

CHEMISTRY NOTES

FORM 3 WORK

NEW SYLLABUS

EXPERIMENTAL TECHNIQUES

Safe ways of disposing laboratory wastes

1. Use of normal waste bins

Controlled wastes include non-recycle plastics, piece of wood, rubber and dirty papers.

2. Draining using a lot of water

Harmless soluble inorganic salts such as magnesium sulphate and sodium sulphate.

3. Use of special controlled waste containers

Sharp metal objects e.g. scalpel, broken laboratory glassware, fine powders inside a bottle, sample tube and items highly contaminated with hazardous chemicals are placed in these containers.

4. Incineration

Wastes that are incinerated include: syringes and needle, all organic solvents, soluble organic wastes, paraffin and mineral oils.

5. Use of fume chambers

This is used for experiments involving poisonous waste gases such as chlorine, sulphur dioxide, hydrogen chloride and other carcinogen solids.

6. Recycling

Each laboratory should have a bin for recyclable materials such as unbroken glass, packing waste and paper.

Designing scientific investigation

I. Problem identification

This is the first step where a problem is recognized and a clear question asked. For example, what causes rusting?

2. Formulation of hypothesis

A hypothesis is either a tentative theory or a tentative explanation for the observation made. A theory is a set of scientific assumptions consistent with one another and supported by evidence but not fully proved. For example, rusting is caused by heat, moisture and air.

3. Experimentation

An experiment is a series of investigations intended to discover certain facts, which may lead to either acceptance or rejection or modification of a hypothesis. For example, an experiment can be designed to investigate conditions necessary for rusting to take place.

4. Observation and data collection

After setting an experiment, careful observations and recordings of all events are done

5. Interpretation of data

The researcher tries to explain the meaning (results) of the data in relation to the aim of the experiment.

6. Conclusion

Conclusions must be made based on the data collected from the experiment. The conclusion is either the confirmation of the hypothesis or a rejection.

Determining purity of substance

Pure substances have fixed (specific) melting and boiling points. **Impurities** causes rise in boiling points in liquids and reduces melting points in solids. The boiling points also depend on the prevailing atmospheric pressure. An experiment can be conducted to investigate the effect of impurities on the boiling point of water.

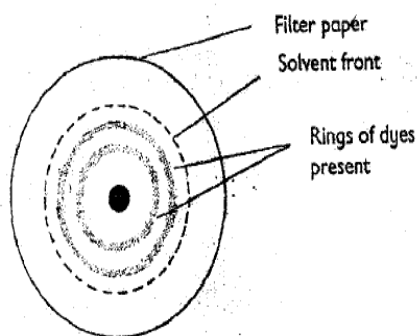
Paper chromatography

Chromatography is a technique for separation of components of a solution by passing it through a medium. Paper chromatography is one of the several types of chromatography.

The components separate readily according to how strong they are absorbed by the medium in comparison on how readily they dissolve in the travelling solution (solvent).

The component that is most soluble moves faster up the paper. An insoluble component remains at the origin. The movement (flow) stops just before the solvent reaches the top of medium (**solvent front**)

Examples of solvents used in chromatography include ethanoic acid, ethanol and propanone.

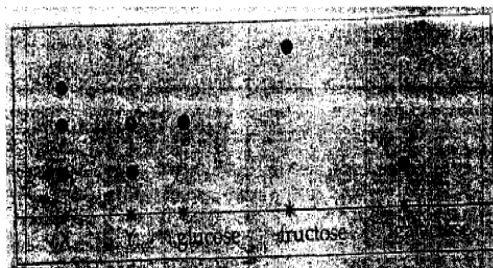


Relative flow value (R_f)

This is the distance travelled by a particular component (spot) in comparison to the distance travelled by the solvent (i.e. solvent front). The ratio of these two distances is called R_f value and it ranges from zero to one.

$$\frac{\text{Distance moved by the substance}}{\text{Distance moved by the solvent front}}$$

A pure substance has only one R_f value.



Which two sugars are present in X and Y?

Tests for ions, gases and water

1. Test for of cations (positively charged ions)

Some commonly tested cations include calcium, aluminium, ammonium, zinc and copper. They react with the reagents aqueous sodium hydroxide and aqueous ammonia solutions.

Cation	3 drops	In excess
Ca^{2+}	White precipitate	Insoluble
Al^{3+}	White precipitate	Insoluble
Zn^{2+}	White precipitate	Precipitates dissolve
Cu^{2+}	Blue precipitate	Dissolves to form deep blue solution
Fe^{3+}	Green precipitate	Insoluble
Fe^{2+}	Red brown precipitate	Insoluble
Pb^{2+}	White precipitate	Insoluble

2. Test for of anions (negatively charged ions)

Common anions tested are carbonate (CO_3^{2-}), chloride, iodide, nitrate and sulphate.

- Dilute acid such as hydrochloric acid and nitric acid is added to a carbonate. Effervescence occurs and carbon dioxide and water are produced.
- Acidified silver nitrate is added to aqueous solution of chloride. A white precipitate forms confirming the presence of chloride ions.
- Nitric acid is added to iodide solution and aqueous lead (II) nitrate. A yellow precipitate forms.

- Sodium hydroxide is added to a nitrate solution and drop aluminium foil. A damp red litmus paper turns blue.

3. Identification of gases

Gases such as ammonia, carbon dioxide, chlorine, hydrogen, oxygen and many others can be tested to confirm their presence.

- To confirm the presence of ammonia introduce a damp red litmus paper into a gas jar containing ammonia. It will turn blue.
- To confirm the presence of carbon dioxide, bubble the gas produced through lime water. The lime water will turn milky.
- To confirm the presence of chlorine gas, hold a moist blue litmus paper at the mouth of a test tube containing chlorine gas. It will turn red and then get bleached.
- To confirm the presence of oxygen, introduce a glowing splint into a jar containing oxygen gas. The splint will relight into a flame.

4. Test for the presence of water

In an experiment, the change of colour of white anhydrous copper (II) sulphate to blue and that of blue anhydrous cobalt chloride to pink is a confirmatory test for water.

Qualitative analysis

Qualitative analysis involves making careful observations, drawing suitable conclusions and inferences based on the observations made. When you are conducting an experiment, the procedures, observations and the conclusions you present in table form or otherwise gives what is called qualitative analysis.

NITROGEN, SULPHUR AND PHOSPHORUS

Sources of Nitrogen

Nitrogen is available in the air. It is also found as a constituent of certain compounds:

- Fertilisers e.g. ammonium nitrate NH_4NO_3
Ammonium phosphate $(\text{NH}_4)_3\text{PO}_4$
- Proteins
- Amino acids

Components of Nitrogen

Nitrogen has a number of components. These include:

- Ammonia
- Acids of nitrogen such as nitric acid, nitrous acid
- Oxides of nitrogen such as nitrogen (I) oxide, nitrogen (II) oxide and nitrogen (IV) oxide

Preparation of Nitrogen

Nitrogen can be prepared in the laboratory. It can also be industrially prepared (isolation from air).

Laboratory preparation of Nitrogen

The two common methods of preparing nitrogen in the laboratory are:

1. Isolation of nitrogen from the air.
2. Reaction of sodium nitrite with any ammonium salt

Physical properties of Nitrogen

- Has no colour
- Has no smell (odourless)
- It is insoluble in water
- It is less dense than air

Chemical properties of Nitrogen

- Nitrogen does not burn or support combustion.
- Litmus papers show no colour change
- Nitrogen has no effect on calcium hydroxide solution
- Nitrogen reacts with a burning magnesium metal to form magnesium nitride (Mg_3N_2).
- Nitrogen does not react with burning sulphur.

Inert character of Nitrogen

Inertness refers to the reluctance of an element to participate in reactions. Nitrogen is a diatomic gas that have strong triple covalent bonds ($\text{N}\equiv\text{N}$) between the atoms of its molecules. These bonds require a lot of energy to break.

Industrial isolation of nitrogen from air

Large scale production of nitrogen is carried out industrially through a process called **fractional distillation** or **liquefaction** of air.

The first step in this process is the removal of impurities such as water vapour, dust particles and carbon dioxide. Carbon dioxide and water vapour which solidify easily would block the pipes if not removed at this stage.

The air then compressed to about 200 atmospheres making it very hot. It is then allowed to expand through a jet causing it to become very cold. Some of it turns into liquid (liquefy). The compression and expansion process is repeated several times and each time the air get much colder. By the time it reaches -200°C , most of the gases, that is, nitrogen, oxygen and argon have become liquid except neon and helium. The two noble gases are thus removed.

Nitrogen and oxygen have different boiling points. This makes it easy to separate them during the liquefaction process. Liquid nitrogen having lower boiling point (-196°C) distills first. Then nitrogen is

collected and stored under pressure in cylinders.

The remaining liquid containing mainly oxygen and argon is further heated. Argon whose boiling point is -186°C distills leaving oxygen with a high boiling point of -183°C . These two gases are collected and stored.

Uses of Nitrogen

- It is used in the Haber process to manufacture ammonia gas
- Used as a coolant to quickly freeze substances. For example it is used for long period of storage of body tissues and bull's semen for future use
- Mending of pipes. Freezes the liquid inside allowing repair work to be carried out.
- Since most bacteria cannot survive in absence of oxygen, it is used in food processing and packaging to provide an inert atmosphere preventing oxidation of the food.
- Because of its inert nature at low temperatures it is pumped in the ships' tanks transporting crude oil to remove any oxygen that can cause explosion.

Nitrogen compounds

1. Ammonia

Ammonia gas can be prepared in the laboratory industrially by the **Haber process**.

Laboratory preparation of ammonia gas

Ammonia gas is usually prepared in the laboratory using an ammonia salt and a base. When a mixture of calcium hydroxide and ammonium chloride is heated ammonia gas is given off. The gas is dried using calcium oxide.



Physical properties of ammonia gas

- It has a pungent choking smell.
- It is colourless.
- It is very soluble in water.

Chemical properties of ammonia gas

- Ammonia does not burn in air nor support combustion
- Ammonia turns moist red litmus paper blue. It is the common alkaline known.
- It reacts with hydrogen chloride gas to form white fumes of ammonium chloride.

Industrial manufacture of ammonia by Haber process

A mixture of nitrogen and hydrogen in the ratio of 1:3 by volume is purified, dried and introduced into the reaction chamber where they are reacted.

High pressure and low temperature give high yield. Most industries use temperatures of 400°C-500°C and pressure of 500 atmospheres. Finely divided iron catalyst is used to increase the rate of reaction. The catalyst is mixed with aluminium oxide to increase its efficiency.

A substance such as aluminium oxide which increases the efficiency of a catalyst is called a **promoter**.

Uses of ammonia

1. Large quantities of ammonia gas are used to make fertilizers.
2. Ammonia gas is used in the manufacture of nitric acid
3. Liquid ammonia is used as a refrigerant in large scale refrigerating plants and factories.
4. Ammonia solution is used as a solvent in cleaning agents e.g. laundries.
5. Ammonia is used in the manufacture of dyes, wood pulp, plastics and fibres such as nylon

Nitric acid is a mineral acid. It is a powerful oxidizing agent. Nitric acid can be prepared in laboratory as well as industrially.

Manufacture of nitric acid by Ostwald Process

Nitric acid is manufactured by a **catalyst** (platinum rhodium gauze) of ammonia.

The first reaction in this process produces nitrogen (II) oxide (nitrous oxide) and water. The heat produced in this step maintains the suitable temperature and therefore no further heating is required.

Nitrogen (II) oxide and steam are cooled by air from the compressor. Nitrogen (II) oxide is further oxidized by the air to produce nitrogen (IV) oxide.

Nitrogen (IV) oxide is then passed through water where it dissolves to form nitric acid (HNO_3) and nitrous acid (HNO_2).

Nitrogen (IV) oxide + water \longrightarrow nitrous acid + nitric acid



Excess air is passed to oxidise nitrous acid to nitric acid.

