DANIEL MPOMJOLI

MALAWI SCHOOL CERTIFICATE OF EDUCATION

DANIEL

Daniel

FORM 3 CHEMISTRY

STUDENTS NOTES

MADALISO

$$\text{percentage yield} = \frac{\text{Actual yield}}{\text{theoretical yield}}$$

find theoretical yield

theoretical yield can be found by finding the RFM

percentage yield

Actual yield

KAYINGA

Table of content

Core Element: Analytical skills in chemistry

1. Experimental techniques.. 3

Core Element: Inorganic compounds

2. Nitrogen, sulphur and phosphorus.. 14

Core Element: Chemical composition of matter

3. Chemical bonding and properties of matter 23

Core Element: Chemical reactions

4. Stoichiometry.. 33
5. Heats of reactions.. 52

Core Element: Organic chemistry

6. Alkanols 58
7. Alkanals and alkanones 67
8. Alkanoic acids.. 74
9. Alkanoates 80
10. Identification of unknown organic compounds 86

Topic 1: Analytical skills in chemistry

Experimental techniques

Chemical wastes

- A chemical waste is a product from a chemical reaction or an expired product which is no longer needed.

Hazardous Chemical wastes hazardous

- A waste that display a hazardous characteristic or is specifically listed by name as hazardous.

Classes of Hazardous Chemical wastes

1. **Ignitable Wastes**
   - These are flammable chemical wastes.
   - They easily catch fire
   - Examples include; ethanol, sodium nitrate, hydrogen gas xylen, acetone etc.
2. Corrosive chemical wastes
   - Aqueous solutions with pH less than or equal to 2 or greater than or equal to 12.5
   - Examples include hydrochloric acid, nitric acid, sodium hydroxide.
3. Reactive Chemical wastes
   - Wastes that react violently when mixed with other substances
   - Examples include; sodium or lithium when reacted with water, cyanides or sulphides bearing wastes when mixed with acids or bases.
4. Toxic Chemical wastes
   - Wastes that are harmful to human health and the environment.
   - Examples include; heavy metals, pesticides, organic compounds.

Disposing Chemical Wastes

- Wastes must be stored in a well-sealed containers labelled "**hazardous wastes**"
- Reactive wastes must be stored in separate containers

Disposing Solid Wastes

✓ Very harmful wastes must be incinerated

✓ Less harmful wastes such as paper, plastics, rubber and wood should be placed in waste bin in the lab and be disposed in rubbish pit or local authority refuse collection area.

✓ Some less harmful wastes such as paper and plastics can be recycled and be made into useful products.

Disposing Liquid Wastes

- ✓ Innocuous (completely harmless) aqueous wastes may be poured into sinks eg NaCl
- ✓ Some wastes can be washed with excess water eg acids and bases, harmless soluble inorganic salt and salts that contain alcohol
- ✓ Recycling for wastes such as elementary mercury, spent acids and bases (weakened by use)
- ✓ Some liquid wastes can be incinerated such as chlorinated solvent.

Note:

- The following wastes must not be discharged into sinks
  - ✓ Raw Chemical wastes (unused chemicals)
  - ✓ Chlorinated hydrocarbons
  - ✓ Brominated wastes
  - ✓ Chlorofluorocarbons (CFCs)
  - ✓ Cyanide wastes
  - ✓ Heavy metals (Hg, Pb, Zn,)
  - ✓ Corrosive wastes eg strong acids and bases
  - ✓ Solvent wastes eg acetone, ethyl ether, benzene, methanol, xylene.
  - ✓ Oil and grease wastes
  - ✓ Ignitable wastes
  - ✓ Reactive wastes
  - ✓ Solid or sticky wastes
  - ✓ Hot liquid and vapour wastes
- These wastes must be collected and managed as hazardous chemical wastes.

Disposing Gaseous Wastes

- Gases such as CO_2, NO, NO_2, H_2S, SO_2, and Cyanides are harmful to both human beings and the environment

  Disposing

- ✓ Do all the experiments that emit these gases in a fume board or in a well-ventilated room or in an open space.

Minimising Chemical Wastes

- ✓ Practicing the concept of source reduction
- ✓ Order small quantities (just the required amount)
- ✓ Share surplus
- ✓ Keep good record of the chemicals available

- ✓ Where possible substitute hazardous chemical with non-hazardous ones
- ✓ Reduce the scale of lab experiments to reduce the volume of waste products.

