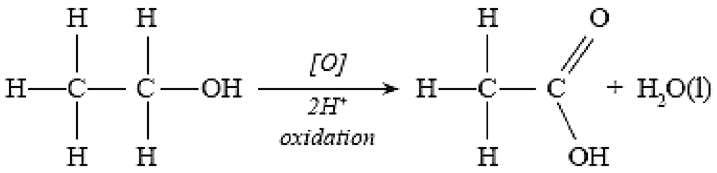
Oxidation by oxidising agents ($KMnO_4$ and $K_2Cr_2O_7$)

Ethanol is **oxidised by oxidising agents such as potassium manganate(VII)** and **potassium dichromate(VI) to form ethanoic acid**. When reacted with ethanol, **acidified purple potassium manganate(VII) turns colourless.** This is due to the **reduction of manganate ions (MnO_4^-) which are purple to manganese ions (Mn^{2+}) which are colourless.**

Acidified orange potassium dichromate(VI) **turns green** due to the reduction of chromate ions ($Cr_2O_7^{2-}$) to chromium ions (Cr^{3+}).

$$\underset{Ethanol}{CH_3CH_2OH(l)} \xrightarrow[2H^+]{MnO_4^-} \underset{Ethanoic\ acid}{CH_3COOH(aq)} + H_2O(l)$$

The structure of the reactant and product in the equation can be represented as:

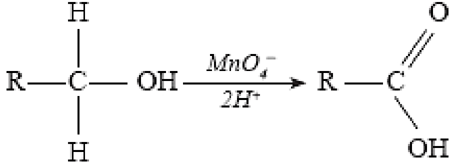


$$H-\overset{\overset{H}{|}}{\underset{\underset{H}{|}}{C}}-\overset{\overset{H}{|}}{\underset{\underset{H}{|}}{C}}-OH \xrightarrow[\substack{2H^+\\ oxidation}]{[O]} H-\overset{\overset{H}{|}}{\underset{\underset{H}{|}}{C}}-C\overset{\nearrow O}{\underset{\searrow OH}{}} + H_2O(l)$$

The oxidising property of the chromate (VI) is used in some **breathalysers** to indicate the level of **alcohol (alkanol) content** in the breath of motorists.

Generally, alkanols are **oxidised by oxidising agents to alkanoic (carboxylic) acids.**

$$\text{Alkanols} \xrightarrow[2H^+]{MnO_4^-} \text{Alkanoic acids}$$



$$R-\overset{\overset{H}{|}}{\underset{\underset{H}{|}}{C}}-OH \xrightarrow[2H^+]{MnO_4^-} R-C\overset{\nearrow O}{\underset{\searrow OH}{}}$$

R- represents an alkyl group

Dehydration by concentrated sulphuric (VI) acid.

Concentrated sulphuric(VI) acid reacts with ethanol at **180**°C to form **ethene and water**. The concentrated sulphuric(VI) acid acts as a **dehydrating** agent.

$$\underset{Ethanol}{2CH_3CH_2OH(l)} \xrightarrow[180°C]{Conc\ H_2SO_4} \underset{Ethene}{CH_2CH_2(g)} + H_2O(l))$$

At a temperature **of 140°C, incomplete dehydration occurs in which an ether** is formed.

$$\underset{Ethanol}{2CH_3CH_2OH(l)} \xrightarrow[140°C]{Conc\ H_2SO_4} \underset{Ether}{C_2H_5OC_2H_5(l)}$$

Uses of Alkanols

1. As solvents, i.e., in the preparation of drugs.
2. As fuels when blended with gasoline to form gasohol.
3. In the manufacture of synthetic fibres, e.g., polyvinylchloride and polythene.
4. As an antiseptic when used under specified concentrations.

5. Ethanol is used as an alcoholic drink only in low concentrations.

Alcohol and Health

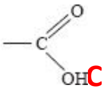
Over-consumption of ethanol causes damage to some body organs such as the liver, the brain and the heart. It also leads to addiction.

When small amounts of methanol are added to ethanol, it causes blindness and may even lead to death. In many parts of the world it is illegal to sell ethanol to persons under the age of eighteen.

Methylated spirit is an alcohol containing about 95% absolute alcohol and 5% methanol. Methanol and a purple dye called methylviolet are added to ethanol to make the methylated spirit unsuitable for human consumption.

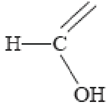
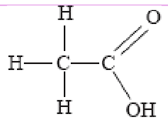
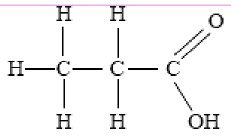
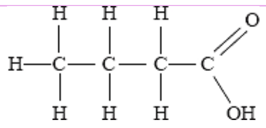
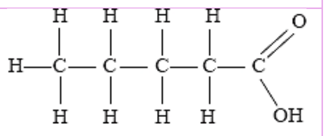
Alkanoic Acids (Carboxylic Acids)

Alkanoic acids belong to a homologous series of organic compounds that contains a **carboxyl** group **(-COOH)** as a functional group.

A carboxyl group has the structural formula 

Alkanoic acids have a general formula **CnH2n+1COOH** where n =0,1,2,3…

The first five alkanoic acids

Alkanoic acid	*Molecular formula*	*Structural formula*
Methanoic	HCOOH	
Ethanoic	CH_3COOH	
Propanoic	C_2H_5COOH	
Butanoic	C_3H_7COOH	
Pentanoic	C_4H_9COOH	

Alkanoic acids are naturally found in **fruits** such as oranges, lemon and pepper. Methanoic acid is present in **nettle leaves** and **insect stings** such as bees and wasps. Ethanoic acid is commonly known as **vinegar**. **Butanoic** acid is found in beef fat (**Butter**). **Hexandioic** acid is found in palm oil and olive oil.

Nomenclature

Alkanoic acids are named by **replacing the "e"** ending of the **corresponding alkane** by the **suffix –oic.** The carbon atom to which the functional group is attached is given position one.

The simplest member of the alkanoic acid series when n = 0 is **$HCOOH$ (Methanoic acid)** and when n = 1 is **CH_2COOH (ethanoic acid).**

Names and formulae of the first ten alkanoic acids

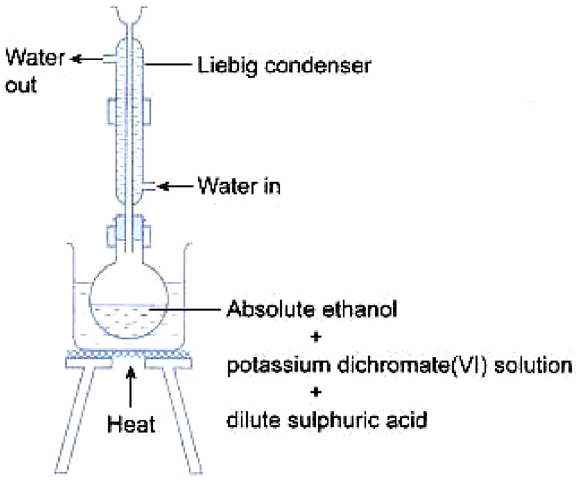
Name of acid	*Molecular formula*	*Condensed structural formula*
Methanoic	CH_2O_2	$HCOOH$
Ethanoic	$C_2H_4O_2$	CH_3COOH
Propanoic	$C_3H_6O_2$	CH_3CH_2COOH
Butanoic	$C_4H_8O_2$	$CH_3CH_2CH_2COOH$
Pentanoic	$C_5H_{10}O_2$	$CH_3CH_2CH_2CH_2COOH$
Hexanoic	$C_6H_{12}O_2$	$CH_3CH_2CH_2CH_2CH_2COOH$
Heptanoic	$C_7H_{14}O_2$	$CH_3CH_2CH_2CH_2CH_2CH_2COOH$
Octanoic	$C_8H_{16}O_2$	$CH_3CH_2CH_2CH_2\ CH_2CH_2CH_2COOH$
Nonanoic	$C_9H_{18}O_2$	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2COOH$
Decanoic	$C_{10}H_{20}O_2$	**$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2COOH$**

Laboratory Preparation of Ethanoic Acids

In the laboratory, ethanoic acid can be prepared by oxidising ethanol using suitable oxidising agents such acidified potassium manganate (VII) or potassium dichromate (VII).

On heating, acidified potassium manganate(VII) oxidises ethanol to ethanoic acid.

$$CH_3CH_2OH \xrightarrow[2H^+]{MnO_4} CH_3COOH(aq) + H_2O(l)$$

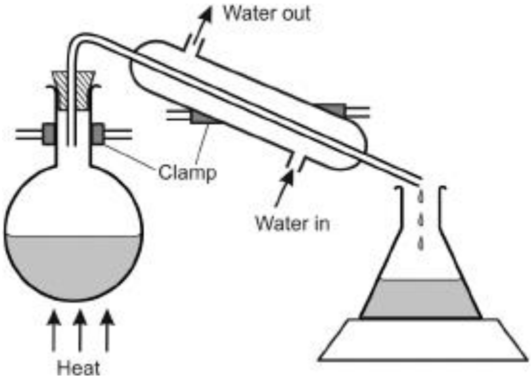


OBSERVATIONS

- During the reaction the **purple solution turns colourless.** The colour of the solution is colourless because the purple manganate(VII) (MnO^{-4}) ions are reduced to colourless manganese(II) (Mn^{2+}) ions.

- When acidified **potassium chromate(VI) is used,** the solution in the flask is **orange before** heating but **after heating it turns green.** The orange colour is due to chromate(VI) ions which are reduced to green chromium(II) ions.

$$C_2H_4(g) + H_2O(g) \xrightarrow{\text{Phosphoric(V) acid}} CH_3CH_2OH(l)$$



To **ensure complete oxidation, excess oxidising agent is used.** The **condenser** ensures any **vapour escaping is condensed back into the flask f**or further reaction.

The solution in the flask contains the oxidising agent, water and ethanoic acid. In order to obtain **pure** ethanoic acid the **mixture is distilled**. The distillate collected at about 118 °C is ethanoic acid which is a **colourles liquid with a sharp smell**.

Properties of Alkanoic acids

Physical properties of alkanoic acids

Name of acid	Molecular mass	State at room temperature	Solubility	Melting point (°C))	Boiling point (°C
Methanoic	46	Liquid	Very high	8.4	101
Ethanoic	70	Liquid	Very high	16.6	118
Propanoic	84	Liquid	Highly soluble	-20.8	141
Butanoic	98	Liquid	Highly soluble	-6.5	164
Pentanoic	112	Liquid	Moderate	-34.5	186
Hexanoic	116	Liquid	Low	-1.5	205

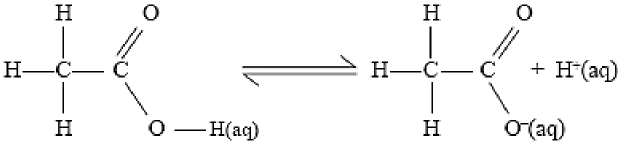
- Alkanoic acids are **soluble** in water because their **molecules are able to form hydrogen bonds with water molecules.** Their **solubility decreases with increase in molecular mass** because of the **decreases** in **polarity** of the acid molecules.
- Melting and boiling points **increase** with **increase in molecular mass.** This is due the **increase** in the **strength of van der Waals' forces.**
- The alkanoic acids have **higher melting and boiling points than their corresponding alkanols** with the same molecular mass **because the alkanoic acids form more hydrogen bonds per molecule than the corresponding alkanol.**

Chemical Properties of alkanoic acids

Acidic properties in solution.

Ethanoic acid has a pH of **about 5** and is a **weak acid because it partially dissociates in solution** to form ethanoate ions and hydrogen ions. The dissociation can be represented as:

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$



Reaction with metals.

Ethanoic acid reacts with **metals** to form **hydrogen gas and a salt** (**organic** salt).

$$2CH_3COOH(aq) + Mg(s) \longrightarrow (CH_3COO)_2Mg + H_2(g)$$

Ethanoic acid *Magnesium* *Magnesium ethanoate*

Alkanoic acids react with metals to form **alkanoate** and hydrogen gas.

Alkanoic acid + Metal ⟶ Metal Alkanoate + Hydrogen gas

Reaction with bases and alkalis (neutralization)

Ethanoic acid neutralises sodium hydroxide forming a salt and water.

$$CH_3COOH(aq) + NaOH(aq) \longrightarrow CH_3COONa(aq) + H_2O(l)$$

Ethanoic acid *Sodium hydroxide* *Sodium ethanoate*

Generally, alkanoic acids react with alkalis to form a salt and water.