Uses of nitric acid

- Manufacture of nitrate fertilizers
- Manufacture of explosives e.g. TNT and dynamite
- Manufacture of dyes and drugs
- Manufacture of plastics
- It is used as an oxidizing agent in textile industries
- It is used in the refining of gemstones
- It is used in the identification of gold. The acid does not react with gold, hence it can be used to distinguish gold from those other substances which have similar appearance.

Manufacture of ammonium nitrate fertilizer

Ammonia + nitric acid \longrightarrow Ammonium nitrate



Determining percentage of nitrogen in ammonium nitrate

Calculate the percentage of nitrogen in ammonium nitrate

Solution

The formulae of ammonium nitrate is NH_4NO_3

$$\begin{aligned} \text{R.F.M. of } \text{NH}_4\text{NO}_3 &= (2 \times 14) + (1 \times 4) + (3 \times 16) \\ &= 80 \end{aligned}$$

$$\text{Fraction of N} = \frac{28}{80}$$

$$\begin{aligned} \% \text{ of N} &= \frac{28}{80} \times 100 \\ &= 35\% \end{aligned}$$

SULPHUR

Sources of sulphur

- Sulphur occurs naturally in volcanic regions
- Hydrogen sulphide found in petroleum, coal deposit and natural gas.

- Metal sulphides e.g. zinc blends (ZnS), iron pyrites (FeS_2), copper pyrites (CuS_2) etc etc.
- Metal sulphates e.g. gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$).

Rhombic sulphur or alpha (α - sulphur)

Rhombic sulphur is also known as **alpha** (α -sulphur). It is yellow crystalline solid with an octahedral shape. Rhombic sulphur is stable below 96°C . Above this temperature, needle-shaped long prism crystals of monoclinic sulphur are slowly formed.

Monoclinic sulphur or beta-sulphur (β - sulphur)

Monoclinic sulphur also known as beta-sulphur (β), is a yellow crystalline solid. It is obtained by allowing molten sulphur to solidify. Long term shaped needle like sulphur form on walls of the container. They can be easily separated from the molten sulphur by pouring off the latter. If monoclinic sulphur is kept at room temperature for few days, it gradually changes back to rhombic sulphur.

Transition temperature of the two allotropes is 96°C .

The other two non-crystalline forms of sulphur are known as **amorphous sulphur** and **plastic sulphur**. Plastic sulphur is a brown rubber-like liquid substance. It forms when the liquid is super cooled (suddenly cooled). The liquid gradually changes to a hard yellow opaque substance (rhombic sulphur).

Table: Changes when sulphur heating

Temperature	Observation
113°C	Sulphur melts to give a mobile amber liquid. As the temperature rises further; the liquid darkens.
180°C	The liquid becomes very thick or viscous. The colour changes gradually from red to black.
250°C	The liquid is so thick and viscous that it cannot flow.
Above 250°C	The colour becomes brighter and the liquid once more becomes mobile.

Physical properties of sulphur

- Sulphur is a yellow, non-metallic solid.

- It has low melting point.
- It is insoluble in water.

- Sulphur is soluble in organic solvents like methylbenzene and carbon disulphide.

Chemical properties of sulphur

- Sulphur reacts with most metals to form metal sulphides.
- Sulphur burns in air or oxygen with a bright blue flame
- Sulphur does not react with dilute acids. However, it is oxidized by hot concentrated oxidizing acids such as sulphuric acid and nitric acid.

Uses of sulphur

- In the manufacturing of sulphuric acid (H_2SO_4).
- For making gun powder and matches.
- For manufacture of fireworks.
- In the vulcanization (hardening) of rubber.
- For manufacturing germicides and fungicides.
- In the manufacture of medicines.
- In photographic development.
- For making bleaching agents.
- In making artificial hair colours or dyes.

Compounds of sulphur

Sulphur has a number of important compounds. These include:

- Fluorides of sulphur e.g. sulphur difluoride.
- Oxides of sulphur such as dioxide, sulphur trioxide and disulphur oxide.
- Hydrides of sulphur.
- Metal sulphides e.g. zinc blende (ZnS) and iron pyrites.

Uses of sulphuric dioxide

- It is used for bleaching, as fumigant and in food preservatives.

- Used in the contact process for the manufacture of sulphuric acid.
- It is used as disinfectant.

Uses of sulphur trioxide

It is mainly used in the industrial manufacture of sulphuric acid.

Manufacture of sulphuric acid by contact process

Stage 1: Production of sulphur dioxide

Sulphur dioxide in a furnace by burning sulphur or roasting metal sulphides in the air.

Stage 2: Conversion of sulphur dioxide to sulphur trioxide

Sulphur dioxide is converted to sulphur trioxide. Vanadium (V) oxide is used as a catalyst because it is cheaper and less easily poisoned by impurities than platinum.

Stage 3: Conversion of sulphur trioxide to sulphuric acid

Finally sulphur trioxide is absorbed in concentrated sulphuric acid forming an oily liquid product called **oleum**. Sulphuric acid is obtained by diluting oleum with water.

Uses of sulphuric acid

- To manufacture fertilisers such as ammonium sulphate.
- In the production of synthetic fibres
- In petroleum refining
- In car batteries
- In cleaning of metals
- To make soaps and detergents
- To make dyes, drugs, paints, explosives and many other compounds.

Manufacture of ammonium sulphate

- (iv) Ionic compounds are insoluble in organic solvents like benzene or propane.
- (v) They are usually crystalline solids. Ionic crystals are brittle.

Properties of covalent compounds

- They do not conduct electricity.
- They have low melting and boiling points. Some covalent compounds such as diamond and graphite have unusually high melting point. Others like large molecules plastics and carbohydrates decompose before their melting points are reached.
- Most are insoluble in water.
- They are soluble in organic solvents like benzene and propane.
- Most of them are gaseous or volatile liquids at room temperature and pressure.
- Most of them have scent.

Electrostatic forces in covalent and ionic compounds

When sugar and sodium chloride were heated in an experiment, sugar quickly melted and decomposed. Sodium chloride did not melt. This occurrence is due to the different strengths of the electrostatic forces between these two substances.

Covalent bonds too exhibit ionic bonding character to some extent. This phenomenon is due to unequal sharing of covalent bond electrons. The more electronegative atom tends to pull the shared electrons toward itself creating partial negative charge on it. The less electronegative atom become partially positive charged. This is what we referred to

ionic bonding character in covalent. This kind of ionic bonding is weak and does not cause much increase in the strength of the bonds in covalent compounds.

Solubility and volatility of ionic and covalent compounds

Substance that can dissolve in a particular solvent such as water is said to be soluble. Different substances may dissolve in a range of solvents or specifically in one. Some compounds do not dissolve at all.

Sodium chloride, an ionic compound, dissolves in water but does not dissolve in paraffin. Water being a polar liquid was able to break down the ionic bond of sodium chloride hence dissolving it. Ionic compound, readily dissolves in water and other polar solvents but do not dissolve in organic solvents such as paraffin.

Water molecules being polar have partial positive charge and partial negative charge. When ionic substance is placed in water, there is strong attraction between the ions of the solute and polar ends of the water molecules. This causes the breakup of the crystal lattice structure and the ions are then surrounded by water molecules, a process called **hydration**.

Covalent compounds, being non polar, readily dissolve in organic solvents. This is why cooking fat dissolved in acetone but would not dissolve in water.

A substance which evaporates readily (like acetone) at normal temperature and atmosphere pressure is said to be **volatile**. Some organic solvents are highly volatile while others are less volatile.

Sulphuric acid is reacted with aqueous ammonia to produce ammonia sulphate fertilizer.



Excess water is evaporated from aqueous ammonium sulphate solution. It is then cooled to crystallize.

PHOSPHORUS

Source of phosphorus

It occurs in many mineral rocks mainly phosphates. Its chief ore is calcium phosphate $\text{Ca}_3(\text{PO}_4)_2$. Phosphorus is also a major constituent of bones.

Allotropes of phosphorus

White phosphorus – It turns pale yellow on exposure to light. It is self-igniting and can catch fire very easily when exposed to air or oxygen. Therefore, it is usually kept under water. It gives off a very poisonous vapour. It is very unstable, most reactive and highly toxic of the two allotropes.

Red phosphorus – It is stable form at all temperatures. It does not catch fire easily unlike the white phosphorus. It may be formed by heating the white phosphorus to 250°C or by exposing white phosphorus to sunlight.

Physical properties

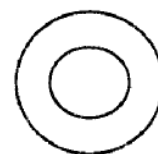
- Phosphorus is a yellow solid at room temperature.
- Melting point is 44°C .
- Boiling point is 280°C .
- Density 1.8gcm^{-3} .
- Does not conduct heat or electricity.

Chemical properties

1. Yellow phosphorus burns in oxygen to form two oxides: **phosphorus pentaoxide** and **phosphorous trioxide**.
2. It reacts with chlorine to form a mixture of two chloride compounds: **phosphorus trichloride** and **phosphorus pentachloride**.

Uses of phosphorus and its compounds

- The red phosphorus is used in the making of “safety” matches.
- Used in making detergents, toothpaste, baking powder and toothpaste.
- Phosphoric acid is used in the manufacture of inorganic fertilisers and soft drinks
- Used as an animal food additive



CHEMICAL BONDING AND PROPERTIES OF MATTER

Properties of Ionic compounds

- (i) Ionic compounds are electrolyte. They conduct electricity in solution or molten state.
- (ii) They have high melting and boiling points.
- (iii) They are generally soluble in water.

Electrical conductivity of covalent and ionic compounds

Aqueous solutions which conduct electric current are called **electrolytes**. Liquids such as acidified water, sodium chloride solution, mineral acids and copper (II) chloride solution conduct electricity.

These solutions have free ions which transfer (conduct) the electric current. The current decomposes the electrolyte as it passes through it.

Aqueous solutions such as sugar and urea do not conduct electricity. These are known as **non-electrolytes**. They exist as discrete molecules that is the covalent bond does not dissociate. This explains why the organic solvents such as ethanol and propanone are non-conductors of electricity.

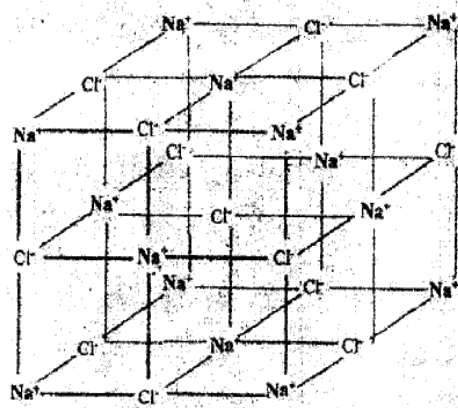
Metals are good conductors of electricity. The only non-metal that conduct electricity is graphite.

They conduct electric current by means of **free electrons**. Other substances conduct electric current in their molten (by means of ions) state like solid lead (II) bromide.

Structure of ionic and covalent compounds

Structure of ionic compounds

The ions are arranged in a regular pattern forming a **crystalline** structure. The crystal lattice consists of the oppositely charged ions with strong electronic attraction between them. There is alternating arrangement of cations and anions forming a **giant ionic** structure.



Effects of structures of ionic compounds on their properties

- (i) The strong electrostatic attraction between the ions in the structure requires large amount of heat to break down. This causes the high melting and boiling points of ionic compounds.
- (ii) They are soluble in water. They are also slightly soluble in some organic solvents such as ethanol, ether and benzene.
- (iii) They conduct electricity in molten or aqueous state but not in solid state
- (iv) The regular octahedral arrangement pattern of alternating anions and cations gives them the rigid crystalline structure.

Structure of covalent compounds

Covalent bonds exist in both elements (non-metallic) and molecular compounds. Examples of covalent elements are: sulphur, phosphorus, carbon and bromine

Covalent compounds include: ethanol, ethanoic acid, oils, ammonia and carbon dioxide.

Types of structures in covalent compounds

(i) Simple molecular structures

In simple molecular structures, the atoms forming the molecules are held together with strong covalent bonds while the molecules themselves are attracted to one another by weak intermolecular forces called **van der Waal's forces** or **hydrogen bonds**.