Scientific Investigation

1. Techniques and suitable apparatus for a variety of simple experiments
   - ✓ Use appropriate apparatus to record a range of measurements eg volume, mass, time, temperature
   - ✓ Use all kinds of burners, water bath, sand bath and electric heaters.
   - ✓ Use pH charts and pH meters
   - ✓ Use titration apparatus
   - ✓ Use distillation apparatus
   - ✓ Use appropriate apparatus to make standard solutions.
   - ✓ Use melting points and boiling points apparatus.
   - ✓ Use of paper Chromatography
   - ✓ Use of the conductivity apparatus
   - ✓ Safely hand solids, liquids and gases including corrosive, irritating, flammable and toxic.
2. **Presenting experimental data from observations and measurements**
   - Data from experiments and observation must be presented in graphs and tables for easy interpretation.

   Tables
   - Show the data in two or more columns which are well labelled with the names of the quantities under study.
   - Each quantity is written together with its units

   Example

Temperature (K)	Volume (cm^3)

Graphs

- Different type of graphs such as bar graph, pie charts and histogram are used to interpret experimental data.
- The most common one is a line graph in which the shape of the graph corresponds to its meaning.

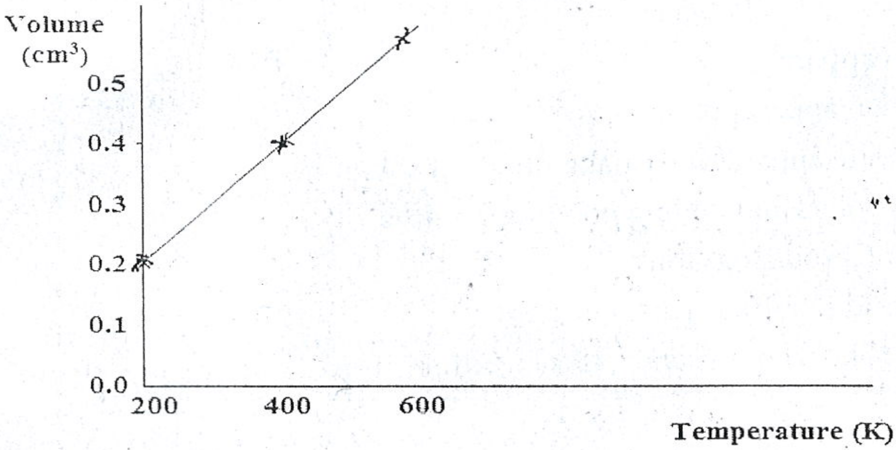
- In a line graph, the independent variable is always plotted on the horizontal axis (X-axis) while the depend variable is shown on the vertical axis (Y-axis)

Important

- The two axes must be well labelled with the name of a quantity and its units .
- Choose appropriate scale for the graph (**apply $^3/_4$rule**)
- The graph should have a clear title.

Example

Graph of Volume against Temperature



Purity of a substance

Pure substance

- Substance with uniform composition throughout ie it is made up of only one kind of particles (atoms, ions, molecules).
- Has no other substances mixed with it and cannot be separated by physical means.
- Has fixed (sharp) melting points and boiling points

Mixture

- Two or more substances that are not chemically combined with each other and can be separated by physical means.
- It melts or boils over a range of temperatures

Purity of a substance can be determined by:

1. Using melting points and boiling points
2. Using paper chromatography

Using melting points and boiling points

- Melting point is the temperature at which a solid melts and becomes a liquid
- Boiling point is the temperature at which a liquid boils and becomes a gas

- No two substances have the same pair of melting and boiling pts. Hence melting and boiling pts can be used to identify a substance.

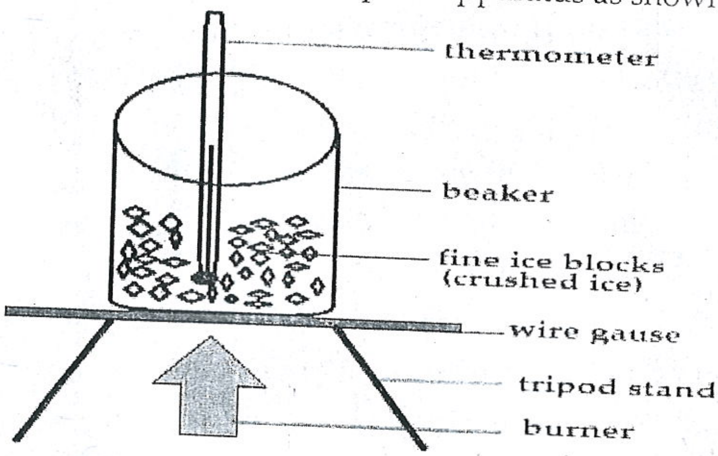
Experiment

Aim: To identify a pure substance using melting points and boiling points

Materials: ice blocks, beaker, burner and thermometer

Procedure:

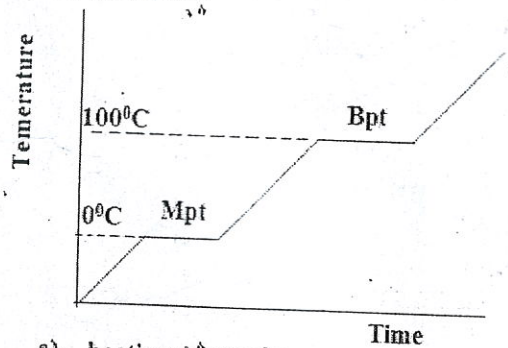
- Put some fine ice blocks in a beaker and set up the apparatus as shown below.



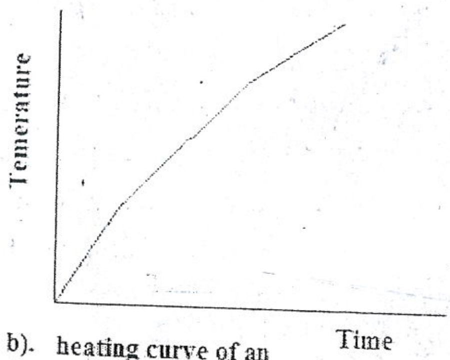
- Measure the initial temperature of the ice
- While stirring, note and record the temperature of the ice every 2 minutes.
- Record the temperature every two minutes until the water starts boiling.

Results and Discussion

- Plot the graph of temperature against time.
- The graph shows a rising gradient which become constant at 0^0C and at 100^0C as shown in **(a)below**



a). **heating curve of a pure substance**



b). **heating curve of an impure substance**

- 0^0C and at 100^0C are the Mpts and Bpts of water respectively.
- When an impurity is added to water the graph shows a rising gradient through out ie no constant temperature as shown in **(b) above**

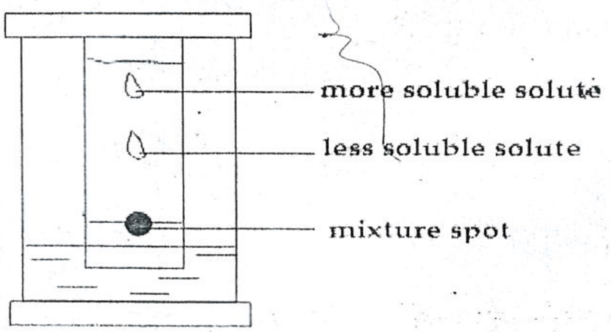
94

Conclusion

- Pure substances such as pure water have a fixed mpt and bpt while impure substances boil over a melt or boil over a range of temperatures.
- Impurities will **lower** the melting point and **raise** the boiling point

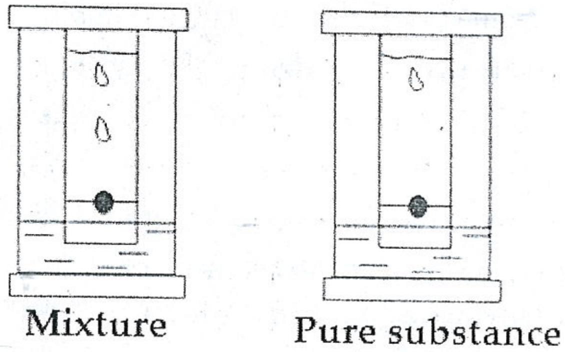
Determination of purity by chromatography

- Chromatography is a method of separating and identifying mixtures;
  - ✓ To separate the solutes in a mixture eg Pigments from plants, dyes from ink, amino acids from proteins for analysis.
  - ✓ Can be used for example to identify artificial dyes in food eg identify poisons and drugs; detect traces of unlawful dyes or other additives in foodstuffs.
  - ✓ To determine the purity of a given substance
- Principle involved depends upon the different in **solubility** of the substances in the mixture in the solvent used.
- A spot of the mixture is placed near the bottom of a piece of chromatography paper and the paper is then placed upright in a suitable solvent, eg water.
- As the solvent soaks up the paper, it carries the mixtures with it.
- Different components of the mixture will move at different rates. This separates the mixture out.
- A dye strongly adsorbed onto paper, and not very soluble in solvent will be left behind
- A dye weakly adsorbed onto paper, and very soluble in solvent will be carried furthest