Alkanoic acid + Alkali ⟶ Metal alkanoate + water

Reaction with carbonates

Ethanoic acid also reacts with sodium carbonate to produce a salt, carbon(IV) oxide and water.

$$2CH_3COOH(aq) + Na_2CO_3(aq) \longrightarrow 2CH_3COONa(aq) + CO_2(g) + H_2O(g)$$

OBSERVATION

Effervescence due to production of carbon (IV) oxide

Reaction with alkanols (esterification)

When ethanoic acid reacts with ethanol, in the presence of a few drops of concentrated sulphuric(VI) acid, a pleasant smelling compound called an ester is formed.

$$CH_3COOH(aq) + CH_3CH_2OH(aq) \longrightarrow CH_3COOCH_2CH_3(aq) + H_2O(l)$$

Ethanoic acid *Ethanol* *Ethylethanoate*

Uses of Alkanoic Acids

1. Used as solvents.
2. In the manufacture of drugs and chemicals.
3. In flavouring of foods, e.g., ethanoic acid (vinegar).
4. In the manufacture of synthetic fibres such as terylene, dacron and nylon.
5. In preservation of food, e.g., benzoic acid.

Detergents

Detergents are **substances which improve the cleaning properties of water**. There are two types; **soapy and soapless** detergents.

Soapy Detergents

Soapy detergents are referred to as **soap**. They are **prepared from either fats or oils**.

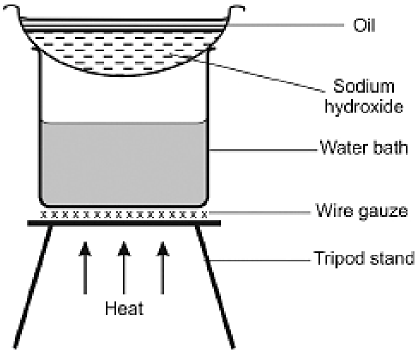
Soap is a **sodium or potassium salt** with a **general formula** $C_nH_{2n+1}COO^-Na^+$ or $C_nH_{2n+1}COO^-K^+$

Fats and oils are **esters**. **Fats** occur naturally in **animals** while **oils** occur both in plants and **animals**. Some examples of oils include **whale oil, groundnut oil, corn oil and coconut oil.** Naturally occurring fats are **butter** from milk, **lard** from **pigs** and **tallow** from **animals**.

Fats are **saturated organic compounds** while **oils are unsaturated.**

Preparation of soap.

In the laboratory, soap can be prepared by mixing 5 cm^3 of castor oil and 20 cm^3 of 4 M sodium hydroxide in an evaporating dish. The mixture is then heated in a water bath for about 20 minutes, stirring continuously while adding small amounts of distilled water. Finally, three spatulafuls of sodium chloride are added, the mixture is stirred and allowed to cool. The mixture is then filtered and the residue washed with cold distilled water.



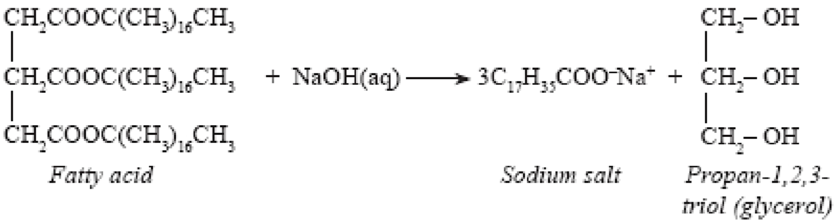
When an alkali is boiled with fat or oil, a **hydrolysis reaction** takes place.

Fat + NaOH ⟶ Soap + glycerol

When a fatty acid is **hydrolysed in the presence of an alkali** the process is referred to as <u>saponification</u>.

In the hydrolysis of a fatty acid, **sodium hydroxide neutralises the acid to form the sodium salt of the acid.**

For example;



$$\begin{array}{l} CH_2COOC(CH_3)_{16}CH_3 \\ | \\ CH_2COOC(CH_3)_{16}CH_3 \\ | \\ CH_2COOC(CH_3)_{16}CH_3 \end{array} + NaOH(aq) \longrightarrow 3C_{17}H_{35}COO^-Na^+ + \begin{array}{l} CH_2-OH \\ | \\ CH_2-OH \\ | \\ CH_2-OH \end{array}$$

The soap formed (sodium octadecanoate) is commonly known as **sodium stearate**.

- **Sodium chloride crystals are added to the mixture to reduce the solubility of the soap in glycerol.** This is known as <u>salting out</u>.
- The soap obtained is **rinsed in distilled water to remove impurities** such as **glycerol**, other **salts** and **unused alkali solution.**

The Mode of Action of Soap in Cleaning

The cleaning property of soap depends on its structure. For example, sodium stearate consists of an oil soluble long hydrocarbon end and a polar end which is water soluble ($C_{11}H_{35}COO^-Na^+$)

A soap molecule has:

(i) A hydrocarbon chain end which is non-polar and has no attraction for water (water – hating) and is called <u>hydrophobic.</u>

(ii) A carboxylate end which is polar and is attracted to water or is 'water-loving' and is called <u>hydrophilic</u>. This end is in fact negatively charged in water.

$$\underbrace{CH_3(CH_2)_{16}}_{\text{non polar}}\quad\underbrace{COO^-Na^+}_{\text{polar}}$$

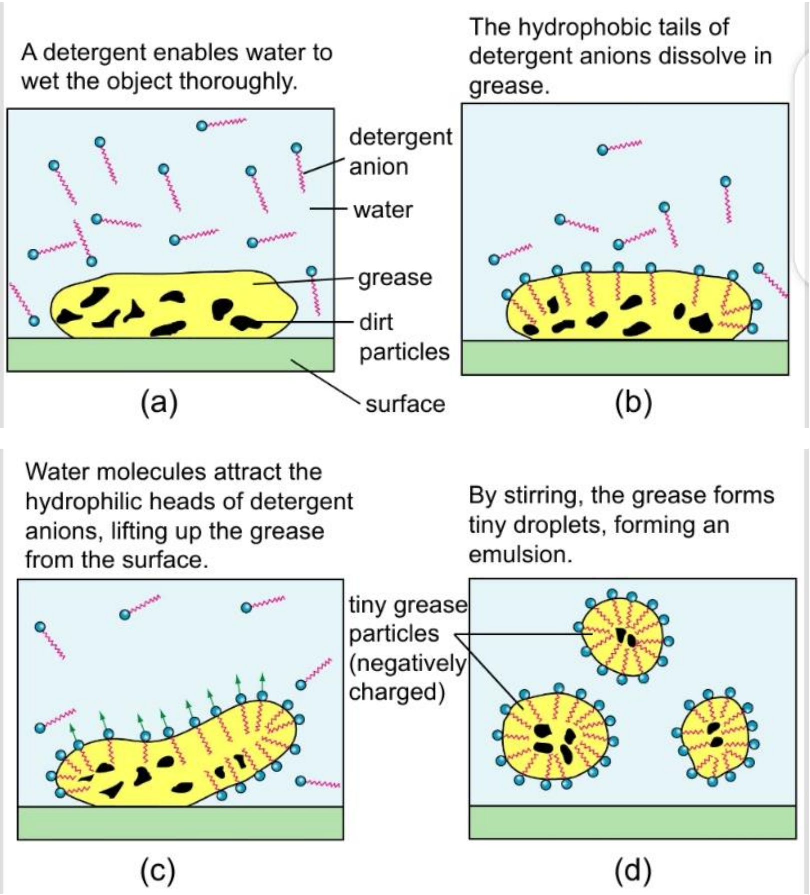
The above can also be represented by a skeletal structure as shown.

The hydrocarbon chain may also be represented by 'R'. Thus **R – COO–Na**+

In water, soap **dissociates** into **carboxylate** and **sodium ions,** $CH_3(CH_2)_{16}COO^-$ and Na^+. The **non-polar end of the carboxylate ion is oil soluble** while the **polar end** is **water soluble.**

Water alone is not sufficient to clean grease stained linen or surfaces because **water cannot dissolve grease as the two are immiscible.**

When soap is **added** to water during washing, the **non-polar hydrophobic end of the carboxylate ions lodge themselves onto the grease** while the **hydrophilic ends stick out in the water**, since they are attracted by water molecules.


A detergent enables water to wet the object thoroughly.

The hydrophobic tails of detergent anions dissolve in grease.

Water molecules attract the hydrophilic heads of detergent anions, lifting up the grease from the surface.

By stirring, the grease forms tiny droplets, forming an emulsion.

Agitation of the material being washed **causes enough of the carboxylate ions** to stick into grease **particles and dislodge them**, this forms **small droplets** with the **water loving ends** poking out (**micelles**). The micelles **cannot coalesce together as they repel one another due to the charge on their surface.** This way they are **washed away by water** when the garment is **rinsed**.

The detergents **'heads'** are attracted by water molecules.

Effect of Hard Water on Soap

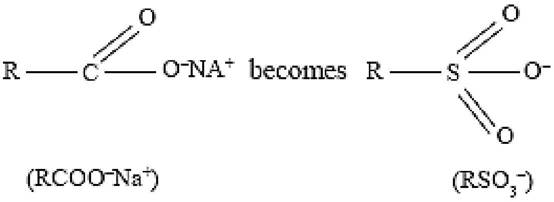
When soap is added to water containing calcium and magnesium ions, a precipitation reaction takes place. For example:

$$2C_{17}H_{35}COO^-Na^+(aq) + Ca^{2+}(aq) \longrightarrow (C_{17}H_{35}COO^-)_2Ca^{2+}(s) + 2Na^+(aq)$$

Lather **does not form until all the calcium or magnesium ions responsible for the hardness have been precipitated.** The precipitate floats on the water as **scum**.

Soapless Detergents

These are detergents in which a **carboxylic group** of the soap is **replaced by an alkyl sulphonate group**. For example:



$$R-C(=O)-O^-NA^+ \text{ becomes } R-S(=O)_2-O^-$$

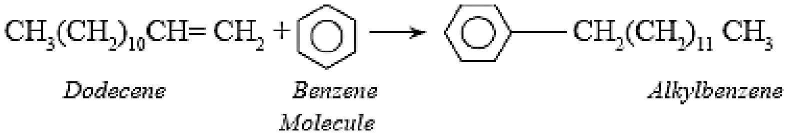
$(RCOO^-Na^+)$ (RSO_3^-)

Where **R represents** the **long hydrocarbon chain**.

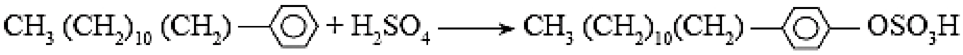
Soapless detergents act like soap in the cleaning process **but they are not affected by hard water unlike ordinary soap.** Therefore, unlike soap, soapless detergents **lather readily with water** since the **corresponding calcium and magnesium salts are soluble.**

Soapless detergents are manufactured **from petroleum products.** They can also be made from **vegetable oil or fats** which **do not contain the benzene ring.**

The initial step in the preparation involves the **heating of the long chain hydrocarbons** with **benzene** molecules at **very high temperature** to produce **alkylbenzene** as the organic part of the detergent.

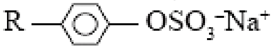


The alkylbenzene is **further reacted with concentrated sulphuric(VI) acid** to form **alkylbenzene sulphonate.**

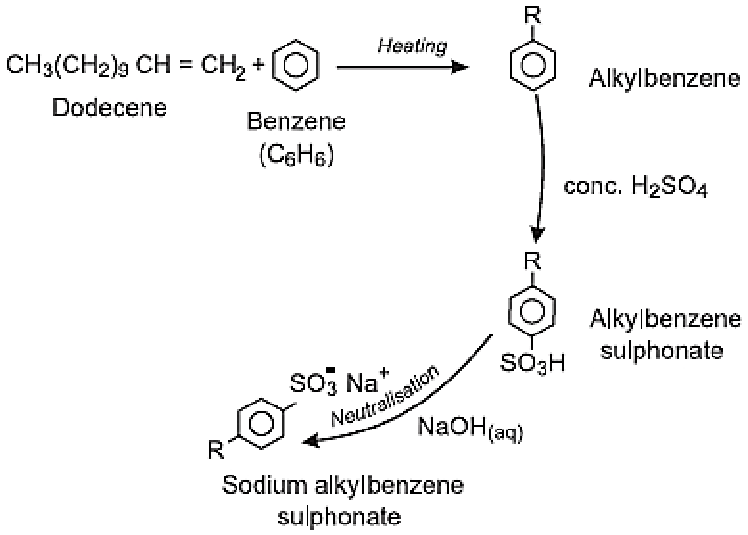


The **alkylbenzene sulphonate** is **reacted** with sodium hydroxide solution to form **sodium alkylbenzene sulphonate** which is the detergent.