Examples of substances with simple molecular structure are carbon dioxide, oxygen, sulphur, water and iodine molecules.

Effects of simple molecular structure

- Low melting and boiling points.
- They are non-polar and generally insoluble in water.
- Since they exist as discrete molecules, they do not conduct electricity.
- The solids are soft due to the weak intermolecular force.

(ii) Giant atomic structure

It is possible for many atoms to link up together forming chains of covalent bonds. When this occurs, **giant molecules** are formed. Commonly known substances with giant atomic structure are diamond, silicon dioxide and graphite. The atoms linking up in the chain form very strong three dimensional structures except for graphite.

Graphite

The layers are joined by weak intermolecular forces, and the layers can easily slide over each other. This makes graphite a soft and slippery substance. The fourth electron is often delocalized. These delocalized electrons are mobile in the structure making a good conductor of electricity.

Graphite is used as lubricant, making of pencils, to reinforce metals and broken bones. It is also used as the positive terminals in dry cells and as electrodes in industry.

Diamond

The four carbon are tetrahedrally arranged and the covalent bonds interlace forming a tetrahedron structure. This makes the structure very strong and rigid. Diamond has no free mobile electrons.

Effects of giant atomic structure on the properties of the substance

- (i) The substances generally have high melting and boiling points.
- (ii) Because of the tetrahedral arrangement and interlacing of the bond, diamond is the hardest known substances. is used in making of drilling and cutting tools.
- (iii) Diamond and silicon dioxide have all their outermost electrons held in fixed positions. Therefore, they are very poor conductors of electricity.

Covalent bonds and dative bonds

In the case, where the shared pair of electrons are equally donated by the atoms involved, the bonding is referred to as **normal** or **pure covalent bond**.

When only one of the atoms involved contributes all the shared pair of electrons then, the bond formed is called **dative covalent bond** or simply as **dative bond**.

Example of dative bonding occurs in the formation of ammonium ion. Nitrogen atom in the molecule of ammonia having a lone pair of electrons donates to hydrogen (H^+) ion.

Polar and non-polar covalent bonds

When the covalent bond is formed between two identical atoms e.g. H-H or Cl-Cl, the electrons are shared equally. If the atoms are not the same e.g. H-Cl, H-O or N-H, the electrons contributed are usually not shared equally. There is an electron displacement towards the more electronegative atom. The **affinity** or attracting power of atom to attract an electron to form ion is called **electronegativity**.

The side to which the electrons are pulled more becomes partially **negative** while the other end becomes partially **positive**. If this happens, in a molecule, such as water, it is said to be **polar**.

The following are electronegativity values of some elements.

F	4.0	↑ Increase in electronegativity
O	3.5	
N	3.0	
C	2.5	
H	2.1	

The most electronegative elements is fluorine.

Intermolecular forces

These are forces of attraction holding together molecules of substances.

Examples of intermolecular forces

- Hydrogen bonding
- Van der Waals forces

Hydrogen bonding

The attraction between the partially negative atom and the partially positive hydrogen constitute the **hydrogen bond** which is usually represented by a dotted line.

Hydrogen bonding can also occur between different molecules e.g. ammonia and water.

Van der Waals forces

They increase with the mass of molecules, but between two molecules their effect is negligible. Their accumulative effects significantly influence some physical properties of covalent compounds e.g. organic compounds.

Effects of intermolecular forces on physical properties

1. Because hydrogen bonds have to be overcome during boiling, water has higher boiling point compared to many other covalent compounds of similar mass.
2. Surface tension of water is great due to hydrogen bonding among the water molecules.
3. Van der Waal's forces significantly contributes to the higher boiling and melting points for some elements moving down the group.
4. The presence of hydrogen bond between water molecules gives it the liquid state. Without these bonds, water would exist in gaseous state at room temperature and pressure.

Allotropy

Allotropy is the existence of an element in more than one form without change of state. The various forms are called **allotropes**. The allotropes exhibit different physical properties and sometimes even varying chemical properties.

Allotropes of oxygen

Oxygen and ozone are allotropes. Ozone has a triatomic molecule and oxygen is a diatomic molecule. Ozone can be made from oxygen. It can also be converted back to oxygen by heat.

Table: Comparison between oxygen and ozone

Oxygen (O ₂)	Ozone (O ₃)
1. Gas at room temperature	Gas at room temperature
2. Density 16 (H=1)	Density 24 (H=1)
3. Not attacked by heat.	Heat converts it to oxygen
4. Oxidising agent	Very strong oxidizing agent
5. Not poisonous	Poisonous
6. Does not react with potassium iodine solution.	Liberates iodine from potassium iodine solution.

Allotropes of carbon

Table: Comparison between diamond and graphite

Characteristics	Diamond	Graphite
Appearance	Colourless, transparent crystals that sparkle in light	Dark-grey, opaque and shining.
Density	Highest (3.5g/cm ³).	Moderate (2.3g/cm ³)
Hardness	Hardest natural substance known.	Soft and silvery.
Electrical conductivity	Does not conduct electricity	Conducts electricity

Carbon exists as diamond, graphite and amorphous carbon. Diamond and graphite are **crystalline**. The amorphous is largely non-crystalline. Amorphous carbon are in the form of charcoal, soot or coke.

Melting and boiling points of metals

Metals generally have high melting and boiling points. This is due to the strong **metallic** bonds between the fixed positive ions embedded in the 'sea' of delocalized electrons.

However, sodium (and the group I elements) is an exception with a low melting point of 98°C. This is because of the few delocalized electrons available for bonding decreasing the strength of the metallic bonds in its structure.

Density of metals

Metals have high densities because of their giant metallic structures except for sodium

and group I metal. The atoms of metals are packed together very closely.

Thermal and electrical conductivity of metals

There is an easy movement of electrons in metals. These electrons explain the great conducting power of solid or liquid metals for heat and electricity.

When one end of the metal is heated, the heat is quickly felt on the other end because the **kinetic** energy of the delocalized electrons is increased, they vibrate much faster, move and hit the nearby electrons resulting in the transfer of heat.

A flow of electrons in a metal atom pushes the electrons nearest to it so that the entry of one electron pushes out another electron to the opposite end of the metal. This leads to conduction of electrical current.

Ductility and malleability

- Most metals are **ductile** i.e. they are capable of being drawn out into wires. Ductility involves slipping of the layers of atoms over each other when the metal is pulled or hit.
- Most metals are also **malleable** i.e. they are capable of being hammered into thin sheets.

Uses of metals

Copper

- It is very ductile metal
- It is an excellent heat conductor
- It is used in arrow heads and other tools which are required not to shatter.
- Used in jewellery and ornaments making.

Aluminium

- It is used in overhead power lines
- It is also used in making bodies and other parts of automobile
- Because of its shiny metallic lustre, it is widely used in jewellery and as food wrapping material. Its shiny surface reflects heat well keeping the food hot for a long time.
- Because of its light and good heat conductivity, aluminium is used in making most cooking utensils.

Zinc

- It is used in batteries because of its good electrical conductivity.
- Many electrical appliances contain zinc.

Iron

- Because of its malleability, it is beaten into thin sheet in the making of roofing sheets, cutting tools such as pangas and hoes.
- Its good heat and electrical conductivity finds useful application in heating appliances such as iron box and cooking pans.

Alloys and their uses

An alloy is a mixture of two or more elements one of which is a metal. An alloy is obtained by mixing the base metal with a given element or elements to confer the desired quality of strength, malleability, weight or hardness among others.

Some of the important uses of alloys

- (i) Copper alloys are very tough. Hence they are used in tools, weapons jewellery and ornament making.
- (ii) Duralumin, in alloy of aluminium with copper and smaller quantities of magnesium, silicon and iron is very strong. Used in the manufacture of aircraft parts.
- (iii) Alloys of iron such as steel improves its tensile strength, lowering its brittleness, makes it tough and hard. These qualities are desirable in making machinery, cutting tools and iron forgery works.



STOICHIOMETRY

Chemical equations show precisely the reactants, products and their combining mole ratios. Stoichiometry of chemical reactions refers to an overall balanced equation.

Chemical Formulae Of Compounds

When elements react to form compounds, they do so by bonding. The capacity or strength of the atom to form the bonds is known as the **valency or combining power** of the element. The valency of an element is determined by the number of electrons in the outermost energy level. The electrons are also referred to as **valence electrons**.

Table: Valency and charges of ions of the group elements

Group	I	II	III	IV	V	VI	VII	VIII
Valency	1	2	3	4	3	2	1	0
Charge on the ion	1+	2+	3+		3-	2-	1-	0

There are some elements that display more than one valency in different reactions.

Table : Elements with multiple valencies

Element	Symbol	Valencies
Carbon	C	2 and 4
Lead	Pb	2 and 4
Iron	Fe	2 and 3
Copper	Cu	1 and 2
Sulphur	S	2, 4 and 6
Cobalt	Co	1, 2 and 3
Manganese	Mn	2 and 7
Chromium	Cr	3 and 6

Radicals and valencies

Radicals are groups of atoms that are bonded together to form complex ions. They cannot exist

in their own. They have either positive or negative charges.

Table : Radicals and their valencies

Radical	Charge on ion	Valency
Ammonium	NH_4^+	1
Hydrogen sulphate	HSO_4^-	1
Hydrogencarbonate	HCO_3^-	1
Hydroxide	OH^-	1
Nitrate	NO_3^-	1
Carbonate	CO_3^{2-}	2
Sulphate	SO_4^{2-}	2
Phosphate	PO_4^{3-}	3

Writing chemical formulae

Example 1

Write the formula for sodium sulphate.

Solution

Step 1

Write the symbols of the elements and radical.

Na SO_4

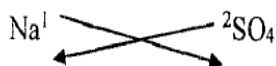
Step 2

Write the valency of the element and radical as superscripts. (simplify the values where possible)

Na^1 $^2\text{SO}_4$

Step 3

Exchange the valencies by writing them below the symbols as shown by the arrows.

**Step 4**

Write the symbols close together.

**Writing chemical equations**

When writing chemical equations ensure the following:

- (i) Correct symbols of the elements
- (ii) Correct state symbols of reactants and products.
- (iii) The equation is balanced.

Example 1

Write balance equation for the reaction of magnesium with oxygen to form magnesium oxide.

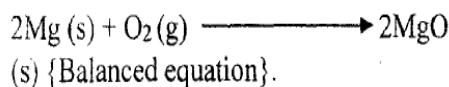
Solution

- i. First write the symbols and formulae of the compound formed.

$$\text{Mg (s)} + \text{O}_2\text{(g)} \longrightarrow \text{MgO (s)}$$
 {Unbalanced equation}.

- ii. Check number of the atoms of element in the equation if they are equal in the (reactants and products) sides. For example in above equation, are more atoms of oxygen on the reactants side than on the products side.

- iii. To balance the equation, we must identify suitable (small) whole numbers and values that will multiply atoms of the elements on both sides of the equation to make them equal. Write the number in front of each symbol. For example:



Note: 1. The plus (+) sign in equation does not mean "addition". It means "react with".

2. The arrow (\longrightarrow) is used to indicate formation of product.

3. Remember to include the state symbols.

(s) means to solid.

(l) means liquid.

(g) means gas.

(aq) means aqueous (substance dissolved in water)

Relative formula mass (R.F.M)

It is the average mass of one atom of the element with $\frac{1}{12}$ th of the mass of one atom of carbon -12.

Relative formula mass is the sum total of relative atomic masses of all the atoms in a compound.

Example 1

Find the relative formula mass (R.F.M) of the following compounds.

Calcium chloride

(RAM: Ca=40 Cl=35.5

Cu=63.5)

Solution

Step 1: Write the correct formula of the compound



Step 2: Add relative atomic masses of the elements in the compound.

$$\begin{aligned} \text{R.F.M of CaCl}_2 &= 40 + (35.5 \times 2) \\ &= 40 + 71 \\ &= 111 \end{aligned}$$

The mole

An Italian physicist Amedeo Avogadro came up with a unit for measuring or counting small particles such as atoms, ions, molecules and electrons, called **mole**.