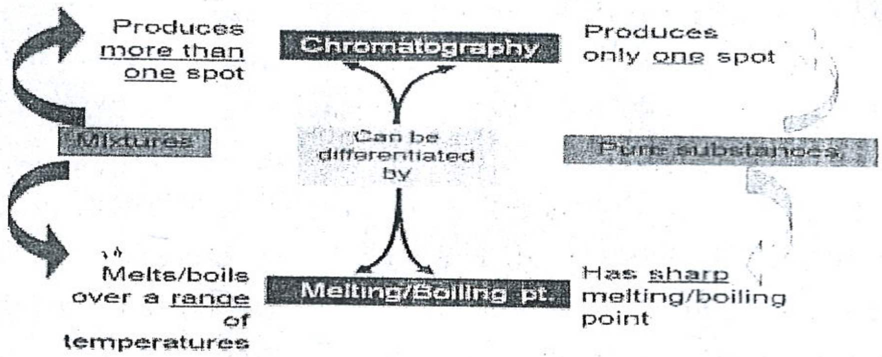


Criteria Of Purity

- A pure substance will give only one spot in the chromatogram



Summary

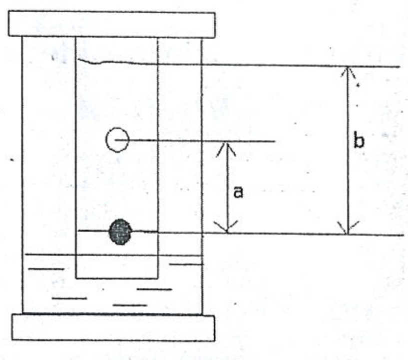


Relative flow values (R_F)

- This is the ratio of the distance travelled by the spot to the distance travelled by the solvent ie

$$(R_F) = \frac{\textit{distance travelld by the spot}}{\textit{distance travelled by the solvent}}$$

- Measure the distance from the starting point to the center of the spot on the chromatography paper (distance a).
- Measure the distance from the starting point to the solvent front (distance b).



From the diagram, $(R_F) = \frac{a}{b}$

- The values obtained from calculations using experimental data are then compared with the values in the data books to identify the substances.

Testing for ions gases and water

1. **Identifying cations**

Ions – Charged Species. Metals tend to form cations and Nonmetals form anions.

Cations – Positively charged ions. eg Na^+, K^+, Ca^{2+}, Ba^{2+}, Fe^{3+}, Cu^{2+},Mg^{2+}, Al^{3+}, NH_4^+

Anions – Negatively charged ions. Eg Cl^-, Br^-, I^-, NO_3^-, SO_4^{2-},

Testing for cations

Using precipitation reactions

- Metallic ions or cations are identified based on the colour and solubility of the hydroxides formed from their reaction with ***aqueous ammonia and sodium hydroxide.*** It is based on the formation of a precipitate

Precipitate

- A solid beneath a liquid. Reactions that form a solid at the bottom of a liquid are called precipitation reactions.
- Many Cations form precipitates when mixed with sodium hydroxide or ammonia solution.

a. **Using aqueous sodium hydroxide**

Procedure

a) Put a 2cm³ of the test solution into test tube.
b) Add a few drops of aqueous sodium hydroxide and shake gently.
c) Observe the colour of any precipitate.
d) Add some more sodium hydroxide solution and shake the test tube.
e) Record whether precipitate dissolves or not and any colour changes

- Some common metal cations that can be identified using these reagents are shown in the table below

Metal cation	Observation with sodium hydroxide aq	Equation for the reaction
Copper (II), Cu^{2+}	Blue precipitate, insoluble in excess	$Cu^{2+}{}_{(aq)} + 2OH^-{}_{(aq)} \rightarrow Cu(OH)_{2\,(aq)}$
Iron(II) Fe^{2+}	Green precipitate, insoluble in excess	$Fe^{2+}{}_{(aq)} + 2OH^-{}_{(aq)} \rightarrow Fe(OH)_{2\,(aq)}$
Calcium, Ca^{2+}	White precipitate, insoluble in excess	$Ca^{2+}{}_{(aq)} + 2OH^-{}_{(aq)} \rightarrow Ca(OH)_{2\,(aq)}$
Zinc Zn^{2+}	white precipitate, soluble in excess	$Zn^{2+}{}_{(aq)} + 2OH^-{}_{(aq)} \rightarrow Zn(OH)_{2\,(aq)}$
Magnesium, Mg^{2+}	white precipitate, insoluble in excess	$Mg^{2+}{}_{(aq)} + 2OH^-{}_{(aq)} \rightarrow Mg(OH)_{2(aq)}$
Aluminium, Al^{3+}	white precipitate, soluble in excess	$Al^{3+}{}_{(aq)} + 3OH^-{}_{(aq)} \rightarrow Al(OH)_{3\,(aq)}$

b. **Using aqueous ammonia**

Procedure

i. Put $2cm^3$ of the test solution into test tube.