The long hydrocarbon chain can be **presented by R.** Thus, the **sodium alkylbenzene sulphonate may be written as:**



The flow diagram below is a summary of the process involved during preparation of soapless detergents.



Soapless detergents are manufactured in **liquid or solid form**.

In order **to improve the cleaning properties of soapy and soapless detergents, tetraoxophosphate materials are added**. The compounds **prevent formation** of compounds with **Ca2+, Mg2+** ions hence **no scum is formed**.

Examples of soapless detergents sold locally are: Omo, Dynamo, Perfix, Persil, Sunlight and Toss.

Comparison between soapless detergents and soap

Soapless detergents	***Soapy detergents***

Very soluble in water and therefore can be used well in hard water.	Form scum in hard water
Cause water pollution because some are non-biodegradable.	Do not cause pollution because they are biodegradable.
Are expensive.	Are cheap.

Pollution Effect of Detergents

Detergents contain **long chains of alkylbenzene** groups which are **difficult to break down through bacterial action**. When large quantities of detergents are released into lakes, rivers and dams, froth forms. Eventually, the forth forms a **protective blanket layer on top of the water preventing air from dissolving in the water**. Lack of oxygen in water causes death of animals due to high biological and chemical oxygen demand.

The **tetraoxophosphate** and other compounds added to remove Ca**2+** and Mg**2+** during cleaning **provides nutrients for aquatic plants**, e.g., algae and water hyacinth. The population of the aquatic plants grows quickly depleting the dissolved oxygen in water. This is referred to as eutrophication.

In order to control pollution by detergents, efforts are being made by manufacturers to use biodegradable materials in detergents, i.e., more unbranched long chain hydrocarbons.

Polymers

A **polymer** is a long chain organic molecule formed when a small chain molecule undergoes self addition reaction.

The unit molecule is referred to as a monomer. Monomers may be of the same molecule or different compounds.

The process through which the monomers combine to form a long chain molecule is known as Polymerisation.

The polymers may be man-made (synthetic) or naturally occurring. **Naturally occurring polymers include rubber, cellulose, wool, silk and starch** while **synthetic polymers include nylon, terylene and polyethene.**

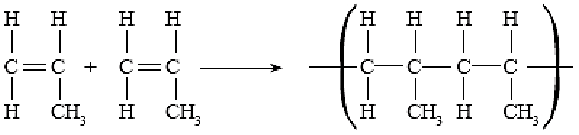
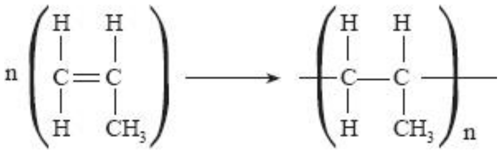
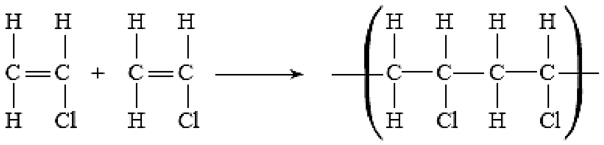
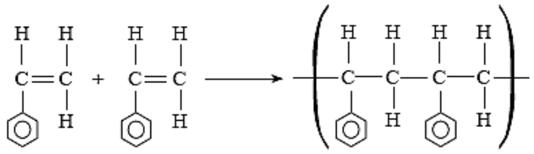
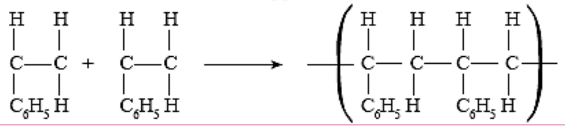
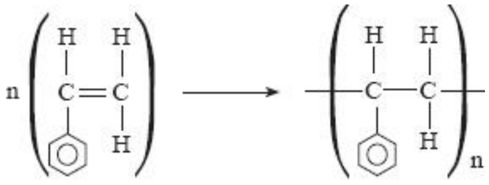
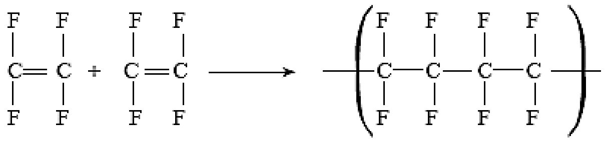
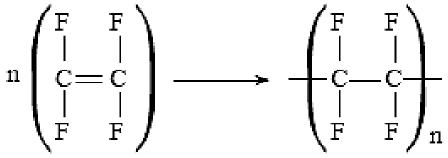
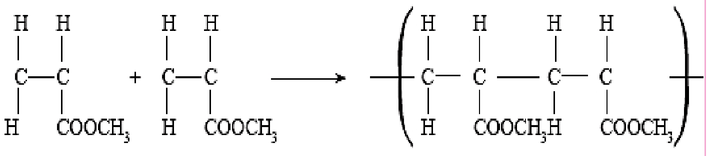
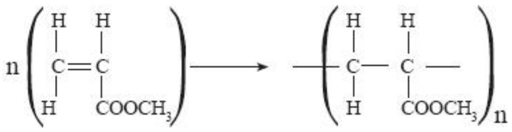
Synthetic polymers are made using two methods; **addition polymerisation** and **condensation polymerisation.**

Addition Polymerisation

Addition polymerisation occurs when **unsaturated molecules (monomers)** join to form long chain molecules (polymers) **without the formation of any other product.** Monomers open and bond with each other.

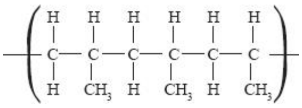
Polymers formed through addition polymerisation are: polythene, polypropene, polyvinylchloride (PVC), polystyrene, polytetrafluoroethene and perspex.

The following equations show the formation of these polymers:

Polymerization process	General structure
(a) Polypropene  *Propene (monomer)* *Propene (monomer)* *Polypropene (polymer)*	
(b) Polychloroethene (Polyvinylchoride)  *Chloroethene* *Chloroethene* *Polychloroethene*	
(c) Polyphenylethene (polystyrene)  *Phenylethene* *Phenylethene* *Polyphenylethene (polystyrene)* Or 	
(d) Polytetrafluoroethene  *Tetrafluoroethene* *Polytetrafluoroethene (PTFE)*	
(e) Polymethylmetacrylate (PMMA)  ***2-methyl propanoate*** ***polymethylmetacrylate*** ***(PMMA)*** ***(methyl metacrylate)***	

A **polymer formed by addition of identical molecules** will have a **relative molecular mass that is an integral multiple of the relative molecular mass of that monomer.** If the relative molecular mass of a polymer and part of its structure is given, the number of monomers forming the polymer can be determined.

For example:

A polymer has the following  structure and has a molecular mass of 4200.

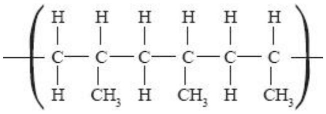
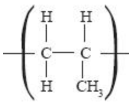
(i) Identify the monomer and draw its structure.

(ii) Determine the relative molecular mass of the monomer.

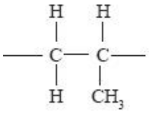
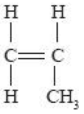
(iii) Determine the number of monomers in the polymer.

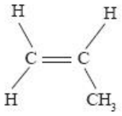
Solution

(a) The monomer is determined by determining the repeating units in the polymer given:

Thus in the polymer: ,  **the repeating unit is:** 

Since the monomer must be an unsaturated hydrocarbon, then the monomer forms the repeating unit.

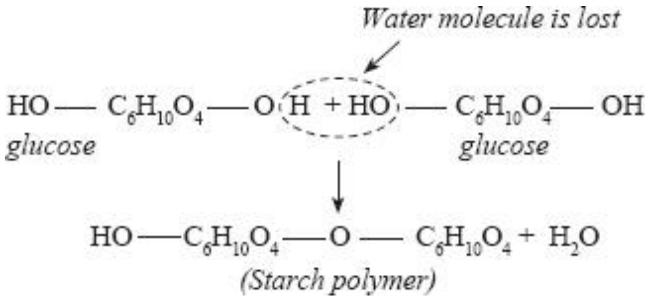
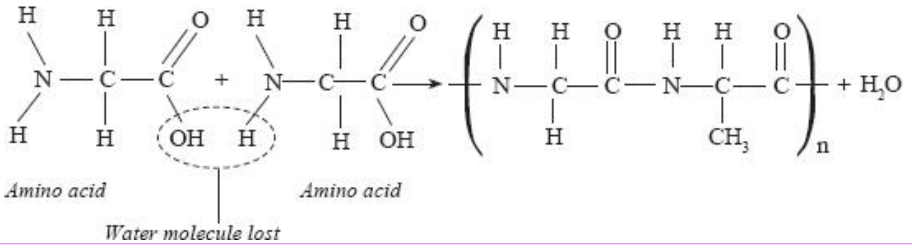
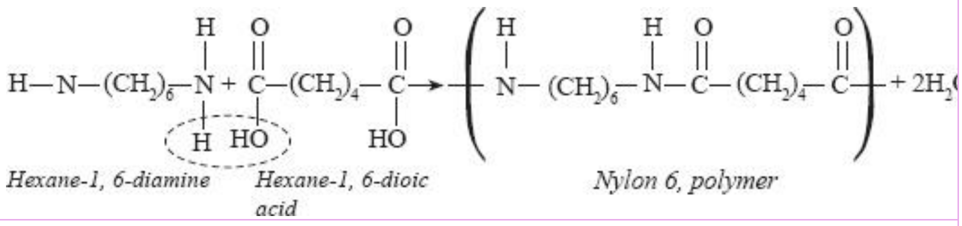
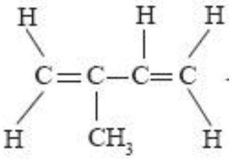
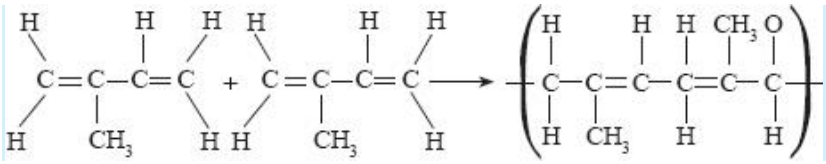
Thus:  is 

(ii) The relative molecular mass of the monomer is: ⇒  $(12 \times 3) + (1 \times 6) = 42$

(iii) Number of monomers = $\frac{\text{Molecular mass of polymer}}{\text{Molecular mass of monomer}} = \frac{4200}{42} = 100$

Condensation Polymerisation

In **condensation polymerisation, identical or different monomers** combine to form **long chain molecules with the loss of small molecules like water.** The monomer **should have two functional groups at each end** so that molecules can join at both ends to form long chain molecules. The functional groups may be **identical or different**. For example:

Name	Polymerization process
(a) Starch	
(b) Protein	
(c) Nylon 6.6	**The polymer is formed from the reaction between hexane -1, 6. diamine and hexane -1, 6 -dioic acid.** 
Rubber Naturally occurring rubber is obtained from the rubber tree.  *isoprene*	**Rubber trees produce a liquid called latex which is collected from cuts made on the trunks of the rubber trees. The latex is then allowed to coagulate resulting in a hydrocarbon polymer which is made up of isoprene (2- methyl but -1, 3-diene).** The monomer polymerises as shown below.  Synthetic rubber can also be obtained by **polymerisation of isoprene**. The product is usually soft hence must be hardened by a process called vulcanisation. This involves **heating the rubber with sulphur.** The sulphur atoms form links between chains of rubber molecules reducing the number

	of double bonds in the polymers. This makes the material **tougher, less flexible and less soft. This improves the quality of rubber.**

Advantages of Synthetic Polymers and Fibres

Synthetic polymers have many advantages over natural materials:

1. They are less affected by acids, alkalis, water and air.
2. They are lighter.
3. They are stronger.
4. They can be moulded into desired shapes easily.
5. They are less expensive.

Disadvantages of Synthetic Fibres

1. Some synthetic fibres burn more readily than natural ores.
2. They do not decompose easily, i.e., are non-biodegradable. This results in environmental pollution.
3. Some synthetic polymers give off poisonous gases when they burn, e.g., polythene gives off hydrogen cyanide and carbon(IV) oxide.

Some polymers and their uses

Polymer	***Uses***
Polyethene	Film wrappers, flexible bottles, electrical wire insulators, water pipes.
Polypropene	Crates, carpets, plastic bottles, chairs, ropes.
Polychloroethene	Floor tiles, car dash boards, cool water pipes, hose pipes, gutters.
Polyphenyethene	Ceiling lines, insulation materials.
Polytetrafluoroethane	Non-stick coating on pressing boxes and cooking utensils.
Perspex	Substitute for glass.