It is a fixed and constant number. It is referred to as **Avogadro's number** or **Avogadro's constant**. This number is represented as $L=6.023 \times 10^{23}$.

Meaning of mole

The mole is the amount of substance that contains as many elementary particles as there are atoms in exactly 12.0g of the carbon-¹² isotope.

Relationship between mole and relative masses of substances

Molar mass

Relative molecular mass and relative formula mass of substance are taken as equivalent to one mole of the substance. If we weigh one mole of a substance, the mass obtained is referred to as molar mass. The molar mass M is the mass of one mole of a substance expressed in grams per mole (g/mol).

Molar mass of any element always contains 6.023×10^{23} particles though the molar masses of the elements are different. This is because of the variation in weights (relative atomic masses) of atoms.

Table : Molar mass and relative atomic mass (RAM) of some elements

Element	Relative atomic mass	Molar mass	Number of particles	Number of moles
Carbon 12	12	12	6.023×10^{23}	1
Hydrogen ¹	1	1	6.023×10^{23}	1
Aluminium	27	27	6.023×10^{23}	1
Sulphur	32	32	6.023×10^{23}	1
Lead	207	207	6.023×10^{23}	1

Table: Relative molecular mass (RMM) and relative formula mass (RFM) of some elements and compounds

Molecules/Compounds	Number of moles	Number of particles	Relative molecular/formula mass	Molar mass (grams)
H ₂	1	6.023×10^{23}	2	2
O ₂	1	6.023×10^{23}	32	32
Al ₂ (SO ₄)	1	6.023×10^{23}	342	342
CO ₂	1	6.023×10^{23}	44	44
CuCO ₃	1	6.023×10^{23}	124	124

Calculations involving moles

Example 1

Find the number of atoms in:

- 3 moles of iron
- 0.25 moles of iron
($L=6.023 \times 10^{23}$)

Solution

- 1 mole of Fe contains 6.023×10^{23} atoms
Let the number of atoms in 3 moles be x

$$\therefore x = \frac{3 \text{ moles} \times 6.023 \times 10^{23}}{1 \text{ mole}}$$

$$18.069 \times 10^{23}$$

$$1.8069 \times 10^{24} \text{ particles}$$

- (ii) 1 mole of calcium contains 6.023×10^{23} atoms

Let the number of atoms in 0.25 moles of Ca be x atoms

Solving for x ,

$$x = \frac{0.25 \text{ mole} \times 6.023 \times 10^{23}}{1 \text{ mole}}$$

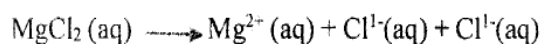
$$= 1.506 \times 10^{23} \text{ atoms}$$

Example 2

- (i) How many magnesium ions are there in 1 mole of magnesium chloride?

Solution

- (i) The formula of magnesium chloride is MgCl_2 . Ions are set free when the salt is dissolved or melted.



1 mole of magnesium chloride when dissolved in water produces 1 mole of Mg^{2+} and 2 moles of Cl^{-} ions.

But 1 mole = 6.023×10^{23} Mg^{2+} ions

\therefore 1 mole of MgCl_2 contains 6.023×10^{23} magnesium ions.

Example

If 54g of aluminium were obtained from aluminium ore, how many atoms aluminium were obtained? ($\text{Al}=27$; $L=6.023 \times 10^{23}$)

Solution

One of the ways to calculate the number of moles (N) is to use the formula

$$N = \frac{\text{Mass}}{\text{RAM or RMM}}$$

or you can use the proportion method also called the first principle method.

First principle method

1 mole of $\text{Al}=27\text{g}$

y mole of $\text{Al}=54\text{g}$

Cross multiply and solve for y

$$27\text{g} \times y = 54\text{g} \times 1 \text{ mole}$$

$$y = \frac{54\text{g} \times 1 \text{ mole}}{27\text{g}}$$

$$= 2 \text{ moles of aluminium}$$

1 mole of $\text{Al}=6.023 \times 10^{23}$ atoms

2 moles = $2 \times 6.023 \times 10^{23}$ atoms

Example

Convert the masses of the following substances to moles:

- (a) 0.12g of carbon atoms.

$$(\text{C}=12, \text{Mg}=24, \text{O}=16)$$

Solution

First principle method

- (a) 12g of carbon atoms contain 1 mole

$$0.12\text{g} \times x \text{ mole} = 12\text{g} \times 1 \text{ mole}$$

$$x = (0.12\text{g} \times \frac{1 \text{ mole}}{12\text{g}})$$

$$= 0.01 \text{ mole}$$

Therefore 0.12g of carbon atoms = 0.01 mole

Conversion of mass (in grams) of compounds to moles

Example 1

Suppose you have:

(a) 32g of sulphur dioxide

How many moles of each of the compounds do you have?

Solution

First principle method

First write the correct formula of the compound. Then work out the mass of mole i.e. molar mass of the compound.

$$\begin{aligned} \text{(a) 1 mole of sulphur dioxide, SO}_2 &= \\ 32 + (16 \times 2) &= 64\text{g} \\ x \text{ moles of sulphur dioxide, SO}_2 &= \\ 32\text{g} &= \\ (32\text{g} \times \frac{1 \text{ mole}}{64\text{g}}) &= \\ &= 0.5 \text{ mole} \end{aligned}$$

Conversion of moles to mass in grams

Example 1

A student bought x of charcoal (carbon). This was the same as 0.1 mole. Find the value of x . ($C=12$).

Solution

First principle method

$$\begin{aligned} 1 \text{ mole of carbon atoms} &= 12\text{g} \\ \therefore 0.1 \text{ moles of carbon atoms} &= x\text{g} \\ \text{Cross multiplying and solve for } x & \\ 1 \text{ mol } x \times \text{g} &= 0.1 \text{ mole} \times 12\text{g} \end{aligned}$$

$$\begin{aligned} \therefore x &= 0.1 \text{ mole} \times \frac{12\text{g}}{1 \text{ mole}} \\ &= 1.2\text{g} \end{aligned}$$

Conversion of moles of molecules or compounds to grams

Example 1

How many grams are there in 9 moles of hydrogen gas?

Solution

First principle method

R.A.M of $H=1$

1 mole of $H_2 = 1 \times 2 = 2\text{g}$

$$\begin{aligned} 9 \text{ moles of } H_2 &= x\text{g} \\ \times &= 2 \times 9 \\ &= 18\text{g} \end{aligned}$$

Molar gas volume

It is the volume occupied by one mole of the gas at standard temperature and pressure.

Experiments have shown that one mole of any gas at 25°C (298K) and 1 atmosphere pressure, occupy 24dm^3 . These conditions are known as **room temperature and pressure** (r.t.p).

It has also been shown that one mole of gas occupies 22.4dm^3 at **standard temperature** 0°C (273K) and **pressure** 1 atm (s.t.p). Because one mole of gas occupies the specified volume, the volume is known as the **molar volume** or **molar gas volume**.

Calculation involving molar volume

Example 1

How many moles of hydrogen gas are there in 24cm^3 of hydrogen at room temperature and pressure?

Solution

1 mole of hydrogen gas occupy 24 dm^3 at r.t.p
 $1000\text{cm}^3 = 1 \text{ dm}^3$

x moles of hydrogen gas occupy 0.024 dm^3 (24cm^3)

Cross multiply and solve for x

$$x = \frac{0.0024\text{dm}^3}{24\text{dm}^3}$$

$$= 0.001 \text{ mole}$$

Example 2

What volume in cm^3 does 0.01 moles of oxygen occupy at s.t.p?

Solution

1 mole of oxygen will occupy 22.4 dm^3 .

0.01 mole of oxygen will occupy $x \text{ dm}^3$

$$x = (22.4 \times \frac{0.01 \text{ mole}}{1 \text{ mole}})$$

$$= 0.224 \text{ dm}^3$$

$$\text{Volume in cm}^3 = 0.224 \times 1000 \text{ cm}^3$$

$$= 224 \text{ cm}^3$$

Water of crystallization

When saturated solutions of some salts e.g. copper (II) sulphate are gently heated and then slowly cooled large crystals will form. These crystals have definite shape with surfaces and edges. They incorporate water molecules within their structures that makes this possible. This water is called **water of crystallization**. Salts with water of crystallization are known as **hydrated salts**.

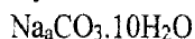
Hydrated ionic compounds have different number of water molecules within them. When hydrated salts are heated they lose their water of crystallization.

Solution

Elements	C	H
Composition of mass in g	12	4
Get moles by dividing mass by R.A.M of each element	$\frac{12}{12} = 1$	$\frac{4}{1} = 4$
Divide the moles by the smallest value	$\frac{1}{1} = 1$	$\frac{4}{1} = 4$
Mole ratio	1	4

Calculating percentage of water of crystallization**Example 1**

Calculate the percentage of water of crystallization in hydrated sodium carbonate,



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

Relative formula mass $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} = 286$

$$\text{Percentage of water} = \frac{10\text{H}_2\text{O}}{\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}} \times 100\%$$

$$= \frac{10 \times 18}{286} \times 100\% =$$

$$62.937\%$$

Empirical formula

Empirical formula of a compound shows the simplest whole number ratio of the moles of atoms of the different elements in a compound.

Example 1

When a sample of hydrogen was analyzed, it was found to contain 12g of carbon and 4g of hydrogen. Determine the empirical formula of this hydrocarbon. (C=12, H=1).

Calculating empirical formula using percentage composition**Example 2**

When a compound containing carbon and hydrogen was analysed, it was found to contain 75% of carbon and 25% of hydrogen. Find its simplest formula. (R.A.M of C=12, H=1).

Solution

Elements	C	H
Composition of mass in %	75	25
Get reacting moles by dividing % by R.A.M	$\frac{75}{12} = 6.25$	$\frac{25}{1} = 25$
Divide by the smallest value	$\frac{6.25}{6.25} = 1$	$\frac{25}{6.25} = 4$
Mole ratio	1	4

Therefore, empirical formula of the compound is CH.

Molecular formula (M.F)

The molecular formula of a compound shows the actual number of atoms of the various elements present in one molecule of that compound.

To determine the molecular formula of a compound, we need to know the empirical formula and the relative molecular mass (R.M.M) or relative formula mass (RFM) of the compound.

Molecular formula is obtained by multiplying the empirical formula by a whole number, n known as the multiplication factor.

$$M.F = (E.F) n$$

Example 1

Sodium metal burns in excess oxygen. It burns with a bright yellow flame and forms a yellow oxide that contains 59% sodium. The relative molecular mass of the oxide is 78. What is the molecular formula of the oxide? (RAM: Na=23, O = 16)

Percentage of oxygen is $100 - 59 = 41\%$.

Element	Na	O
Composition of mass in %	59	41
Get reacting moles by dividing mass by R.A.M	$\frac{59}{23} = 2.57$	$\frac{41}{16} = 2.56$
Divide by the smallest value to get mole ratio	$\frac{2.57}{2.56} = 1$	$\frac{2.56}{2.56} = 1$
Mole ratio	1	1

The empirical formula (E.F) of the oxide of the oxide is NaO.

The relative mass of empirical, $NaO = 23 + 16 = 39$

The molecular mass is given by

$$(NaO) \times n = 78$$

$$39 \times n = 78$$

$$n = \frac{78}{39}$$

$$= 2$$

Therefore molecular formula is $(NaO)_2$ which is Na_2O_2 .

Concentration of solutions

The term **concentration** refers to the **amount of solute dissolved in specific volume of solvent**.

Concentration is usually expressed in number of moles of solute dissolved per 1000cm^3 of solution.

$$1000\text{cm}^3 = 1\text{dm}^3 = 1 \text{ litre.}$$

When one mole of a solute is dissolved in water and the volume of the solution made up to 1dm^3 solution is said to be a **molar solution**.

A molar solution contains **1 mole of solute per dm^3 solution**. A concentration of one mole per cubic decimeter is usually written at **1 mol/dm^3 or 1 mol dm^{-3} or 1M** . Concentration expressed this way is also referred to as molarity of the solution. The molarity of solution is abbreviated as M.