ii. Add a few drops of aqueous ammonia.

iii. Observe the colour of any precipitate.

iv. Add excess aqueous ammonia and shake the test tube.

v. Record whether precipitate dissolves or not and any colour change

Table of Results

Metal cation	Result with aqueous ammonia
Calcium, Ca^{2+}	No precipitate, very slight white precipitate
Zinc Zn^{2+}	white precipitate, soluble in excess (precipitate dissolves)
Magnesium, Mg^{2+}	white precipitate, white precipitate remain in excess(insoluble)
Aluminium, Al^{3+}	No precipitate or very slight white precipitate, insoluble in excess

copper(II) Cu^{2+}: Blue precipitate insoluble in excess

Test for ammonium ions

- To the test solid or solution add a little amount of dilute sodium hydroxide solution and gently heat.
- If Ammonia (alkali gas) is given off, the unknown substance contained ammonium ions
- The equation for the reaction is

$$NH^+_{4(aq)} + OH^-_{(aq)} + \rightarrow NH_3 + H_2O$$

Anion
→ Halide
→ sulphate
→ Nitrate.
→ carbonate.

2. **Identifying aqueous anions**

a. **Halides**

Test for halide ions

i. Silver nitrate test for halide ions (chloride, bromide and iodide)

ii. Put around 1 – 2 cm^3 of a solution of the test compound into a test tube.

iii. Add a few drops of dilute nitric acid

iv. Then add a few drops of silver nitrate solution.

v. observe

halide ion	observation with silver nitrate	Reaction equation
Chloride	white precipitate forms	$Ag^+_{(aq)} + Cl^-_{(aq)} \longrightarrow AgCl_{(s)}$
bromide	cream precipitate forms or pale yellow	$Ag^+_{(aq)} + Br^-_{(aq)} \longrightarrow AgBr_{(s)}$
Iodide	yellow precipitate forms	$Ag^+_{(aq)} + I^-_{(aq)} \longrightarrow AgI_{(s)}$

Pirco J.P

b. Sulphates

Test for sulfate ions

- Add a solution of barium chloride (or nitrate) to the solution under test.
- If a precipitate forms add dilute hydrochloric acid

Observation

- A white precipitate of barium sulphate forms in barium chloride which does not re-dissolve in acid.

$Ba^{2+} + SO_4^{2-} \rightarrow BaSO_{4(S)}$ **(insoluble in acids)**

c. Nitrates

Test for the nitrate ion

- Add a few drops of dilute sodium hydroxide are added to the test compound
- Followed by some aluminium powder.
- Warm the solution in a Bunsen flame
- Test any gas given off using red litmus paper.

Results

A gas is given off which turns the litmus blue. This shows that the gas is ammonia

$8Al_{(s)} + 3NO^{-}_{3(aq)} + 5OH^{-} + 2H_2O_{(L)} \rightarrow 3NH_3 + 8AlO^{-}_{2(aq)}$

d. Carbonates

Test for carbonate ions

- Add dilute hydrochloric acid to the solid salt (or to a solution) under taste.
- Bubbles of gas form. The gas turns the limewater milky, which shows that it is carbon dioxide.
- Equations for the reaction
  - $2\,HCl_{(aq)} + Na_2CO_{3(aq)} \rightarrow 2\,NaCl_{(aq)} + H2O_{(l)} + CO2_{(g)}$
  - $CO3^{2-}{}_{(aq)} + 2H^{+}{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

3. Tests for gases

i. Ammonia

a. Reaction with hydrogen chloride gas

- Carefully mix ammonia gas with hydrogen chloride gas
- White fumes of ammonium chloride form as follows

$NH_{3(g)} + HCl_{(g)} \rightarrow NH^{+}{}_4Cl^{-}{}_{(s)}$

b. Using red litmus

- If tested with red litmus, the indicator changes to blue ie ammonia is an alkaline gas

ii. Hydrogen

- Collect the gas in a tube and hold a lighted sprint to it
- The gas burns with a squeaky pop

pop sound

iii. **Chlorine**

- Hold damp litmus paper in the gas (chlorine is poisonous do this in a fume board)
- If the indicator paper turns red and then white (bleached)then the gas is chlorine

iv. **Carbon dioxide**

- Bubble the gas through lime water
- Lime water turns cloudy or milky if the gas is carbon dioxide

v. **Oxygen**

- Collect the gas in a test tube and hold a glowing splint to it
- If the splint immediately bursts into a flame the gas is oxygen

vi. **Sulphur dioxide**

- Hold a piece of filter paper soaked in acidified potassium dichromate solution on top of the gas. (in a fume board)
- If the paper turns from orange to green, then the gas is sulphur dioxide