Nylon	Substitute for cotton in the textile industry, ropes carpets, brushes, fishing nets, parachutes.
Rubber	Vehicle tyres.

Review Exercises

1. 2006 Q7 P1

A group of compounds called chlorofluorocarbons have a wide range of uses but they also have harmful effects on the environment. State one:

(a) Use of chlorofluorocarbons. (1 mark)

(b) Harmful effect of chlorofluorocarbons on the environments. (1 mark)

2. 2006 Q5 (P2)

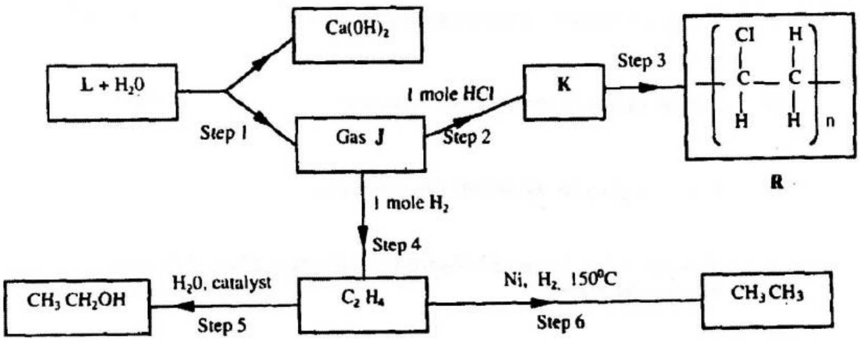
(a) What name is given to a compound that contains carbon and hydrogen only? (½ mark)

(b) Hexane is a compound containing carbon and hydrogen.

(i) What method is used to obtain hexane from crude oil? (1mark)

(ii) State one use of hexane (1mark)

(a) Study the flow chart below and answer the questions that follow.



(i) Identify reagent **L**. (1mark)

(ii) Name the catalyst used in step **5.** (1mark)

(iii) Draw the structural formula of gas **J**. (1mark)

(iv) What name is given to the process that takes place in step **5**? (½ mark)

(v) What name is given to the process that takes place in step 5? (½ mark)

(b) (i)write the equation for the reaction between aqueous sodium hydroxide and aqueous ethanoic acid. (1 mark)

(ii) Explain why the reaction between 1g of sodium carbonate and 2M hydrochloric acid is faster than the reaction between 1g of sodium carbonate and 2M ethanoic acid. (2 marks)

3. 2007 Q20 P1

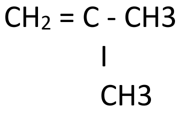
An alkanol has the following composition by mass: hydrogen 13.5%, oxygen 21.6% and carbon 64.9%

(a) Determine the empirical formula of the alcohol. (C=12.0; H=1.0, O=16.0). (2marks)

(b) Given that the empirical formula and molecular formula of the alkanol are the same, draw the structure of the alkanol. (1 mark)

4. 2007 Q2 P2

(a) Give the systematic names of the following compounds

(i) 

$CH_2 = \underset{\underset{CH3}{|}}{C} - CH3$ (1 mark)

(ii) $CH_3CH_2CH_2C \equiv CH$ (1 mark)

(b) State the observations made when Propan-1-ol reacts with:

(i) Acidified potassium dichromate (VI) Solution, (1 mark)

(ii) Sodium metal. (1 mark)

(c) Ethanol obtained from glucose can be converted to ethane as shown below

$$C_6H_{12}O_6 \xrightarrow{Step\ I} C_2H_5OH \xrightarrow{Step\ II} CH_2 = CH_2$$

Name and describe the process that takes place in steps **I** and **II**. (3 marks)

(d) Compounds **A** and **B** have the same molecular formula $C_3H_6O_2$. Compound **A** liberates carbon (IV) oxide on addition of aqueous sodium carbonate while compound **B** does not. Compound **B** has a sweet smell. Draw the possible structures of:

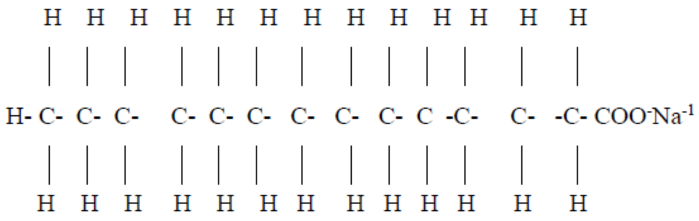
(i) Compound A (1 mark)

(ii) Compound B (1 mark)

(e) Give two reasons why the disposal of polymers such as polychloroethane by burning pollutes the environment. (2 marks)

5. 2008 Q4 P1

The structure of a detergent is:



(a) Write the molecular formula of the detergent. (1 mark)

(b) What type of detergent is represented by the formula? (1 mark)

(c) When this type of detergent is used to wash linen in hard water, spots (marks) are left on the linen. Write the formula of the substance responsible for the spots (1 mark)

6. 2008 Q1 P2

(a) Biogas is a mixture of mainly carbon (IV) oxide and methane.

(i) Give a reason why biogas can be used as a fuel. (1mark)

(ii) Other than fractional distillation, describe a method that can be used to determine the percentage of methane in biogas. (3marks)

(b) A sample of biogas contains 35.2% by mass of methane. A biogas cylinder contains 5.0 kg of the gas.

Calculate the;

(i) Number of moles of methane in the cylinder. (Molar mass of methane=16)

(2marks)

(ii) Total volume of carbon (IV) oxide produced by the combustion of methane in the cylinder (Molar gas Volume=24.0 dm^3 at room temperature and pressure).(2marks)

(c) Carbon (IV) oxide, methane, nitrogen (I) oxide and trichlorofluoromethane are greenhouse gases.

(i) State one effect of an increased level of these gases to the environment. (1 mark)

(ii) Give one source from which each of the following gases is released to the environment;

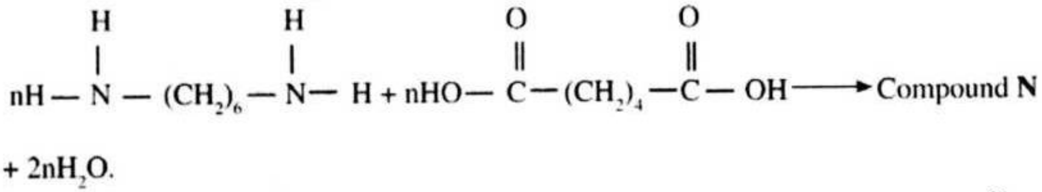
I. Nitrogen (I) oxide (1 mark)

II. Trichlorofluoromethane. (1 mark)

7. 2009 Q25

(a) Draw the structure of compound N formed in the following reaction (1 mark)

(b) Give one use of the following compound N (1 mark).



$$nH-\overset{H}{\underset{|}{N}}-(CH_2)_6-\overset{H}{\underset{|}{N}}-H+nHO-\overset{O}{\overset{\|}{C}}-(CH_2)_4-\overset{O}{\overset{\|}{C}}-OH\longrightarrow \text{Compound } N + 2nH_2O.$$

8. 2009 Q2 P2

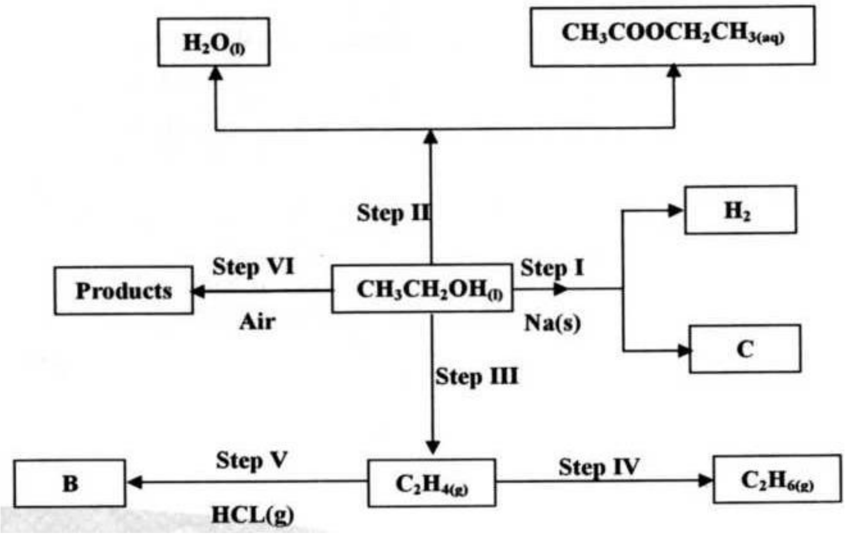
(a) Draw the structure of the following compounds (2 marks)

(i) 2 -methlybut-2-ene

(ii) heptanoic acid

(b) Describe a physical test that can be used to distinguish between hexane and hexanol (2 marks)

(c) Use the flow chart to answer the questions that follow.



(i) Name

I. The type of reaction that occurs in the step II (1mark)

II. Substance B (1mark)

(ii) Give the formula of substance C (1 mark)

(iii) Give the reagent and the condition necessary for the reaction in step (IV)

9. 2010 Q13 P1

Some animal and vegetable oils are used to make margarine and soap. Give the reagents and conditions necessary for converting the oils into:

(a) Margarine (2 marks)

(b) Soap (1 mark)

10. 2010 Q21 P1

The use of CFCs has been linked to depletion of the ozone layer.

(a) What does CFC stand for? (1 mark)

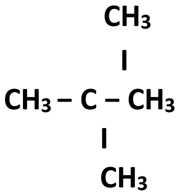
(b) Explain the problem associated with the depletion of the ozone layer (1 mark)

(c) State another environment problem caused by CFCs. (1 mark)

11. 2010 Q2 (P2)

(a) Give the name of the following compounds:

(i)

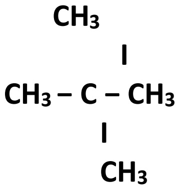


$$\begin{array}{c} CH_3 \\ | \\ CH_3-C-CH_3 \\ | \\ CH_3 \end{array}$$

(1 mark)

(ii) $CH_3C\equiv C\,CH_2CH_3$ (1 mark)

(b) Describe a chemical test that can be carried out in order to distinguish between

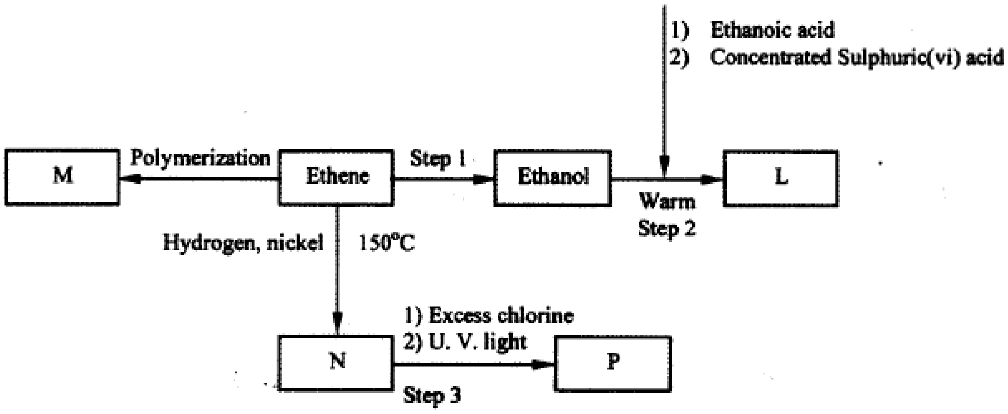


$$\begin{array}{c} CH_3 \\ | \\ CH_3-C-CH_3 \\ | \\ CH_3 \end{array}$$

and $CH_3C\equiv C\,CH_2CH_3$

(2 marks)

(c) Study the flow chart below and answer the questions that follows



(i) Name the compounds: (2 marks)

I. **L**

II. **N**

(ii) Draw the structural formula of compound **M** showing two repeating units (1 mark)

(iii) Give the reagent and the conditions used in step I (1 mark)

(iv) State the type of reaction that take place in: (2 marks)

I. Step **2**

II. Step **3**

(d) The molecular formula of compound **P** is $C_2H_2Cl_4$. Draw the two structural formulae of compound **P** (2 marks)

12. 2011 Q14 P1

Two organic compounds **P** and **Q** decolourise acidified potassium manganate (VII) solution; but only **P** reacts with sodium metal to give a colourless gas. Which homologous series does compound **P** belong? Give a reason. (3 marks)

13. 2011 Q15 P1

Soap dissolves in water according to the equation below;