Expressing concentration of solutions

The three common ways of expressing concentration of solutions are:

- (i) Grams per volume.
- (ii) Moles per volume, usually in dm^3 .
- (iii) Concentration in percentage.

Concentration in grams per dm^3

The concentration (or molarity) can also be expressed in grams per cubic decimeter (g dm^{-3}). We may first calculate the concentration in mol/dm^3 and then convert to g/dm^3 .

Example 1

53g of anhydrous sodium carbonate were dissolved in 2 dm^3 . Calculate the concentration in g/dm^3 . (Na=23, C=12, O=16).

Solution

Using first principle method

Formula of sodium carbonate is Na_2CO_3

Step 1: Work out relative formula mass of Na_2CO_3
 $= (2 \times 23) + (1 \times 12) + (3 \times 16) = 106$

Step 2: Calculating the number of moles as follows:

$$1 \text{ mole of } \text{Na}_2\text{CO}_3 = 106\text{g}$$

Cross multiply and solve for x

$$\begin{aligned} x &= (53\text{g} \times \frac{1 \text{ mol}}{106\text{g}}) \\ &= 0.5\text{moles} \end{aligned}$$

Step 3: Calculate the concentration in mol/dm^3 using the number of moles.

0.5 mole was dissolved in 2 dm^3 .

y moles were dissolved in 1 dm^3 .

Cross multiply and solve for y

$$y = (0.5 \text{ mole} \times \frac{1 \text{ dm}^3}{2 \text{ dm}^3})$$

$$= 0.25 \text{ mol dm}^{-3}$$

Step 4: Finally, convert mol dm^{-3} to g dm^{-3} as follows:

$$1 \text{ mole of } \text{Na}_2\text{CO}_3 = 106\text{g}$$

$$\begin{aligned} 0.25 \text{ mole of } \text{Na}_2\text{CO}_3 &= 0.25 \text{ mole} \times \frac{106\text{g}}{1 \text{ mole}} \\ &= 26.5\text{g} \end{aligned}$$

Therefore, concentration expressed in g per $\text{dm}^3 = 26.5\text{g dm}^{-3}$

Concentration in moles per volume (mol/m^3)

- A solution contains 2 moles in 1000cm^3 (1 dm^3), has a concentration of 2 mol/dm^3 or simply 2M .
- A solution which contains 0.5 moles of solute dissolved in 1 dm^3 solution is expressed as 0.5 mol dm^{-3} or 0.5M .

Concentration expressed as a percentage

Concentration of solutions can also be expressed as a **percentage**. This is very common with strong mineral acids. It is indicated on the bottle label, for example 50% concentration or 98% concentration.

Preparation of standard solutions

A solution whose concentration is known is called a **standard solution**.

Suppose your teacher asks you to prepare 250cm^3 of 1 mol dm^{-3} (1M) sodium hydroxide for a chemistry lesson. How would you prepare this solution?

How to prepare a standard solution

Step 1: Work out required moles.

1 M means 1 mole in 1000cm^3 .

Therefore,

If $1000\text{cm}^3 = 1 \text{ mole}$

$$250 \text{ cm}^3 = x \text{ moles}$$

Cross multiply and solve for x

$$x = 1 \text{ mole} \times \frac{250 \text{ cm}^3}{1000 \text{ cm}^3} \\ = 0.25 \text{ moles}$$

Step 2: Work out R.F.M of sodium hydroxide (NaOH)

$$\text{R.F.M of NaOH} = 23 + 16 + 1 \\ = 40$$

$$1 \text{ mole of NaOH} = 40$$

$$0.25 \text{ moles of NaOH} = y \text{ g}$$

Cross multiply and solve y

$$\therefore y = (40 \text{ g} \times \frac{0.25}{1 \text{ mole}}) \\ = 10 \text{ g}$$

Step 3: Now weigh 10g of sodium hydroxide pellets and put in a 250 cm^3 volumetric flask, dissolve in distilled water and make the solution up to 250 cm^3 mark with distilled water.

Concentration and dilution

When one solvent e.g. water is added to a given concentrated solution, it is said to being diluted. Now if the initial concentration is taken as C_1 and the initial volume as V_1 the number of moles of solute is obtained as:

$$C_1 \times V_1 = \text{Moles}$$

For the diluted solution, let concentration be C_2 and the new volume by V_2 the number of moles is obtained as follows:

$$C_2 \times V_2 = \text{Moles}$$

Since number of moles of solute remain the same in the initial solution and diluted solution.

Then

$$C_1 \times V_1 = C_2 \times V_2 \text{ or}$$

$$C_1 V_1 = C_2 V_2$$

This relationship is called the **dilution law**.

Examples 1

How much water must be added to 5 cm^3 of 6.0 M HCl to dilute the acid to 0.1 M ?

Solution

$$C_1 = 6.0 \text{ M} \quad V_1 = 5.0 \text{ cm}^3$$

$$C_2 = 0.1 \text{ M} \quad V_2 = ?$$

$$\text{Using } C_1 V_1 = C_2 V_2$$

$$V_2 = \frac{C_1 V_1}{C_2}$$

$$V_2 = 6.0 \times \frac{5.0}{0.1}$$

$$= 300 \text{ cm}$$

$$\text{Therefore volume of water to be added } (300-5) \text{ cm}^3 \\ = 295 \text{ cm}^3$$

Acid-base titration

Acid reacts with a base to form salt and water. This reaction is referred to as **neutralization** reaction.

Neutralization reactions can be done very precisely by a process known as **titration** in a laboratory.

Titration is the process of adding a solution from a burette into a standard solution. An indicator is used to show the point at which the reaction becomes complete. The point at which a reaction is complete is called the **end-point** or **equivalent point**.

The volume of the solution being added from the burette is called **titre**. The volume transferred by the pipette into the conical flask is known as an **aliquot**.

Calculating percentage yield of a reaction

The actual amount of product formed during reactions can be expressed in percentage.

Percentage yield is calculated using the following formula:

Percentage yield

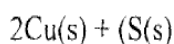
$$= \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

Example 1

In an experiment, 2.4g of copper metal were heated with excess sulphur to yield 2.25g of copper (I) sulphide - $\text{Cu}_2\text{S}(\text{s})$. What is the percentage yield of this reaction?

(Cu = 63.5, S=32)

Solution



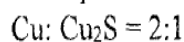
From the equation,

$$\text{Mole of copper that reacted} = \frac{\text{mass of copper given}}{\text{molar mass of copper}}$$

$$= \frac{2.4}{63.5}$$

$$= 0.0378 \text{ mole}$$

Mole ration in the equation can be used to calculate the expected number of moles of the product.



$$\text{Expected moles of Cu}_2\text{S to be formed} = \frac{0.0378}{2} \text{ mole}$$

$$0.0189 \text{ mole}$$

$$\text{R.M.F of Cu}_2\text{S} =$$

$$(2 \times 63.5) + 32 = 159$$

$$1 \text{ mole of Cu}_2\text{S} =$$

$$159\text{g}$$

$$\text{Expected yield of Cu}_2\text{S} =$$

$$159\text{g} \times 0.0189 = 3.00519\text{g}$$

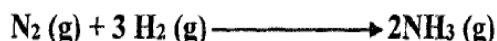
$$\text{Percentage yield of Cu}_2\text{S} =$$

$$\frac{2.25\text{g}}{3.0051\text{g}} \times 100\%$$

$$74.87\%$$

Limiting Reagent

The reactant that is completely used up in reaction while the other reactant is still available is called the **limiting reagent**. The unfinished reactant (reagent) is known as excess **reagent/reactant**.



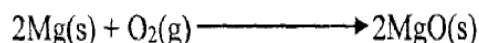
According to the stoichiometric equation above, one mole of nitrogen gas reacts with three moles of hydrogen gas to yield two moles of ammonia gas.

Now, if one mole of nitrogen and three moles of hydrogen are reacted, it means one mole of hydrogen would be left unreacted.

In this case the amount of product formed by the reaction is limited/controlled by the **amount of the limiting reagent (reactant)**. The limiting reagent in this reaction is nitrogen and the excess reagent is hydrogen.

Example 1

Magnesium and oxygen react to form magnesium oxide according to the equation.



If 6g of magnesium and 2g of oxygen are available for a reaction, determine the limiting reagent, the excess reagent and the mass of magnesium oxide formed.

(O = 16, Mg = 24)

Solution

$$\begin{aligned} \text{Moles of Mg used} &= \frac{\text{mass of Mg given}}{\text{molar mass of Mg}} \\ &= \frac{6\text{g}}{24\text{g}} \end{aligned}$$

$$= 0.25 \text{ mole}$$

$$\begin{aligned} \text{Moles of O}_2 \text{ used} &= \frac{2\text{g}}{32\text{g}} \\ &= 0.0625 \text{ moles} \end{aligned}$$

From the equation, the mole ratio is:

2 moles of Mg require 1 mole of O_2 for a complete reaction.

0.25 moles of Mg would require

$0.25/2 = 0.125$ moles of oxygen.

The amount of oxygen available is 0.0625 mole, which is less than the required amount of 0.125 mole. Therefore, oxygen is the limiting reagent.

If 0.0625 moles of O_2 is used; mole of Mg required is.

$(0.0625 \times 2) = 0.125$ moles of Mg.

However, amount of Magnesium available is 0.25 mole, which means magnesium is in excess, by $(0.250 - 0.125) = 0.125$ mole. Magnesium is therefore the excess reagent.

Moles of MgO produced is equivalent to moles of Mg used = 0.125 mole.

R.F.M of MgO = $24 + 16 = 40$.

Mass of magnesium oxide formed = $N \times RFM$
 $= 0.125 \times 40 = 5.0g$



CHEMICAL REACTIONS

Heats of reactions

When a chemical reaction takes place, either heat is involved for the reactants to react or heat is released during the reaction to produce the products. These two types of heat are the one that are called **heats of reaction**.

When potassium nitrate is dissolved in water, heat energy is absorbed from its surroundings.

This is an endothermic reaction. Loss of energy is a result of interaction between a system and its surrounding. The beaker and its content is the system.

When sodium hydrogen is dissolved in water, heat energy is released to the surroundings. The surroundings is the thermometer, the beaker, the air around and your palm. This is the reason why your palm feels warm. The temperature of the solution is higher than that of the water and pellets before dissolution. This is an **exothermic reaction**.

Exothermic and endothermic reactions

An exothermic reaction is one that releases heat and cause the surrounding temperature to rise while an endothermic reaction is one that absorbs heat and causes the surrounding temperature to fall.

Table : Some examples of exothermic and endothermic reactions

Exothermic process	Endothermic process
Formation of ice cubes (freezing water).	Melting ice cubes.
Condensation of rain water vapour.	Evaporation of water.
A burning candle.	Melting solid salts.
Mixing water and strong acid.	Mixing water and ammonium nitrates.
Adding water to anhydrous salt.	Losing water from a hydrated salt.
Combining atoms to form a molecule.	Splitting a molecule to individual atoms.
Neutralization reaction.	Photosynthesis.
Ionization of atoms.	Baking bread.

Enthalpy of neutralization

The term **enthalpy** refers to heat content of substance.

An experiment can be conducted where an acid of a known volume, concentration and temperature is poured into a beaker. A base of known volume, concentration and temperature can then be added

to the acid and the final temperature of the two recorded.

Temperature rise would be seen in these neutralization reactions. This shows that these reactions are exothermic reactions. ΔH is therefore negative.

Table : Results of temperatures reading in neutralization reaction

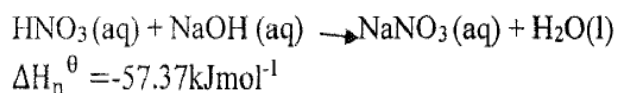
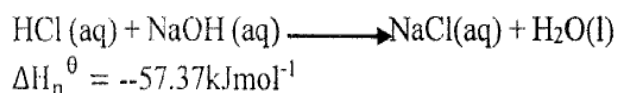
	°C	K
Temperature of the acid, T_1		
Temperature of the alkali, T_2		
Average temperature of the two solutions, $\frac{T_1+T_2}{2} = T_3$		
Highest temperature of solution reached, T_s		
Temperature change, $(T_s - T_3) = \Delta T$		

It is possible to calculate the enthalpy change, using the known volumes.