4. **Testing for the presence of water**

- Mix some anhydrous copper sulphate (white powder) with a test liquid.
- If a blue solution is formed then the liquid is water

Or

- Mix some anhydrous cobalt (II) chloride (blue powder) with a test liquid.
- If a pink solution is formed then the liquid is water

hold the litmus paper in the gas if the indicator paper turns ~~red~~ to stains red then white then the gas is

collect the gas in a tube and hold the light splint into it.

carbon dioxide
- bubble the gas in the lime water

Topic 2: Nitrogen, sulphur and phosphorus

1. Nitrogen

Sources of nitrogen

- About 78% of nitrogen occurs naturally in air where it can be obtained by fractional distillation of liquid air
- Nitrogen keeps on circulating between the air , the soil and living things in a set of processes called nitrogen cycle.

Assignment 1:

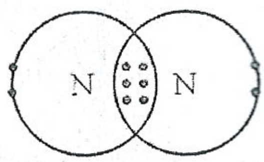
"with the aid of a well labelled diagram describe and explain the processes in the nitrogen cycle"

Properties of nitrogen

- Colourless, odourless gas
- 78% by volume in air
- It is a member of Group V with an electron configuration of 2,8,5
- exists as a diatomic molecule N_2
- It is slightly soluble in water

Inertness and Reactivity of Nitrogen

- Nitrogen gas molecule comprises of two nitrogen atoms covalently bonded together by a very strong triple bond. This strong bond makes nitrogen unreactive



Strong N≡N bond

- However at high temperatures nitrogen can reacts with other elements
  i. Reaction with reactive metals, Li and Mg, to form *nitrides.*
     - ✓ $3Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$
     - ✓ $6Li_{(s)} + N_{2(g)} \rightarrow 2Li_3N_{(s)}$
  ii. Reaction with oxygen to form nitrogen oxides
     - ✓ $N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$
  iii. With hydrogen at special conditions
     - ✓ $N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)$, Haber Process

Uses of nitrogen

- ✓ Liquid nitrogen is a coolant

- ✓ Most important use is in the manufacture of ammonia and nitrogenous fertilizers
- ✓ Quick freeze food, freeze liquid in damaged pipes and in shrink fitting
- ✓ To flush out food packaging and keep food fresh
- ✓ Use in the form of nitrous oxide as an anaesthec

Ammonia

- It is the most important hydride of nitrogen. Its solution in water, aqueous ammonia, is an important common alkali

Preparation of ammonia

i. **Laboratory preparation**

- In a laboratory Ammonia can be prepared by heating an ammonium salt with an alkali .
- Calcium hydroxide and ammonium chloride is a convenient mixture

$2NH_4Cl_{(s)} + Ca(OH)_{2\,(s)} \rightarrow 2NH_{3\,(g)} + CaCl_{2\,(aq)} + 2H_2O_{(l)}$

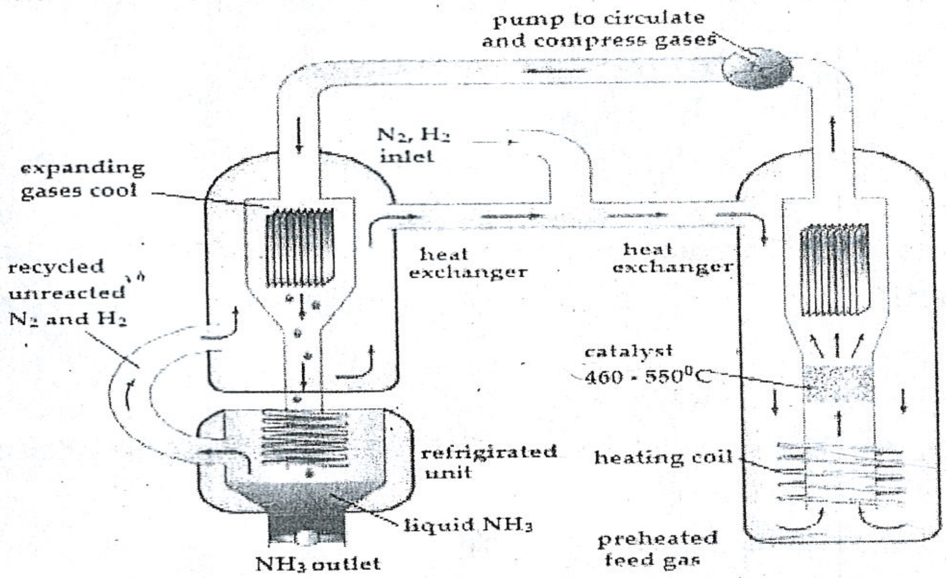
- The gas is then dried by passing it through a tower containing calcium II oxide (quicklime)