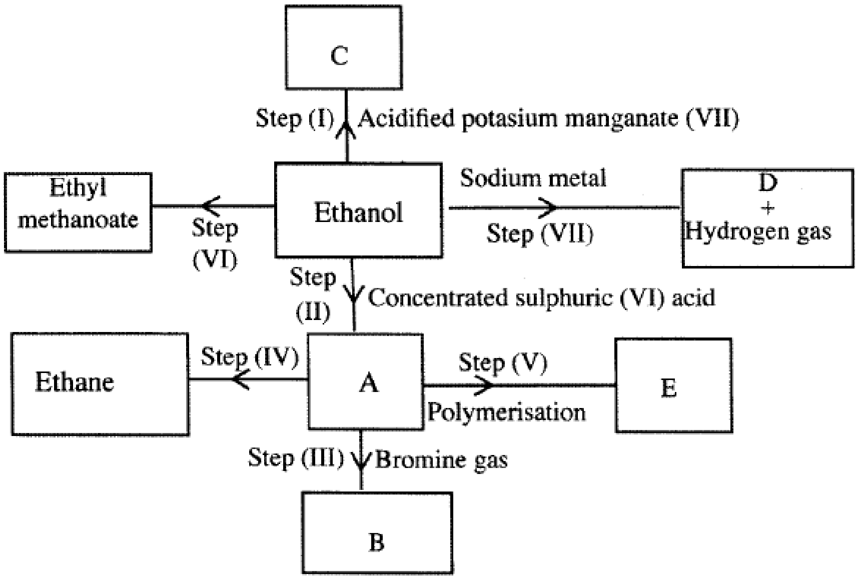
$NaSt\ (aq) \longrightarrow Na^+ (aq) + St^-$ where St^- is the stearate ion

(a) Write the formula of the scum formed when soap is used in hard water (1 mark)

(b) Write the ionic equation for the reaction that occurs when sodium carbonate is used to remove in hardness in water. (1 mark)

14. 2011 Q6 P2

(a) Study the flow chart below and answer the questions that follow.



(i) I. What observation will be made in Step **I**? (1 mark)

II. Describe a chemical test that can be carried out to show the identity of Compound **C** (2 marks)

(ii) Give the names of the following: (2 marks)

I. **E** ………………

II. Substance **D** ………………

(ii) Give the formula of substance **B** (1 mark)

(iii) Name the type of reaction that occurs in: (1 mark)

I. Step (II) ………………………

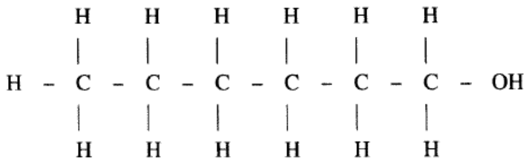
II. Step (IV) ………………………

(iv) Give the reagent and conditions necessary for Step (IV). (2 marks)

Reagent …………………………

Conditions ………………………

(b) (i) Name the following structure



$$\begin{array}{ccccccccccccccc} & & H & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & C & - & C & - & OH \\ & & | & & | & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & H & & H & & \end{array}$$

(1 mark)

(ii) Draw the structure of an isomer of pentene. (1 mark)

15. 2012 Q21 P1

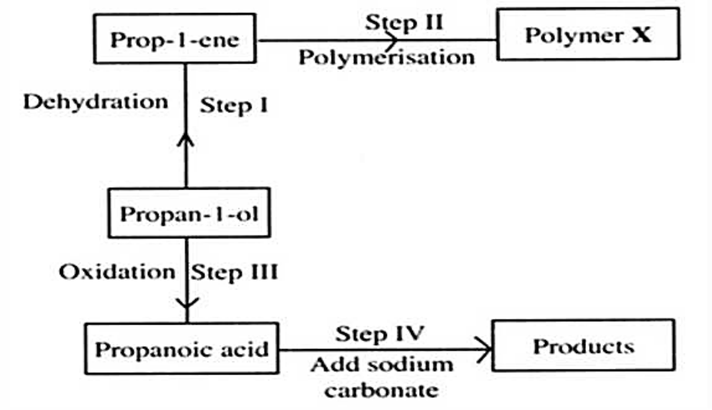
Give two uses of the polymer polystyrene. (1 mark)

16. 2012 Q1 P2, 2016 Q6 P1

(a) Draw the structural formula for all the isomers of $C_2H_3Cl_3$ (2 marks)

(b) Describe two chemical tests that can be used to distinguish between ethene and ethane.
(4 marks)

(c) The following scheme represents various reactions starting with propanol-1-ol. Use it to answer the questions that follow.

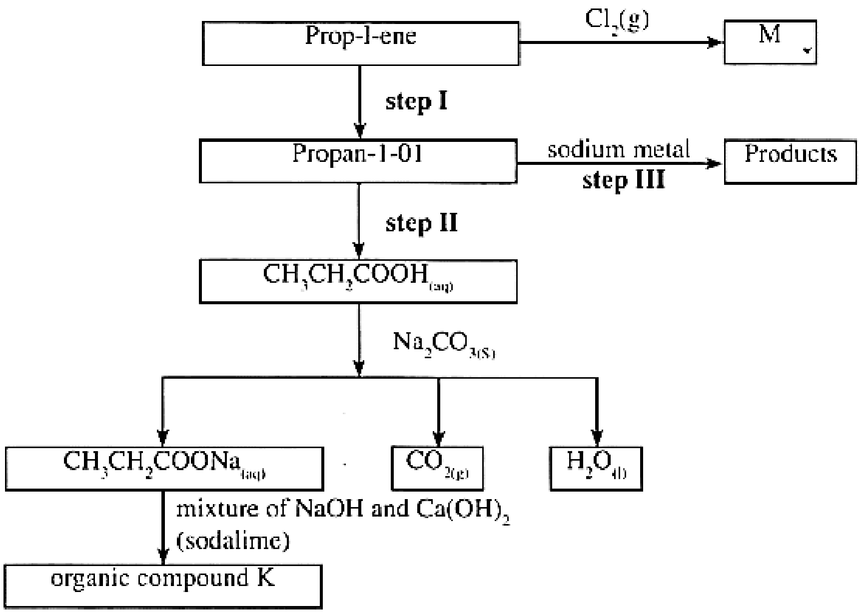


(i) Name one substance that can be used in step 1 (1 mark)
(ii) Give the general formula of X (1 mark)
(iii) Write the equation for the reaction in step IV (1 mark)
(iv) Calculate the mass of propanol-1-ol which when burnt completely in air at room temperature and pressure would produce 18 dm^3 of gas.
(C = 12.0; O = 16.0; H = 1.0; Molar gas volume = 24 dm^3) (2 marks)

17. 2013 Q7 P2

(a) Give the systematic names for the following compounds
(i) CH_3CH_2COOH; (1 mark)
(ii) $CH_3CH_2CH_2CHCH_2$; (1 mark)
(iii) $CHC\ CH_2CH_3$; (1 mark)

(b) Study the flow chart below and use it to answer the questions that follow

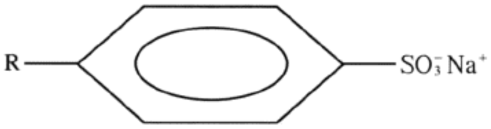


(i) Identify the organic compound K. (1 mark)
(ii) Write the formula of M (1 mark)

(iii) Give one reagent that can be used in
  (a) Step I; (1 mark)
  (b) Step II . (1 mark)

(iv) Write the equation of the reaction in step III (1 mark)

(c) The structure below represents a type of a cleansing agent.



Describe how the cleansing agent removes grease from a piece of cloth. (3marks)

18. 2014 Q9 P1

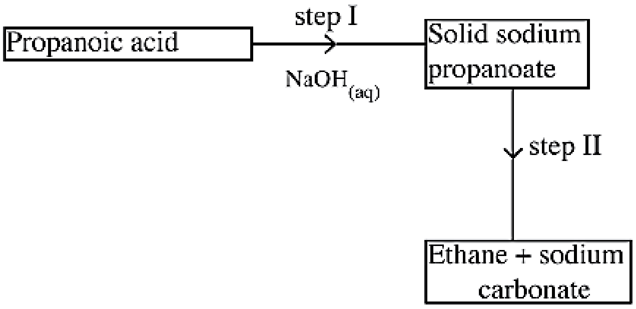
The table below shows the relative molecular masses and boiling points of pentane and ethanoic acid.

	Relative molecular mass	Boiling point (°C)
Pentane	72	36
Ethanoic acid	60	118

Explain the large difference in boiling point between ethanoic acid and pentane. (2 marks)

19. 2014 Q23 P1

Study the flow chart below and answer the question the follow.



(a) Name the process in step I. (1 mark)
(b) Identify the reagent in step II. (1 mark)
(c) Give one use of ethane. (1 mark)

20. 2014 Q26 P1

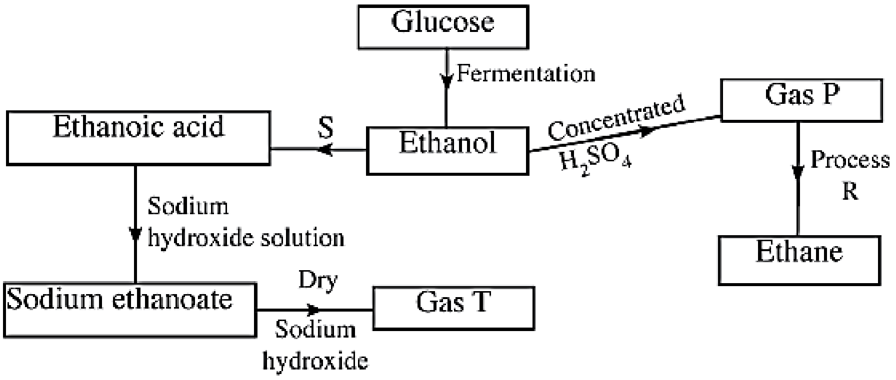
Cotton is a natural polymer. State one advantage and one disadvantage of this polymer. (2 marks)

21. 2014 Q3 P2

(a) Draw the structures of the following.

(i) Butan -1- ol (1 mark)
(ii) Hexanoic acid. (1 mark)

(b) Study the flow chart below and answer the questions that follow



(i) State the conditions necessary for fermentation of glucose to take place. (1 mark)
(ii) State one reagent that can be used to carry out process S. (1 mark)
(iii) Identify gases: (2 marks)

I. P: ………………………
II. T: ……………………

(iv) How is sodium hydroxide kept dry during the reaction (v) Give one commercial use of process R. (1 mark)

(c) When one mole of ethanol is completely burnt in air, 1370kJ of heat energy is released. Given that 1 lire of ethanol is 780 g , calculate the amount of heat energy released

when 1 litre of ethanol is completely burnt (C = 12.0; H=1.0; 0=16.0)
(3 marks)

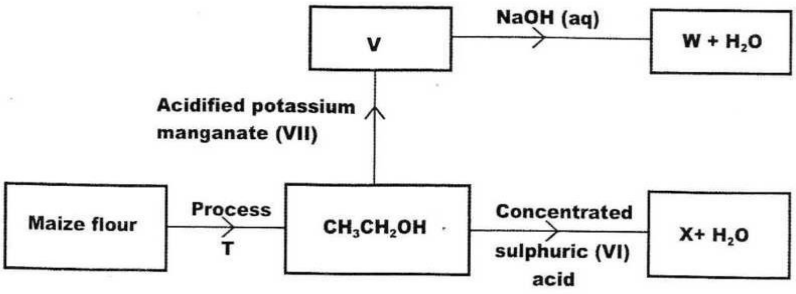
(d) State two uses of ethanol other than as an alcoholic drink. (2 marks)

22. 2015 Q1b P1

Describe a chemical test that can be used to distinguish butanol from butanoic acid. (2 marks)

23. 2015 Q22 P1

Study the flow chart below and use it to answer the questions that follows.



(a) Name process T (1 mark)
(b) Give the formula of W. (1 mark)
(c) State two uses of X (1 mark)

24. 2015 Q2 P2

(a) Draw the structure of the folio wing compounds. (2 marks)
  (i) Butanoic acid;
  (ii) Pent-2-ene.

(b) Explain why propan-1-ol is soluble in water while prop-1-ene is not. (Relative molecular mass of propan-1-ol is 60 while that of prop-1-ene is 42). (2 marks)

(c) What would be observed if a few drops of acidified potassium manganate (VII) were added to oil obtained from nut seeds? Explain. (2 marks)

(d) State one method that can be used to convert liquid oil from nut seeds into solid.