Table : Specimen results

	°C	K
Temperature of hydrochloric acid, T_1	22.5	295.5
Temperature of sodium hydroxide, T_2	22.0	295.0
Average temperature of the two solutions, T_3	22.25	295.25
Highest temperature reached, T_s	35.5	308.5
Temperature change, $\Delta T = (T_s - T_3)$	13.25	13.25

For neutralization reactions involving strong acids and strong bases, the enthalpy change in the reactions is almost a constant value.



Whenever enthalpy (heat changes) are measured under standard conditions, thus temperature as 25°C (298K) and standard pressure as 1

atmosphere, they are called **standard enthalpies** of the reactions. The special symbol θ is used as a superscript for ΔH .

Examples of standard enthalpies of reactions

ΔH_f^θ the standard enthalpy of formation.

ΔH_c^θ the standard enthalpy of combustion.

ΔH_n^θ denotes the standard enthalpy of neutralisation.

Energy level diagrams for exothermic and endothermic reactions

$$\begin{aligned}\text{Total energy required} & \checkmark \\ & = -860 \text{ kJ mol}^{-1}\end{aligned}$$

Since bond formation is an exothermic reaction, the value should be written as -860 kJ mol^{-1}

Subtracting the two total values, $-860 - +675 = -185 \text{ kJ} = \Delta H$

Hence the reaction is exothermic. There are two moles in the product section.

$$\Delta H = \frac{-185}{2} = -92.5 \text{ kJ mol}^{-1}$$

The energy stored in the bonds of the molecules of a substance is called **potential energy**. When hydrogen atoms combine to form hydrogen molecules, energy is released in form of heat energy. The reaction is **exothermic**. Due to the energy released, the hydrogen molecules formed contain less energy and are said to be stable.

Due to energy absorbed when separating the molecules, free atoms possess more energy than the molecules. Therefore, free atoms are highly unstable. This is why hydrogen does not exist as separate atoms.

The energy required to break covalent bonds between atoms have been determined experimentally, hence are known values. These energies are called **bond dissociation** energies.

The covalent bond between the hydrogen atoms is the strongest. This is because the hydrogen atoms are very small and the atomic radius is very small, therefore, the nuclei attract the shared pair of electrons very strongly. Hence, the bond requires a lot of energy to break.

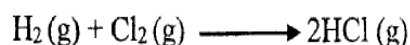
Enthalpy of formation

Enthalpy of information is the enthalpy change that occurs when one mole of a substance is formed from its elements under standard conditions.

Enthalpy of formation of hydrogen chloride

Determine of bonds to be broken and the new bonds to be formed.

Step 1: First write the reaction equation as follows:



Or



Bonds broken new bonds formed

H – H bond dissociation energy given is 436 kJ
 $= 1 \times 436 = +436 \text{ kJ}$

Cl – Cl bond dissociation energy given is 242 kJ
 $= 1 \times 242 = +242 \text{ kJ}$

Total energy needed $= 436 + 242 = +678 \text{ kJ}$

To total energy released during bond formation is obtained as follows:

H – Cl bonds dissociation energy is 431 kJ each
 $= 2 \times 431 = -862 \text{ kJ}$

Total energy released $= -862 \text{ kJ}$

We got more energy when making new bonds than spent when breaking bonds in this case. Heat of reaction, $\Delta H_r = \text{energy difference}$.

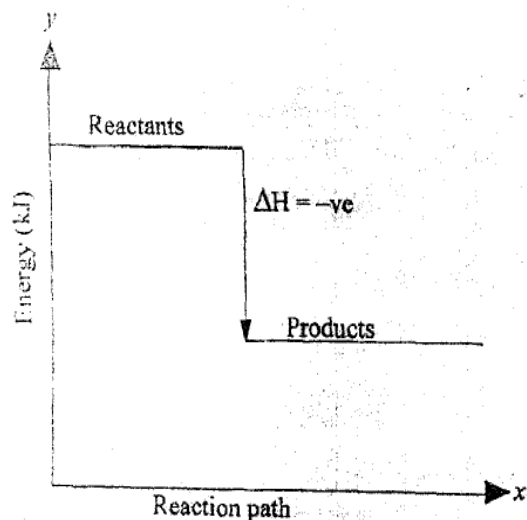
$$\begin{aligned}\therefore \Delta H_r & = 862 - 678 \\ & = -184 \text{ kJ}\end{aligned}$$

$\therefore 184 \text{ kJ}$ of energy is released when 2 moles of hydrogen chloride are formed.

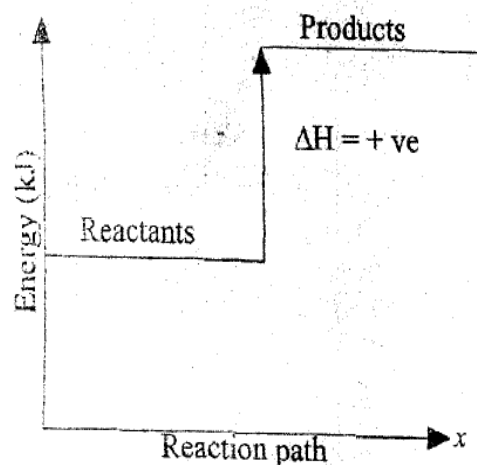
Therefore, the enthalpy of formation for HCl is,
 $\frac{-184 \text{ kJ}}{2} = -92 \text{ kJ mol}^{-1}$

An energy diagram is a graphical representation of the heat change involving products and reactants after the reaction

If the heat change is **negative** (-ve), then that particular reaction is **exothermic**.



If the heat content of the products is higher than that of the reactants then the heat change will be a **positive** (+ve) value and the reaction is **endothermic**.



Energy changes in Bond breaking and formation processes

Bond breaking absorbs energy therefore it is an **endothermic** process.

In bond formation, energy in form of heat is released, therefore it is **exothermic** process. Enthalpy changes in chemical processes can be determined through experiments.

The difference between the energy absorbed and energy released is the overall **enthalpy change** of reaction.

Bond energies calculation

Every chemical bond has a particular energy associated with it. Known bond energies would be used to calculate the overall energy changes of reactions.

Consider the following general reaction where diatomic molecular of X_2 react with diatomic molecules of Y_2 .



Table : Specimen bond energies

Bond	Energy required or released (kJ mol ⁻¹)
X-X	435
Y-Y	240
X-Y	430

To break bond between X-X, the energy required
= 435 kJ mol⁻¹

To break bond between Y-Y, the energy required
= 240 kJ mol⁻¹

Total energy required
= + 675 kJ mol⁻¹

Since bond breaking is endothermic, the value should be written as +675 kJ mol⁻¹

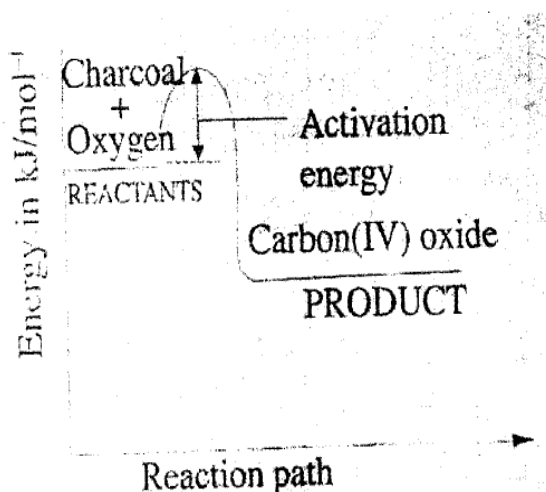
Forming a bond between X-Y, release energy
= 430 kJ mol⁻¹

Forming another bond between X-Y, release energy = 430 kJ mol⁻¹

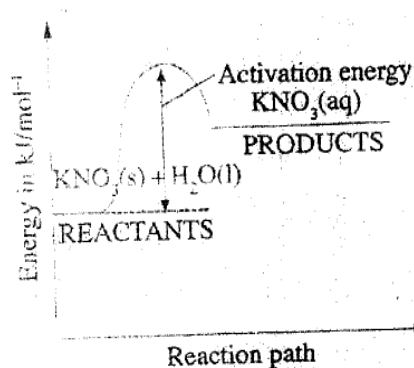
Energy level diagrams showing activation energy

Activation energy (E^a) is the minimum amount of energy required for a reaction to take place.

Without activation energy, reaction cannot proceed. Lighting fire wood before you cook provides energy needed to initiate the reaction between carbon and oxygen, which provides heat used during cooking i.e.



The figure below shows that dissolution of potassium nitrate is an endothermic process and activation energy is needed for reaction to begin.



The general formula of alkanols is ROH , where R is an **alkyl** group. Alkyl group is a name given to a

ALKANOLS

Homologous series is a group of molecules with the same general formula and the same functional group.

Characteristics of Alkanols (alcohols) as a homologous series

- Members of this series contain -OH as the functional group.
- All the names of alkanols end with -ol .
- They have a general formulae of ROH or $\text{C}_n\text{H}_{2n+1}\text{OH}$.
- Each member differs from the preceding one by a $\text{-CH}_2\text{-}$ group (methylene group).
- They show gradual change in physical properties such as melting points and boiling points as the hydrocarbon chain gets bigger.
- They have similar physical and chemical properties.

Functional groups of alkanols

The **functional group** is an atom or combination of atoms which gives an organic molecule its distinctive and characteristic properties. Alkanols are saturated organic compounds with a functional group -OH (hydroxyl group).

Any molecule which contains the functional group -OH is a member of the alkanol family.

General formula and nomenclature of alkanols

group of atoms remaining where one of the hydrogen atom of an alkane is removed.

We can also use the general formula $\text{C}_n\text{H}_{2n+1}\text{OH}$.

Molecular and structural formula of alkanols

Molecular formula shows the actual number of atoms of each element in a molecule, with any functional group present.

Structural formula shows the arrangement of atoms carbon by carbon, with the attached hydrogen and functional groups.

Skeletal formula which show the bonds of the carbon skeleton only, with any functional groups but, the hydrogen and carbon atoms are not shown.

Condensed formula and the skeletal formula

Name	Condensed formula	Skeletal formula
Methanol	CH ₃ OH	
Ethanol	CH ₃ CH ₂ OH	
Propan-1-ol	CH ₃ CH ₂ CH ₂ OH	
Butan-1-ol	CH ₃ (CH ₂) ₃ CH ₂ OH	
Pentan-1-ol	CH ₃ (CH ₂) ₄ CH ₂ OH	

Classification of alcohols

Alcohols can be classified as primary, secondary or tertiary. The classification is based on the location of the hydroxyl group.

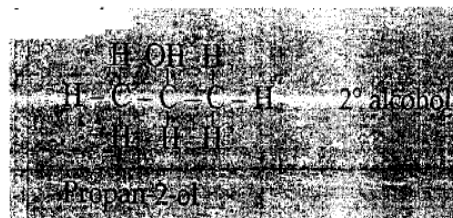
The structural differences between primary, secondary and tertiary alcohols cause differences in their chemical behaviours. Alcohols molecules may contain one, two, three or many hydroxyl groups.

Primary alcohol

If the hydroxyl group is attached to a carbon atom attached to only one other carbon atom, then the alcohol is a primary alkanol (1°).

Secondary alcohol

When the hydroxyl group is attached to a carbon atom is bonded to two other carbon atoms, then the alcohol is called a secondary (2°).



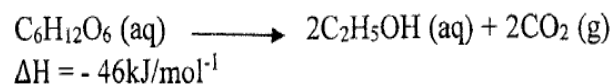
Tertiary alcohol

If the carbon atom bonded to the hydroxyl group is also bonded to three other carbon atoms, then the alcohol is tertiary alcohol (3°).