ii. **Industial preparation**

- Ammonia, NH_3, is produced commercially by the Haber Process.

$$N_{2(g)} + 3H_{2(g)} \underset{460^0C\ 200atm}{\overset{Fe}{\rightleftharpoons}} 2NH_{3(g)} \qquad \Delta H = -92KJmol^{-1}$$

Haber Process



- N_2 and H_2 are pumped into a chamber.
- The pre-heated gases are passed through a heating coil to the catalyst bed.

- The catalyst bed is kept at 460 - 550 °C under high pressure.
- The product gas stream (containing N_2, H_2 and NH_3) is passed over a cooler to a refrigeration unit.
- In the refrigeration unit, ammonia liquefies, but not N_2 or H_2.

Physical properties of ammonia

- A colourless, pungent gas that easily liquefied (b.p. –33°C)
- Has a strong chocking smell
- It is less dense than air
- Extremely soluble in water to form a weakly alkaline solution

Chemical properties of ammonia

a. Reacts with hydrogen chloride to form white smoke of ammonium chloride

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

b. Reaction with water where it dissociates into ammonium and hydroxide ions

$$NH_3(aq) + H_20(l) \rightarrow NH_4^+(aq) + OH^-(aq)$$

c. Reaction with nitric acid to form ammonium nitrate(fertilizer)

$$NH_3(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq)$$

d. Reaction with oxygen

$$4NH_3(g) + 3O_2(g) \rightarrow 2N_2(g) + 6H_2O(g)$$

Uses of ammonia

- ✓ It is used in the manufacture of fertilizer eg ammonium nitrate
- ✓ It is used in softening water
- ✓ It is used in making nitric acid
- ✓ It is used in making plastics
- ✓ Making explosives
- ✓ It is used in making ammonium chloride which is used in dry cell

Nitric acid (HNO_3)

- A very strong acid
- It is a colourless liquid when pure.
- Turns yellow because of dissolved NO_2 formed from the decomposition of HNO_3 when exposed to sunlight.

Laboratory preparation of nitric acid

- In a laboratory nitric acid is prepared by heating potassium nitrate with concentrated sulphuric acid.

$$NO^-_3 + H_2SO_{4(aq)} \rightarrow HSO^-_{4(aq)} + HNO_{3(aq)}$$

- Nitric acid distils over and and is corrected in a water cooled receiver
- Some nitric acid decomposes and form brown fumes of nitrogen dioxide which gives the collected nitric acid a yellow colour

Industrial preparation of nitric acid

- Nitric acid is produced from ammonia in a process called Ostwald process which is in three stages

Stage 1

- Mixture of air & ammonia heated to 230°c and is passed through a metal gauze made of platinum (90%) & Rhodium (10%).
- Reaction produces a lot of heat energy..
- Energy is used to keep reaction vessel temp at 800oc.
- Reaction produces nitrogen monoxide (NO) and water.

Ammonia + oxygen → Nitrogen monoxide + water

$$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(g)}$$

Stage 2

- Colourless nitrogen monoxide gas produced from 1st stage is then reacted with oxygen from the air to form brown nitrogen dioxide gas (NO_2).

Nitrogen monoxide + oxygen → Nitrogen dioxide

$$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$$

Stage 3

- The nitrogen dioxide is then dissolved in water to produce nitric acid.