(1mark)

(e) Describe how soap is manufactured from liquid oil from nut seeds. (3 marks)

(f) 0.44 g of an ester **A** reacts with 62.5 cm3 of 0.08 M potassium hydroxide giving an alcohol **B** and substance **C**. Given that one mole of the ester reacts with one mole of the alkali, calculate the relative molecular mass of the ester. (2 marks)

25. 2016 Q2 P1 ***R***

An alkanol has the following composition by mass; hydrogen 13.5%, oxygen 21.6% and carbonate 64.9%

(a) Determine the empirical formula of the alkanol. (C=12.0, H=1.0, O=16) (2 marks)

(b) Given that the empirical formula and the molecular formula of the alkanol are the same, draw the structure of the alkanol (1 mark)

26. 2017 P1 Q20.

Study the flow chart in **Figure 5** and answer the questions that follow.

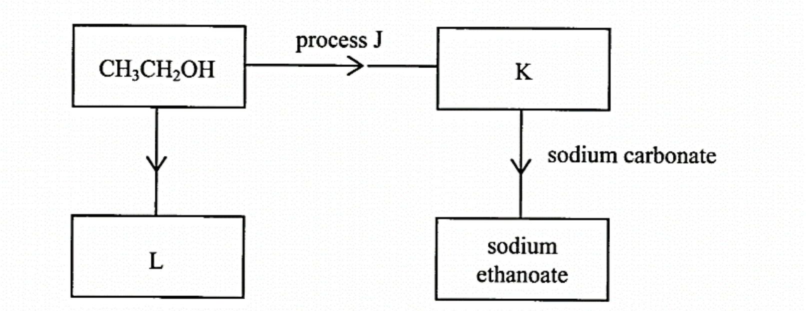


Figure 5

(c) Identify substances **K** and **L**.

K: (1 mark)

L: (1 mark)

(d) Name one reagent that can be used to carry out process J. (1 mark)

27. 2017 P1 Q28.

When an aqueous solution of compound X was mixed with a few drops of bromine water, the colour of the mixture remained yellow.

When another portion of solution X was reacted with acidified potassium dichromate (VI), the colour of the mixture changed from orange to green.

(a) What conclusion can be made from the use of:

(i) bromine water? (1 mark)

(ii) acidified potassium dichromate (VI)? (1 mark)

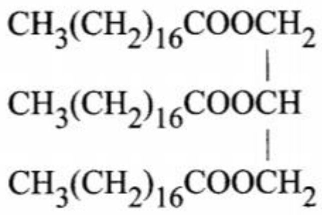
(b) Solution X was reacted with a piece of a metal and a colourless gas was produced. Describe a simple experiment to identify the gas. (1 mark)

28. 2017 P2 Q1.

(a) Name the homologous series represented by each of the following general formulae.

(i) C_nH_{2n-2} (1 mark)
(ii) C_nH_{2n} (1 mark)

(b) Compound G is a tri-ester.



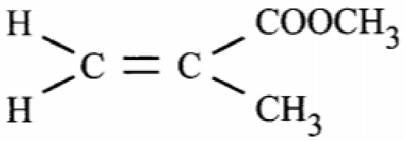
$$\begin{array}{c} CH_3(CH_2)_{16}COOCH_2 \\ | \\ CH_3(CH_2)_{16}COOCH \\ | \\ CH_3(CH_2)_{16}COOCH_2 \end{array}$$

Compound G

(i) Give the physical state of compound G at room temperature. (1 mark)
(ii) G is completely hydrolysed by heating with aqueous sodium hydroxide.
  (I) Give the structural formula of the alcohol formed. (1 mark)
  (II) Write a formula for the sodium salt formed. (1 mark)
  (III) State the use of the sodium salt. (1 mark)

(c) Ethyne is the first member of the alkyne family.
(i) Name two reagents that can be used in the laboratory to prepare the gas. (1 mark)
(ii) Write an equation for the reaction. (1 mark)

(d) Perspex is an addition synthetic polymer formed from the monomer,



(i) What is meant by addition polymerisation? (1 mark)
(ii) Draw three repeat units of Perspex. (1 mark)
(iii) Give one use of Perspex (1 mark)
(iv) State two environmental hazards associated with synthetic polymers. (1 mark)

29. 2018 P1 Q3.

The following are formulae of organic compounds. Use the formulae to answer the questions that follow:

$CH_3CH_2CH_2OH$
CH_3COOH
$CH_3CH_2CH_2CH_3$
CH_3CCCH_3

(a) Select:
(i) two compounds which when reacted together produce a sweet-smelling compound. (1 mark)
(ii) an unsaturated hydrocarbon. (1 mark)
(b) Name the compound selected in (a) (ii). (1 mark)

30. 2018 P2 Q1.

The diagram in **figure 1** shows some natural and industrial processes. Study it and answer the questions that follows

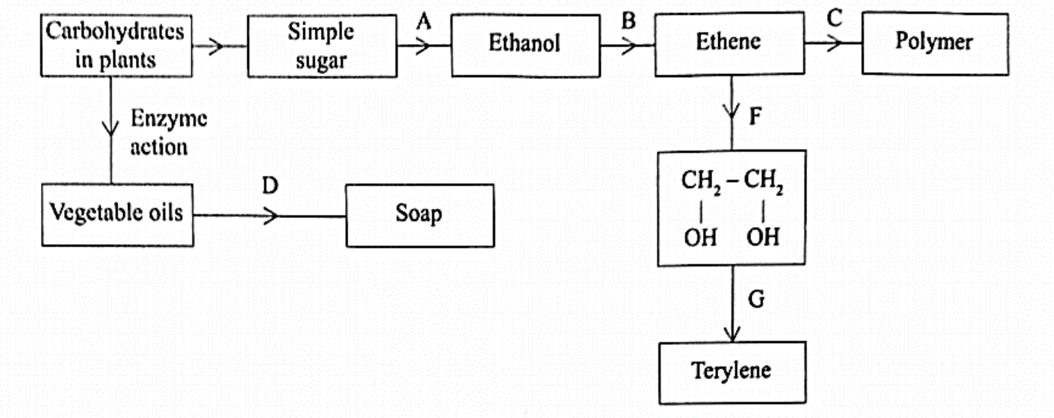


Figure 1

(a) Identify the processes labelled: (2 marks)

A..................

B..................

C..................

D..................

(b) State the reagents and conditions required for processes B and D.

(i) Process B:

Reagent................ (1 mark)

Conditions .... (1 mark)

(ii) Process D:

Reagent............... (1 mark)

Conditions .... (1 mark)

(b) Describe how process D is carried out. (2 marks)

(c) State two additives used to improve the quality of soap. (1 mark)

(d) State the reagents required in steps **F** and **G**.

(i) F.................................... (1 mark)

(ii) G.................................... (1 mark)

(e) Draw the structure of terylene. (1 mark)

(d) (i) Name the polymer formed in step **C**. (1 mark)

(ii) State one disadvantage of the polymer formed in (d) (i). (1 mark)

31. 2019 P1 Q4.

A monomer has the following structure.

$CH{=}\,CH_2$

|

C_6H_5

(a) Draw the structure of its polymer that contains three monomers. (1 mark)

(b) A sample of the polymer formed from the monomer has a molecular mass of 4992.Determine the number of monomers that formed the polymer. (C= 12; H= 1.0). (2 marks)

32. 2019 P2 Q1.

(a) Alkanes are said to be saturated hydrocarbons.

(i) What is meant by saturated hydrocarbons? (1 mark)

(ii) Draw the structure of the third member of the alkane homologous series and name it. (2 marks)

(b) When the alkane, hexane, is heated to high temperature, one of the products is ethene.

(i) Write the equation for the reaction. (1 mark)

(ii) Name the process described in (b). (1 mark)

(c) Study the flow chart in **Figure 1** below and answer the questions that follow.

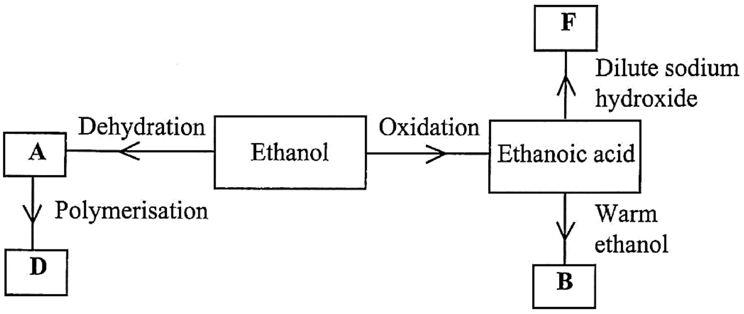


Figure 1

(i) Identify A. (1 mark)

(ii) State one physical property of B. (1 mark)

(iii) Draw the structure of D. (1 mark)

(iv) Give a reason why D pollutes the environment. (1 mark)

(v) Write an equation for the formation of F. (1 mark)

(d) Describe an experiment which can be used to distinguish butene from butanol. (2 marks)

TOPIC SEVEN

RADIOACTIVITY

Table of Contents

Objectives

By the end of the topic the learner should be able to:

(a) Define radioactivity, half-life, radioisotopes and nuclides.

(b) State the types of radioactivity, name the particles emitted during radioactive decay and state their properties.

(c) Carry out simple calculations involving half-life (t- ½).

(d) Write balanced nuclear equations.

(e) Distinguish between nuclear fission and fusion.

(f) State the uses of some radioisotopes and dangers associated with radioactivity.

RADIOACTIVITY

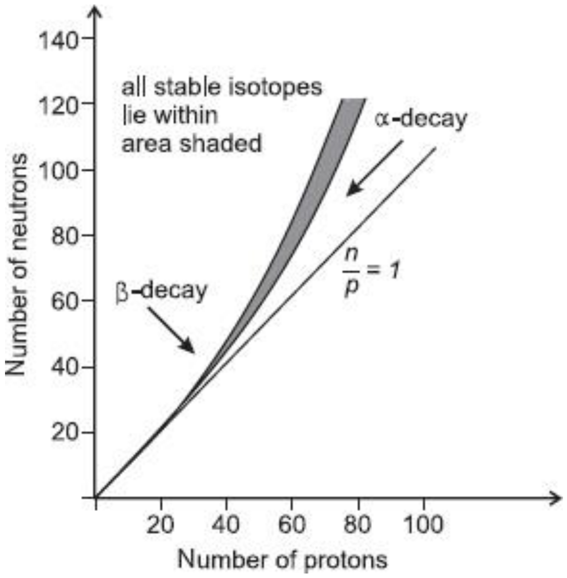
Stability of Isotopes of Elements

A particular nucleus of an atom characterised by its protons and neutrons is known as a nuclide.

Nuclides can either be stable or unstable. Most stable nuclides have a neutron/proton(n/p) ratio of about 1 : 1. For example, $^{20}_{40}$Ca has 20 neutrons and 20 protons, hence the n/p ratio is 20 : 20= 1 : 1.

For nuclides of high atomic number, the n/p ration increases progressively up to about 1.6 : 1 which is still within the stability region. Above this value the nucleus becomes too large and unstable.

The unstable nuclide normally undergoes some changes by emitting radiation in form of particles and energy. Such emissions results in the production of new nuclides of completely different composition.



Radioactivity as a Nuclear Reaction

Radioactivity is the process where an unstable nuclide breaks up to yield another nuclide of different composition with emission of particles and energy.

Substances, which undergo radioactivity, are said to be <u>radioactive</u>. Isotopes, which are radioactive, are known as <u>radioisotopes</u>.

The spontaneous disintegration of radioactive nuclides is known as <u>radioactive decay</u>. ***Radioactivity is a nuclear process and not a chemical reaction.***

Differences between chemical and nuclear reactions

Nuclear Reaction	***Chemical Reaction***
1. Takes place within the nucleus and involves neutrons and protons.	Takes place on the outer energy levels containing the valence electrons. Does not involve protons and neutrons.
2. Releases large amounts of heat energy.	The heat energy released is much less.
3. Not affected by environmental factors such as temperature.	Are affected by environmental factors such as temperature and pressure.

Types of Radioactivity

There are two types of radioactivity namely natural and artificial.

Natural radioactivity occurs when radioactive nuclei **split spontaneously** yielding a new nuclide with the emission of radiation and energy.

For example, the nucleus of uranium– 238 undergoes natural radioactive decay to form Thorium – 234 and some radiation.

Artificial radioactivity occurs when large stable nuclides are **bombarded with fast moving high-energy particles**. In the process, the nuclides split into relatively smaller nuclei with emission of radiation and energy.

Types, Characteristics and Properties of radiations

When radioactive nuclides disintegrate, they emit radiations namely alpha particles (α), beta (β) and gamma (γ).