Preparation of ethanol

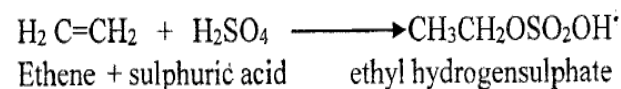
(a) Preparing ethanol by fermentation of sugars

The reactions whereby sugars are converted to alcohol and carbon dioxide are referred to as **fermentation reactions**. Ethanol is prepared in the laboratory by the fermentation of glucose or any other sugar solution using enzymes like yeast.



(b) From reaction between ethene and steam

Alcohols are produced industrially by hydrating alkenes in the presence of an acid catalyst such as concentrated sulphuric acid. Cold concentrated acid reacts with ethene as shown below:

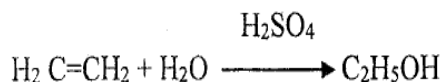


When cold water is added and the mixture warmed the ethyl hydrogensulphate is hydrolyzed to form ethanol.

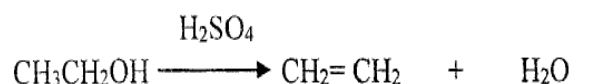


Ethyl hydrogensulphate \rightarrow Ethanol + Sulphuric acid

Overall reaction



The reactions are reversible, because alkanols will undergo elimination reaction.



Ethanol Ethene Water

Secondary and tertiary are more easily dehydrated than primary alcohols.

Properties of alkanols

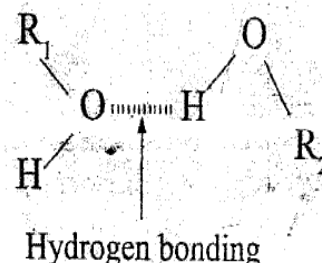
Generally alkanols have neutral pH. They are poor conductors of electricity. The first three are miscible in water. The rest are immiscible. Table 5.6 shows the trends in physical properties of alkanols.

Physical properties of alkanols

- (i) The melting point and boiling points of alkanols increase down the homologous series.
- (ii) Solubility of alkanols decreases down the group.
- (iii) Density of alkanols increases down the group.
- (iv) The higher alkanols containing 4 to 10 carbon atoms are viscous.
- (v) Volatility decreases down the group.

The melting and boiling points of alkanols increase with **increase in relative molecular mass**. As the molecular mass increase, the intermolecular forces of

attraction (**hydrogen bonding** and **Van der Waals forces**) between molecules increase. As these forces increase, more energy is needed to overcome them. This increases the melting and boiling points.



Because of these hydrogen bonds, alkanols have higher melting points than hydrocarbons of the same number of carbon atoms.

The causes of ethanol's anomalous properties is hydrogen bonding that forms between OH groups of adjacent ethanol molecules, giving ethanol (alkanols) relatively high intermolecular forces therefore its low volatility.

Hydrogen bonding between OH groups also has an effect on the viscosity of alcohol particularly those with more than one OH group in their molecule.

For example ethanol has viscosity of $1.06 \times 10^{-3} \text{ Nsm}^{-2}$ at 298°K , about the same as water, but propane -1,

2, 3 -triol (glycerin) is very thick and sticky with viscosity of $942 \times 10^{-3} \text{ Nsm}^{-2}$ at the same temperature, because of interaction between its molecules which carry three OH groups.

Chemical properties of alkanols

a. Reacts with air or oxygen

Ethanol + oxygen \longrightarrow Carbon dioxide + Water

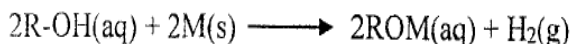


b. React with sodium metal

Ethanol + sodium \rightarrow Sodium ethoxide + Hydrogen



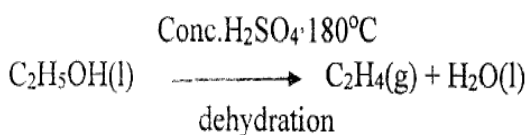
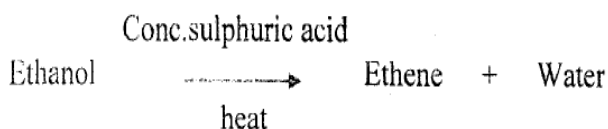
The general equation of the reaction of alkanols with metals is:



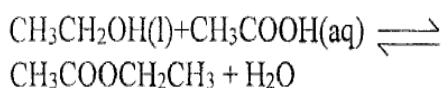
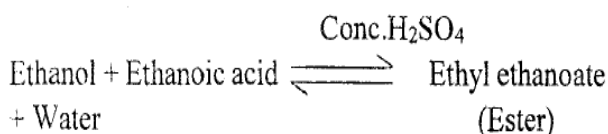
Where, R – alkyl group of the reacting alkanols.

M – metal such as potassium or sodium.

c. Concentrated sulphuric acid reacts with ethanol at 180°C to form ethene and water. This is a **dehydration** reaction.

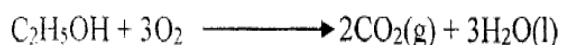


d. Alkanols react with alkanoic acids to form compounds called **esters** and water.

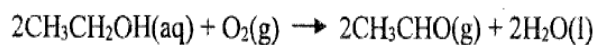
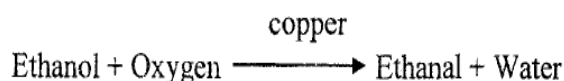


Oxidation of ethanol

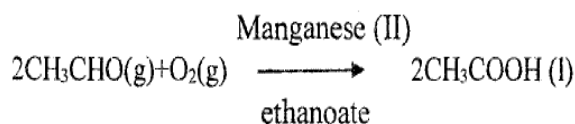
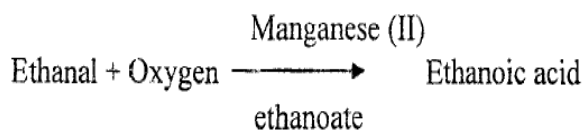
a. Oxidation of ethanol using atmospheric oxygen



Oxidation can also be carried using copper catalyst in the first stage.



In a second stage, manganese (II) ethanoate is used as a catalyst.

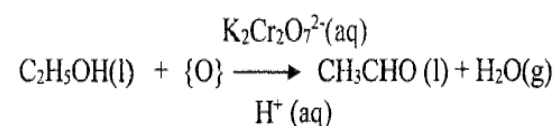
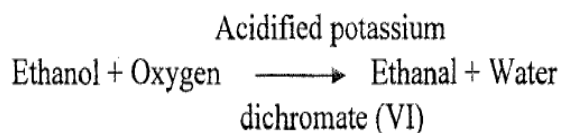


b. Oxidation of ethanol using oxidizing agents

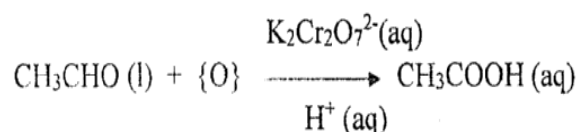
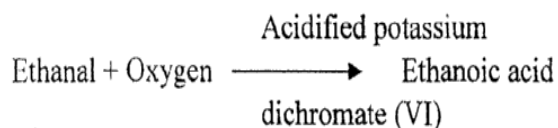
When acidified potassium dichromate (VI) ($K_2Cr_2O_7$) is added as an oxidizing agent to ethanol, the colour changes from orange to green.

When acidified potassium manganate (VIII) ($KMnO_4$) is added to ethanol, the purple colour changes to colourless. The oxidation of ethanol is a two-step process.

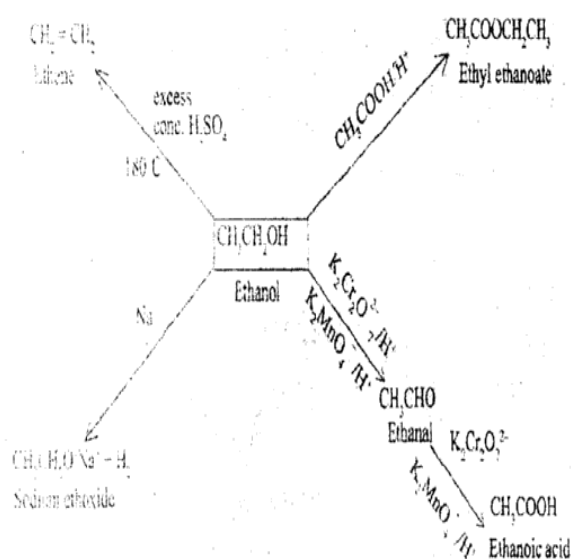
1. Ethanol is first converted to **ethanal**. This is a slow process.



- Then rapid conversion of ethanol to ethanoic acid.



Summary of reaction of ethanol



Uses of ethanol

- Used as a solvent e.g. it is a good solvent for iodine.
- Used in the manufacturing of varnishes, ink, glues and paints.
- Used in the manufacture of deodorants, perfumes, etc.
- Ethanol is used in alcoholic drinks.
- As an antiseptic in specified concentration.

- As a fuel when it is blended with gasoline to form gasohol.

Dangers of excessive consumption of alcoholic drinks

- Alcoholism (addiction) which can destroy physical and emotional health.
- Causes of some cancers are directly attributed to excess alcohol consumption which include: liver cancer, mouth cancer, pharynx, esophagus and pancreatic cancers.
- Cardiac problems, people who drink excessively can develop high blood pressure. Heart muscle damage as a result of drinking too much alcohol can cause fatal heart failure or stroke.
- Pregnant women can suffer miscarriages if they continue to take alcohol or give birth to babies that have fetal alcohol syndrome.
- Alcohol as depressant can influence one to commit suicide or suffer serious injury.
- Liver cirrhosis, this is disease associated with long-term alcohol abuse.
- Risky sexual behaviours leading to unwanted pregnancy, HIV and AIDS and other sexually transmitted infections.
- Loss of employment leading to poverty.

ALKANALS AND ALKANONES

Alkanals and alkanones molecules are characterized by the presence of a carbonyl group. A carbonyl group consists of an oxygen atom bonded to a carbon atom. $\text{C}=\text{O}$. It has an ability of undergoing a wide variety of reactions. This group is present in alkanals (**aldehydes**) and alkanones (**ketones**).

Alkanals

The general formula of alkanals is $RCHO$. Where R represents an alkyl group. In methanol, R represents hydrogen (H).

The other general formula of alkanals is $C_nH_{2n}O$.

Alkanals have alkyl carbon chains (R-) attached to the carbonyl functional group except in the case of methanal (the first member of the group) where the hydrogen atom attaches to the functional group to form methanal. The functional group in alkanals is always at the end of the carbon chain.

Nomenclature of Alkanals

Name	Condensed formula	Structural formulae	Molecular formulae
Methanal	$HCHO$	$\begin{array}{c} O \\ \\ H-C-H \end{array}$	$HCHO$
Ethanal	CH_3CHO	$\begin{array}{c} H & O \\ & \\ H-C & -C-H \\ \\ H \end{array}$	CH_3CHO

Sources of alkanals

- Traces of alkanals are found in essential oils which often contribute to their favourite odours.
- Most sugars are derivatives of alkanals.
- Traces of methanal are present in the air. This is as a result of photochemical oxidation of hydrocarbons.
- Oxidation of primary alcohols under controlled conditions.

Physical properties of alkanals

- All alkanals are liquid except methanal which is gas.
- Alkanals are colourless and have a characteristic smell.

- Alkanals solubility decreases down the homologous series.
- Alkanals do not have ability to form hydrogen bonds like alcohols.
- The melting and boiling points of alkanals generally increase with increase in RMM.

Physical properties of ethanal

- Ethanal is a colourless liquid.
- Has a pungent fruity apple colour.
- Soluble in water.
- Has a low boiling point of $20^\circ C$.

Chemical properties of Alkanals

1. It burns with a blue flame in air or oxygen to produce carbon dioxide and water.
2. Ethanal is a neutral substance
3. Ethanal oxidizes to ethanoic acid in the presence of Acidified potassium manganate (VII) and acidified potassium dichromate (VII).
4. Ethanal can undergo addition reaction with hydrogen cyanide (HCN) to form 2 hydroxy propano nitrile. Ethanal also undergoes addition reaction when shaken with saturated solution of sodium hydrogensulphite ($NaHSO_3$).
5. Alkanal can also undergo addition reaction with hydrogen in presence of nickel metal catalyst or platinum metal catalyst to form primary alcohols. Addition of hydrogen to a substance is called **reduction**.
6. **Polymerization**-alkanals can be readily converted to a variety of addition polymers just like alkenes.