Characteristics of Radiations

Alpha radiations are particles which are positively charged helium nuclei, He**2+**. They are represented as $^{4}_{2}He^{+}$ in nuclear equations. Alpha particles have an electrical charge of 2**+**. They are the heaviest of the three radiations.

Beta radiations are particles which are negatively charged. They are represented as $^{0}_{-1}e$. They are electrons, which originate from within the nucleus and not from the outer energy levels. They are formed when a neutron changes into a proton within the nucleus.

$$\underset{\text{neutron}}{^{1}_{0}n} \longrightarrow \underset{\text{proton}}{^{1}_{1}p} + \underset{\text{electron}}{^{0}_{-1}e}$$

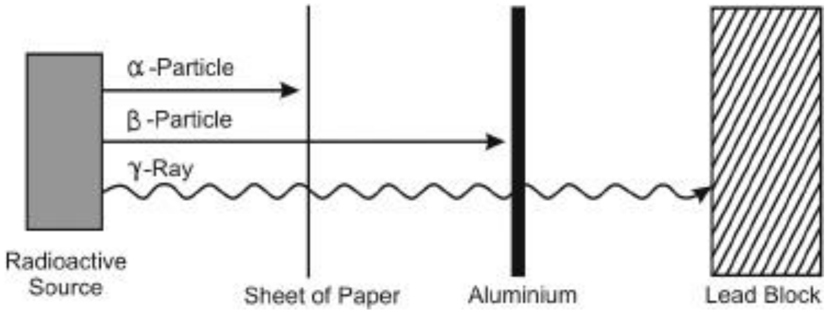
The –1 is not an atomic number. It represents the charge on the particles. Beta particles have an electrical charge of **–**1.

Gamma radiations are high energy rays. They do not have an electrical charge. They are not emitted on their own but normally accompany the emission of alpha or beta particles.

Properties of Radiations

Penetrating power

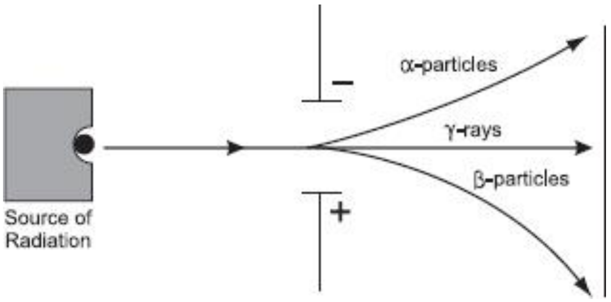
Alpha particles have very **low penetrating** power. They do not pass through a sheet of paper. The have a range of only a few centimetres in air.



Beta particles have **higher penetrating power**. They can pass through a sheet of paper but are stopped by a sheet of aluminium foil.

Gamma **rays** have the **highest penetrating power.** They pass through a sheet of paper and a sheet of aluminium. They are however stopped by a thick lead block.

Deflection by an Electric Field



Alpha particles are **positively charged**. They are therefore **attracted to the negative plate** in an electric field. Since they are **heavy** they are only **slightly deflected**.

Beta particles are **negatively charged** and so are **deflected to the positive plate** in an electric field. The **deflection is greater** than that of alpha particles due to their much **smaller masses**.

Gamma rays do not have a charge and are therefore **unaffected** by an electric field.

Ionising Effect of Radiations

Alpha particles have **very high ionising** power. They produce a **large number of ions** as they pass through gases due to their slow speed and high charge. Their slow speed enables the alpha particles to be in contact with target atoms for a longer time.

Beta particles have **lower ionising power** and produce fewer ions in gases.

Gamma rays have **very low ionising** power.

Radioactive Decay and Half-Life

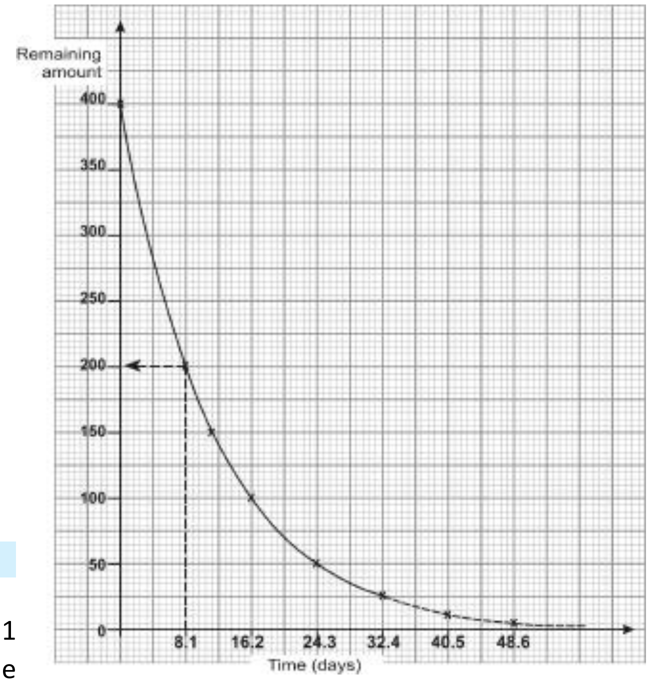
When a radioactive nuclide decays, a new nuclide is formed. As disintegration proceeds, fewer and fewer unstable atoms remain.

The Table below shows the radioactive decay for a sample of 400 g of iodine – 131

Decay of iodine -131

Time (days)	*Mass (g) of radioactive iodine remaining*
0	400
8.1	200
16.2	100
24.3	50
32.4	**25**

A graph of the data given is shown alongside.



From the graph it can be seen that after every 8.1 days the amount of substance remaining is half the previous amount.

The fractions of the original amount remaining after:

(i) 8.1 days is $\frac{1}{2}$

(ii) 16.2 days if $\frac{1}{4}$

(iii) 24.3 days is $\frac{1}{8}$

From the graph, it is also observed that the remaining amount never reduces to zero. The time taken for the mass to reduce to half the previous value is a constant, i.e., 8.1 days. This constant value is referred to as the half-life. It is denoted by $t\frac{1}{2}$.

The half-life *of a radioactive isotope is the time taken for a given mass or number of nuclides to decay to half its original mass or number.*

From the sample of 400 g iodine – 131 given in the example, the amount remaining after the first half-life (8.1 days) will be 200 g. After the second half-life (16.2 days) it will be 100g. This can be illustrated as shown.

$$400\text{ g} \xrightarrow[8.1\ days]{1^{st}\ half\text{-}life} 200\text{ g} \xrightarrow[8.1\ days]{2^{nd}\ half\text{-}life} 100\text{ g} \xrightarrow[8.1\ days]{3^{rd}\ half\text{-}life} 50\text{ g}$$

The amount remaining after the first half-live is half the previous amount. The remaining amount after successive nuclide decay can be worked out by using the formula;

Remaining amount = $\left(\frac{1}{2}\right)^n$ × original amount, where n is the number of half-lives undergone. For example the remaining amount after 24.3 days (3 half-lives) is $\left(\frac{1}{2}\right)^3 \times 400 = \frac{1}{8} \times 400 = 50\text{ g}$.

Given the remaining amount, the original amount can also be determined. For example: if 3g of $^{257}_{103}\text{Lr}$ Lr whose half -life is 8 seconds remain after undergoing radioactive decay for 32 seconds.

The number of half-life = $\frac{32}{8} = 4$

Using the formula :– $\left(\frac{1}{2}\right)^4 \times$ original amount (x) = 3

Original amount $\times \frac{1}{16} = 3$

$= 48\text{ g}$

Alternatively, the original amount can be found using a step by step method thus;

$$3\text{ g} \xrightarrow[8\ s]{4^{th}\ half\text{-}life} 6\text{ g} \xrightarrow[8\ s]{3^{rd}\ half\text{-}life} 12\text{ g} \xrightarrow[8\ s]{2^{nd}\ half\text{-}life} 24\text{ g} \xrightarrow[8\ s]{1^{st}\ half\text{-}life} 48\text{ g}$$

The original amount is 48 g.

Each radioactive isotope decays at its own rate and therefore has its own half- life. Half-life $\left(t\frac{1}{2}\right)$ range from fraction of a second to millions of years.

For example:

Nuclide	half-life
$^{212}_{84}Po$	2.96×10^{-7} seconds
$^{257}_{103}Lr$	8 seconds
$^{234}_{91}Pa$	1.14 minutes
$^{234}_{90}Th$	24.5 days
$^{1}_{1}H$	12.3 years
$^{14}_{6}C$	5570 years
$^{234}_{92}U$	4.5×10^{9} years

The shorter the half-life, the faster the rate of decay of the nuclide. The longer the half-life, the slower the rate of decay of the nuclide. The rate of radioactive decay is unaffected by any chemical or physical change.

The half-life of radioactive isotopes is one of the factors used to determine their application. For example, carbon-14 is used to determine the age of dead organic matter. The process is known as carbon-dating. The carbon in plants which is taken up during photosynthesis contains a small portion of radioactive carbon-14. When the plant dies, the carbon-14 in the dead plant continues decaying hence the amount decreases. Carbon-14 has half-life a of about 5,600 years therefore by determining the carbon content in the dead material it is mathematically possible to determine the age of the sample.

Sodium-20 with a half-life of 0.3 seconds is used to detect leakages because it has a short half-life.

Nuclear Reactions

When an alpha particle is emitted from the nucleus, both the atomic number and the mass number of the nuclide decrease. A new nuclide is formed as represented by the following equation

$$^{234}_{91}Pa \longrightarrow {}^{229}_{89}Ac + {}^{4}_{2}He$$

The equation above is referred to as a nuclear equation.

The sum of the mass number on the right hand side of the equation is equal to the mass number of the left hand side. The sum of proton numbers on the right hand side equals the number of protons on the left hand side.

When a beta particle is emitted, the atomic number increases by one and the mass number remains constant. This is because:

(i) During beta emission, a neutron spontaneously changes into a proton and emits an electron. That is:

$$^{1}_{0}n \longrightarrow {}^{1}_{1}P + {}^{0}_{-1}He$$

The positive charge increases in the process.

(ii) The mass of the particles emitted is negligible.

Consider the following nuclides, which emit beta particles.

(i) $^{14}_{6}C \longrightarrow {}^{14}_{7}N + {}^{0}_{-1}e$

(ii) $^{226}_{88}n \longrightarrow {}^{226}_{89}Ac + {}^{0}_{-1}e$

The emission of gamma rays which is a form of energy always accompanies other radioactive emissions. They are produced when the remaining particles in the nucleus re-organise themselves into more stable arrangements. Gamma rays are not shown when writing nuclear equations because they have no effect on mass number and atomic number of a nuclide.

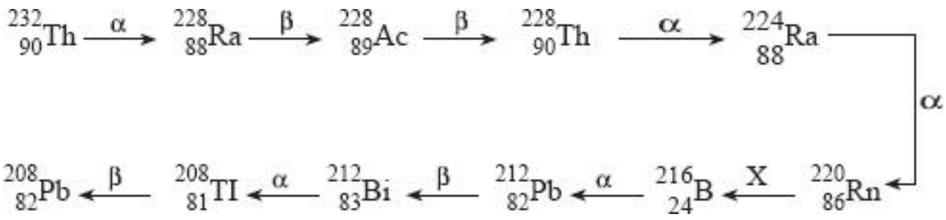
For example:

$$^{236}_{92}U \longrightarrow {}^{141}_{56}Ba + {}^{92}_{36}Kr + 3^{1}_{0}n \text{ + energy}$$

Radioactive Decay Series.

A radioactive decay series represents the sequential and continuous disintegration of unstable nuclide until a stable nuclide is formed. For example starting with thorium–232.

The following series is obtained;



In this decay series, natural decay ends when stable lead-208 is formed. Notice that the $\frac{n}{p}$ ratios of lead – 208 is about 1.5:1 which is in the stable region. This decay series is known as the Thorium decay series.

Nuclear Fission

Nuclear fission ***is the splitting process a heavy nuclide undergoes when bombarded by a fast-moving neutron***.

The large unstable nuclide absorbs a neutron and immediately splits up into two approximately equal fragments together with a number of smaller particles such as a neutron. During this process, much energy is liberated.

For example:

$$^{236}_{92}U + {}^{1}_{0}n \longrightarrow {}^{141}_{56}Ba + {}^{92}_{36}Kr + 3^{1}_{0}n \text{ + energy}$$

The mass of the products added together equal the mass of the original nuclide plus the mass of the neutron. The three neutrons produced further bombard more nuclei and more neutrons are formed. This results in a chain reaction.

Energy is liberated due to the fact that the total mass of the products is slightly less than the mass of the initial nuclides. This difference in mass is radiated as energy. A very small loss in mass results in an enormous amount of energy being released.