Test for alkanals

(a) Brandy test

The reagent is an orange powder. Formation of a yellow precipitate shows a positive identification of a carbonyl group.

(b) Tollens test

A silver mirror on the inside of the test tube. A tollen test will follow a brady test to distinguish alkanal from alkanone.

(c) Fehling's test or Benedicts solution

Fehling's solutions when mixed with ethanal change colour from blue to green to orange precipitate and then red precipitate (copper mirror). The ethanal forms ethanoic acid

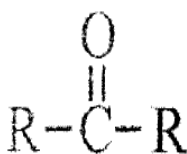
Uses of ethanal

1. It is a raw material in production of cellulose acetates, ethylacetate and athanoic acid.
2. It is used as a reducing agent in mirror production.
3. Used as a solvent in the rubber tanning and paper industries.
4. Used as a denaturant for alcohol.
5. Used in fuel composition.
6. Used in manufacture of perfumes.
7. Used in manufacture of plastics and synthetic rubber.
8. Formalin contains 40% methanal aqueous solution and thus used as a preservative.

ALKANONES

Alkanones have a carbonyl group between two alkyl groups (R and R')

The general formula of alkanones is

**Nomenclature of alkanones**

Names, formulae and structures of the first four alkanones

Name	Condensed formulae	Structural formulae	Molecular formulae
Propanone	CH_3COCH_3	$\begin{array}{c} \text{H} & \text{O} & \text{H} \\ & & \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\ & & \\ \text{H} & & \text{H} \end{array}$	$\text{C}_3\text{H}_6\text{O}$
Butanone	$\text{CH}_3\text{CH}_2\text{COCH}_3$	$\begin{array}{c} \text{H} & \text{H} & \text{O} & \text{H} \\ & & & \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ & & & \\ \text{H} & \text{H} & & \text{H} \end{array}$	$\text{C}_4\text{H}_8\text{O}$
Pentan-2-one	$\text{CH}_3\text{CH}_2\text{CH}_2\text{COCH}_3$	$\begin{array}{c} \text{H} & \text{H} & \text{H} & \text{O} & \text{H} \\ & & & & \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ & & & & \\ \text{H} & \text{H} & \text{H} & & \text{H} \end{array}$	$\text{C}_5\text{H}_{10}\text{O}$
Penta-3-one	$\text{CH}_3\text{CH}_2\text{COCH}_2\text{CH}_3$	$\begin{array}{c} \text{H} & \text{H} & \text{O} & \text{H} & \text{H} \\ & & & & \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ & & & & \\ \text{H} & \text{H} & & \text{H} & \text{H} \end{array}$	$\text{C}_5\text{H}_{10}\text{O}$

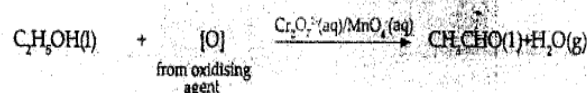
Sources of alkanones

- By-products obtained when body fat provides energy instead of glucose.
- The oxidation of secondary alcohols under controlled conditions.

Oxidation of secondary alcohols to form alkanones**Example**

Oxidation of propan-2-ol by acidified potassium dichromate (VI) produces propanone.

Ethanol + Oxygen \longrightarrow Ethanal + water

**Physical properties of Alkanones**

- All alkanones are liquids.
- Alkanones are polar because of the carbonyl group.
- Solubility of alkanes decreases down the homologous series.
- Boiling points and melting points increase with increasing molecular mass.

Physical properties of propanone

- It is a colourless liquid.

- It is flammable.
- Miscible with water, alcohol and ether.
- It has low boiling point of 56°C.
- Has a characteristic sweet odour.

Preparation of alkanones

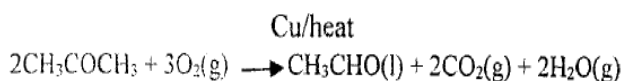
- Alkanones can be prepared by oxidation of alkanols.
- By dehydrogenation of propan-2-ol using heated copper.
- By dry distillation of calcium ethanoate (calcium acetate).
- By **Wacker** process where propene and air mixture is passed between a solution of copper (II) chloride and palladium (II) chloride between 20°C and 60°C.

Chemical properties of alkanones

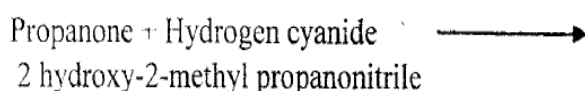
- It burns in oxygen with a blue flame to produce carbon dioxide and water.



- No effect on both red and blue litmus papers.
- Alkanones are difficult to oxidise. They cannot be oxidized under conditions which alkanals get oxidized. Hot copper metal catalyses oxidation of propanone to ethanal, carbon dioxide and water.

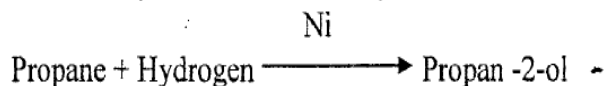


- Propanone will undergo **addition reaction** with hydrogen cyanide, HCN, to give 2-hydroxy-2 methyl propanonitrile.



- Propanone will also undergo addition reaction when shaken with saturated aqueous sodium hydrogen sulphite (NaHSO_3).

- Reduction reaction:** Alkanones are reduced by hydrogen in the presence of a platinum or nickel metal catalyst to form secondary alcohols.



- Polymerization:** Alkanones are not reactive enough to polymerise easily.

Test for alkanones

(a) Brady's test

A yellow or orange precipitate is formed which gives a positive test for an alkanone (ketone). Both alkanals and alkanones give positive test with Brady's test.

(b) Tollen's test

Positive test indicate a silver mirror. A negative test result indicates the compound is alkanone i.e. silver mirror does not form.

(c) Fehling's test or Benedicts solution

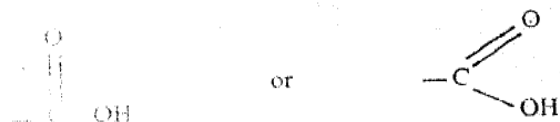
If the colour of the solution does not change alkanone is present. Alkanone will not change the colour of Fehling's solution.

Uses of propanone

- Propanone is used as a solvent for plastics and synthetic fibres.
- Propanone is used as nail varnish remover.
- It is capable of dissolving plastic, glasses, jewelry, rayon's garment, pens and pencils.
- Propanone is used as a solvent for both polar and non-polar compound.

ALKANOIC ACIDS

Alkanoic acids are organic acids. They are also called **carboxylic acids**. They have a **carboxyl group** ($-\text{COOH}$) as their functional group.



General formula and nomenclature of alkanoic acids

Alkanoic acids form homologous series with a general formula $\text{R}-\text{COOH}$ or $\text{C}_n\text{H}_{2n+1}\text{COOH}$

The formula of the first member, methanoic acid, is HCOOH . The second member of the alkanoic acids series is **ethanoic acid**. It has two carbon atoms. Since one of the carbon atoms is already attached to the functional group, the value of n in the general formula $\text{C}_n\text{H}_{2n+1}\text{COOH}$ is 1.

Table: Natural sources of alkanoic acid

Name	Where found or used
1. Citric acid	Citrus fruits, such as oranges, lemons
2. Tartaric acid	Grapes, health salts, baking powder
3. Lactic acid	Sour milk
4. Ethanoic acid	Vinegar
5. Methanoic acid	In ant, bee and nettle stings
6. Carbonic acid	Coke, lemonade and other fizzy drinks
7. Butanoic acid	Cheese
8. Tannic acid	Tea

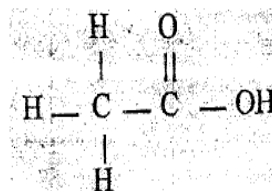
Preparation of alkanoic acids

The alkanol is mixed with excess acidified potassium dichromate (VII) and warmed.

In the first step an alkanal and water are formed. The second step involves further oxidation of alkanal to alkanoic acid. Overall reaction is:



The structural formula is represented as:



Ethanoic acid

Alkanoic acids as homologous series

- They have the functional group $-\text{COOH}$ attached to the hydrocarbon chain.
- Their general formula is $\text{C}_n\text{H}_{2n+1}\text{COOH}$ or RCOOH , where $n=0, 1, 2$.
- Members of the series differ from the next by a $-\text{CH}_2-$ group.
- They have similar chemical properties.
- They show a trend in their physical properties such as increase or decrease in melting points, boiling points and densities.



Physical properties of alkanoic acids

- The first ten straight chain alkanoic acids are liquids at room temperature.
- The first five members are soluble in water and the rest are insoluble.
- The boiling points and melting points of alkanoic acids gradually increase as the length of the hydrocarbon chain increases.
- They conduct electric current though they are not strong acids.

Chemical properties of alkanoic acids

IDENTIFICATION OF UNKNOWN ORGANIC COMPOUNDS

Deducing families and formulae of Organic compounds

1. Empirical formula of organic compound

2. RMF and structural formula

The relative molecular formula of a compound is calculated using the **relative molecular mass** of the compound provided.

We can also tell the identity of organic compound by looking at its structure.

3. Chemical reactions and the products in identification of organic compounds.

Each homologous series has certain chemical characteristics dependent on its functional group or reactions particular to it.

Tests for organic compounds

The table below is a summary of tests for alkane, alkene, alkanol, alkanoic acid, alkanal and alkanone

Octyl ethanoate	$\text{CH}_3 \text{COO} (\text{CH}_2)_7 (\text{CH}_3)$	Orange
Propyl ethanoate	$\text{CH}_3 \text{COOCH}_2 \text{CH}_3$	Pear
Ethyl butanoate	$\text{CH}_3 \text{CH}_2 \text{CH}_2 \text{COO} \text{CH}_2 \text{CH}_3$	Pineapple

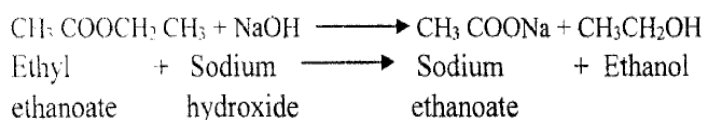
2. Alkanoates can be prepared synthetically by condensation reaction of alkanoic acids with alkanols.

Physical properties of esters

- Esters are often volatile with pleasant odour (smell).
- Small chain are fairly soluble in water.
- Esters are soluble in organic solvents.

Chemicals properties of esters

Hydrolysis of esters - when esters are heated under reflux with a dilute alkali such as sodium hydroxide solution it is hydrolysed. For example:



The products are a **carboxylate salt** and an alcohol which can be distilled off easily.

Uses of alkanoates

1. Many esters are used as flavours and perfumes.
2. Smaller esters are quite volatile and are used as solvents in adhesive, inks and paints.
3. Pentyl ethanoate is used in nail varnish.
4. Esters are used in making of soaps.

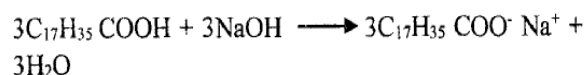
Saponification

The hydrolysis of esters using alkaline is referred to as **saponification**.

Soap making

When a fat is boiled with strong alkaline solution it is hydrolysed to give glycerol (propan-1, 2, 3 triol) and soap (sodium octadecanoate). The process is called **alkaline hydrolysis**.

Each part of the acid reacts with NaOH to form the salt i.e. sodium octadecanoate ($\text{C}_{17}\text{H}_{35}\text{COO}^- \text{Na}^+$) which is the soap. This is neutralization reaction.



Sodium octadecanoate (soap)

The general formula for ordinary soap is $\text{C}_n\text{H}_{2n+1} \text{COO}^- \text{Na}^+$

Hydrophobic means the dislike of the water. **Hydrophilic** means having a strong affinity for water.

Identification test table

[illegible]

A flow diagram to identify some selected organic compounds