This energy can be tapped and utilised to generate electrical energy and other forms of energy under controlled conditions. The enormous heat is used to heat water to produce steam which then turns turbines to produce electricity. The energy of an atomic bomb is due to nuclear fission of uranium–235.

Nuclear Fusion

Nuclear fusion occurs when nuclei combine together when they are made to collide at high velocity resulting in the formation of a heavy nucleus and release of energy.

Some sub-atomic particles such neutrons are also released during nuclear fusion. The energy released after fusion causes other nuclide to collide and a chain reaction occurs. For example:

$$^{3}_{1}H + {}^{2}_{1}H \longrightarrow {}^{4}_{2}He + {}^{1}_{0}n + \text{energy}$$

+ energy

The sum of mass numbers and atomic numbers of the products equal the sum of mass numbers and atomic number of the reactants.

The energy released during nuclear fusion can be harnessed and converted into other forms of energy such as electrical energy.

The hydrogen bomb works on the principle of nuclear fusion.

Nuclear fission and fusion are similar in that:

1. In both cases a large quantity of energy is released.
2. Both processes results in chain reactions.
3. In both cases sub-atomic particles such as neutrons accompany the process.
4. The energy released can be harnessed and converted into other forms of useful energy such as electrical energy.
5. The large amount of energy produced in both reactions can be very destructive when misused such as in nuclear warfare.

Applications of Radioactivity

Besides production of energy, other uses of radioactivity include the following:

Medical applications

- Used to destroy cancerous tissue when a patient is exposed to correct dose of radiation from radioactive nuclide such as a cobalt–60 and caesium–137, it is possible to destroy cancerous growth without serious damage to other tissues of the patient.
- Sterilisation of surgical instruments using gamma radiation.
- Radioactive iodine (iodine–131) is used in patients with defective thyroid to enable doctors to follow the path of iodine through the body.
- Used to monitor growth in bones and healing of fractures.
- For providing power in heart pace setters.
- Detecting leakages in underground water or oil pipes without digging them out.

Agricultural applications

Radioactivity may be used in monitoring:

- Photosynthesis and related process; Carbon(VI) oxide containing radioactive carbon –14 is used, and the path of this carbon can be followed during the growth of the plant.
- Absorption of phosphate fertilisers; radioactive phosphorus can be used to determine the rate of absorption of the fertiliser.

Other uses

- Preservation of food by exposing micro-organisms to gamma radiation which kills them.
- Gauging the thickness of thin metal and paper sheets.
- Measuring the level of food in canned and packed food.
- Determining the age of archeological materials in fossils from carbon –14 dating.
- Manufacture of nuclear weapons and atomic bombs.

Dangers of Radioactivity

Environmental Pollution

Environmental pollution occurs when radioactive materials emit radiation into the atmosphere. Long term exposure to low dosages of these radiations can cause genetic mutation in living tissue leading to anaemia, bone cancer and other forms of cancer.

Disposal of nuclear waste is of particular concern since some of the radioactive materials have very long half-life.

Testing of nuclear weapons in the oceans also causes environmental pollution since plants and other living organisms may take in the radioactive materials released in the water.

When not put into proper use, radioisotopes can be used as weapons of mass destruction as it happened in the cities of Hiroshima and Nagasaki, Japan during the Second World War.

Control of Environmental Pollution

Environmental pollution can be controlled by proper use, storage and disposal of radioactive materials as well as regular checks of equipment which emit radiations.

Revision Exercise

1. 2006 Q 4

(a) Complete the nuclear equation below. (1 mark)

$$^{37}_{18}\mathrm{A} + \ldots\ldots\ldots \longrightarrow {}^{37}_{17}\mathrm{B}$$

(b) (b) State one:

(i) Use of radioisotopes in agriculture. (1 mark)

(ii) Danger associated with exposure of human beings to radioisotopes. (1 mark)

2. 2007 Q 14 P1

(a) Distinguish between nuclear fission and nuclear fusion. (2 marks)

(b) Describe how solid wastes containing radioactive substances should be disposed of. (1 mark)

3. 2008 Q 24 P1

(a) A radioactive substance emits three different particles. Give the symbol of the particle with the highest mass. (1 mark)

(b) (i) Find the values of **Z_1** and **Z_2** in the nuclear equation below. (1 mark)

$$^{Z1}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{37}_{17}\mathrm{B} + {}^{140}_{Z2}\mathrm{Xe} + 2\,{}^{1}_{0}\mathrm{n}$$

(ii) What type of nuclear reaction is represented in represented in b (i) above? (1 mark)

4. 2009 Q 6d P2

Naturally occurring uranium consist of three isotopes which are radioactive.

Isotope	$^{234}\mathbf{U}$	$^{235}\mathbf{U}$	$^{238}\mathbf{U}$
Abundance	0.01%	0.72%	99.27%

(i) Which of these isotopes has the longest half-life? Give reasons. (1 mark)

(ii) Calculate the relative atomic mass of uranium. (2 marks)

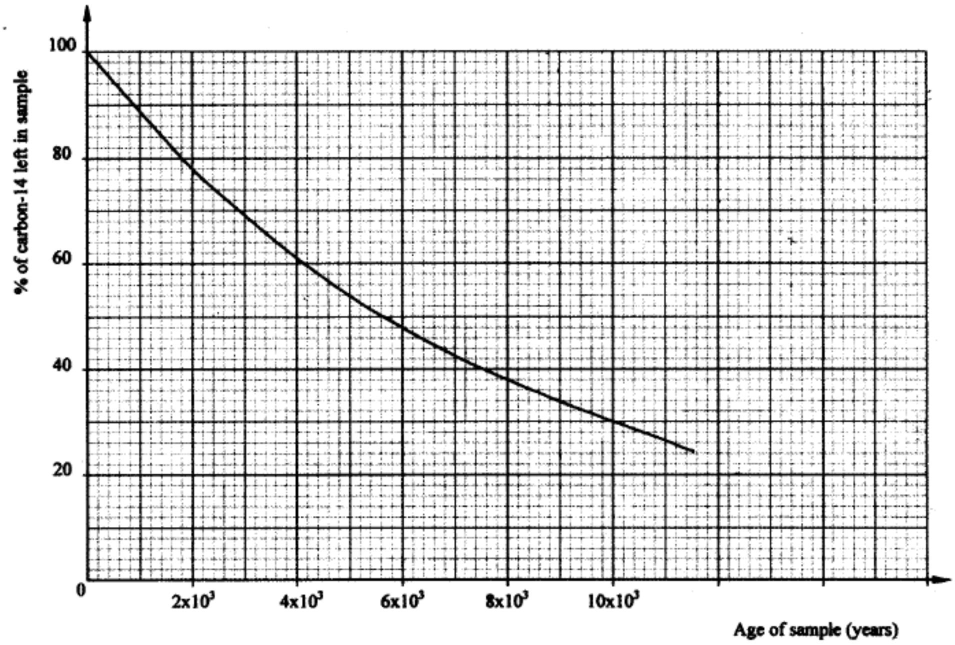
(iii) $^{235}_{92}\mathbf{U}$ is an alpha emitter. If the product of the decay of this nuclide is thorium (Th), Write a nuclear equation for the process. (1 mark)

(iv) State one use of radioactive isotopes in the paper industry. (2 marks)

5. 2010 Q 9 P1

Carbon -14, $^{14}_{6}\mathbf{C}$, is used in carbon dating. It decays to form nitrogen, $^{14}_{7}\mathbf{N}$

The graph below shows the amount of carbon -14 left in a sample against its age in years.



(a) Write a nuclear equation for the decay process of carbon -14. (1 mark)

(b) From the graph, determine the;

- (i) Half-life of carbon -14; (1 mark)
- (ii) Percentage of carbon -14 in a sample whose age is 1950 years. (1 mark)

6. 2011 Q 2 P1

(a) Complete the nuclear equation below: (1 mark)

$$^{131}_{53}\mathbf{I} \longrightarrow {}^{131}_{54}\mathbf{I} + \_\_\_\_\_\_$$

(b) The half-life of $^{131}_{53}\mathbf{I}$ is 8 days. Determine the mass of $^{131}_{53}\mathbf{I}$ remaining if 50 grams decayed for 40 days. (1 mark)

(c) Give one harmful effect of radioisotopes. (1 mark)

7. 2012 Q9 P1

120g of iodine - 131 has a half-life of 8 days decays for 32 days. On the grid provided, plot a graph of the mass of iodine - 131 against time. (3 marks)

8. 2013 Q8 P1

Draw a labelled diagram to illustrate how alpha, beta and gamma radiations can be distinguished from each other.
(3 marks)

9. 2013 Q6 P2

(a) Distinguish between a neutron and proton. (1 mark)

(b) What is meant by a radioactive substance? (1 mark)
(c) State two dangers associated with radioactive substance in the environment. (2 marks)

(d) The two isotopes of hydrogen, deuterium and tritium T react to form element y and neutron particles, according to the equation below;

$$^{2}_{1}\mathrm{D} + ^{3}_{1}\mathrm{T} \longrightarrow ^{a}_{b}\mathrm{Y} + ^{1}_{0}\mathrm{n}$$

(i) What is the atomic:
(I) Mass of Y? (1 mark)
(II) Number of Y? (1 mark)

(ii) What name is given to the type of reaction undergone by the isotopes of hydrogen? (1 mark)

(e) (i) What is meant by half-life of a radioactive substance? (1 mark)
(ii) 288 g of a radioactive substance decayed to 9 g in 40 days. Determine the half-life of the radioactive substance. (2 marks)

10. 2014 Q8 P1

(a) Complete the nuclear reaction below. (1 mark)

$$^{226}_{88}\mathrm{Q} \longrightarrow ^{222}_{86}\mathrm{P} +$$

(b) State two uses of radioisotopes in health. (2 marks)

11. 2015 Q19 P1

A radioactive substance weighing M kg took 1900 years for the original mass to reduce to 15 kg. Given that the half-life of the radioactive substance is 380 years;
(a) Determine the original mass of the radioactive substance. (2 marks)
(b) State two uses of radioactivity in medicine. (1 mark)

12. 2017 P1 Q2.

Calculate the values of X and Y in the following nuclear equation.

$^{239}_{92}\mathrm{U} \longrightarrow ^{X}_{Y}\mathrm{Th} + 2\alpha + 2\beta$ (2 mark)

13. 2017 P2 Q7.

The decay rates of a sample of a radioisotope of bismuth at different time intervals is indicated in the following table.

Time (hours)	0	5	10	15	20	25
Rate of disintegration in counts s^{-1}	730	570	455	365	292	232

(a)

(i) Draw a graph of disintegration rate against time. (3 marks)
(ii) Determine the half-life of bismuth. (1 mark)
(iii) What would be the effect on the curve if half the amount of sample of bismuth were used. (1 mark)

(b) Radioactivity has several applications. State one application of radioactivity in:

(i) Medicine (1 mark)
(ii) Agriculture (1 mark)
(iii) Tracers (1 mark)
(iv) Nuclear power station (1 mark)

(c) State two dangers associated with radioactivity. (2 marks)

14. 2018 P1 Q 19.

(a) Give the symbols of the two charged particles emitted by a radioactive isotope. (2 marks)

(b) An isotope $^{210}_{82}\mathbf{Pb}$ disintegrates by emitting two beta particles. Determine the mass number and atomic number of the resulting nuclide. (1 mark)

15. 2019 P1 Q 22.

The diagram in **figure 6** shows radiations emitted by a radioactive sample.

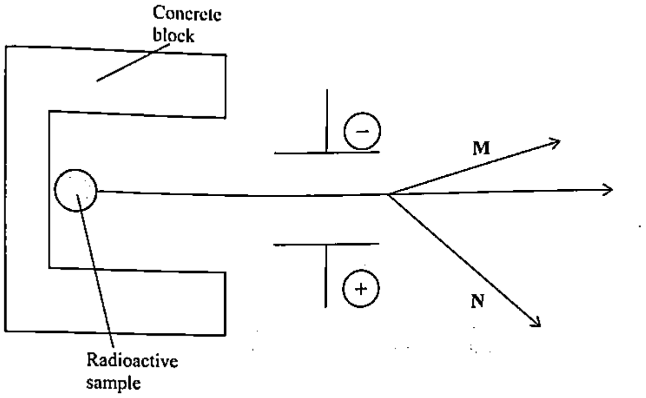


Figure 6

(a) Identify radiations:

(i) M .... (1 mark)
(ii) N .... (1 mark)

(b) Explain what would happen when a sheet of paper is placed in the path of the two radiations. (1 mark)