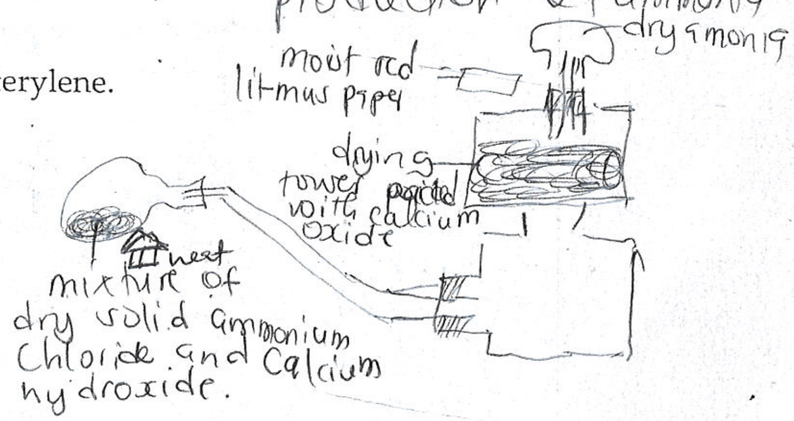
Nitrogen dioxide + water → Nitric acid + nitrogen monoxide

$$3NO_{2(g)} + H_2O_{(l)} \rightarrow 2HNO_{3(aq)} + NO_{(g)}$$

Uses of Nitric acid

- Nitric acid produced is used in the manufacture of the following:
  - ✓ Artificial fertilisers – Ammonium nitrate.
  - ✓ Explosives, Dyes and drugs.
  - ✓ Artificial fibres, such as nylon and terylene.
  - ✓ Used in treatment of metals.
  - ✓ Used in making drugs.
  - ✓ It is used as a laboratory reagent.

production of ammonia



Sulphur

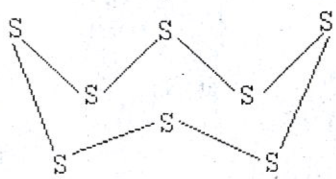
- Sulphur is a non-metal element of group 6
- It is a brittle yellow solid.
- Has an electron configuration of 2,8,6 and atomic mass of 32.065 a.m.u.

Sources of sulphur

1. Volcanic regions
2. Crude oil such as organic sulphur compounds and natural gases such as hydrogen sulphide.
3. Metal ores, such as Iron sulphide (FeS2), Lead sulphide (PbS), Zinc sulphide (ZnS).
4. Tissues of living plants. For example the substance that makes people cry when slicing onion is a sulphur compound. And the smells of garlic, mustard, cabbage are due to sulphur compounds.

Physical properties of sulphur

1. Has a low melting point and boiling point
2. It is insoluble in water but soluble in other solvents such as carbon tetrachloride and benzene
3. It does not conduct electricity and heat
4. 2. It is made up of crown shaped molecules, each with eight atoms, S8.

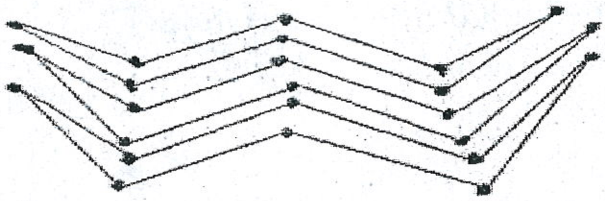


5. It has two allotropes; rhombic sulphur and monoclinic sulphur.

- Allotropes are different forms in which an element can exist in the same physical state.

1. **Rhombic sulphur (alpha sulphur or $\alpha - sulphur$)**

- Is a yellow crystalline solid with an octahedral shape
- The molecules in rhombic sulphur are packed more closely; interlock

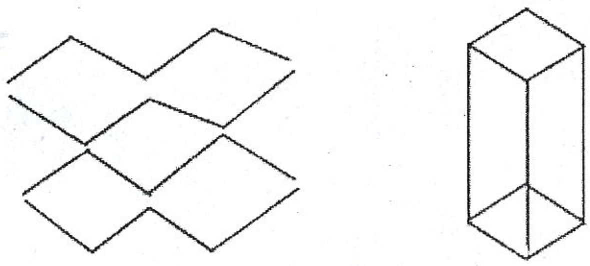


molecules are interlocked

- Formed when a solution of sulphur in methylbhenzine is heated and cools to a temperature below 96^{0}C
- Rhombic sulphur is stable below 96^{0}C

2. **Monoclinic sulphur**

- Long prism shaped needles with its molecules not closely packed.



- Formed when a solution of sulphur in methylbhenzine is heated and cools to a temperature **above** 96^{0}C
- It is stable above 96^{0}C
- Slowly reverts to rhombic sulphur when the temperature of the solution falls below 96^{0}C.

Chemical Properties of Sulphur

1. Sulphur reacts with metals to form sulphides.
   a. Sulphur reacts with magnesium to form magnesium sulphide;
   $Mg + S \rightarrow MgS$.
   b. It also reacts with Iron to form Iron Sulphide;
   c. $Fe + S \rightarrow FeS$.
2. It burns in oxygen to produce Sulphur dioxide;
   $S + O2 \rightarrow SO2$

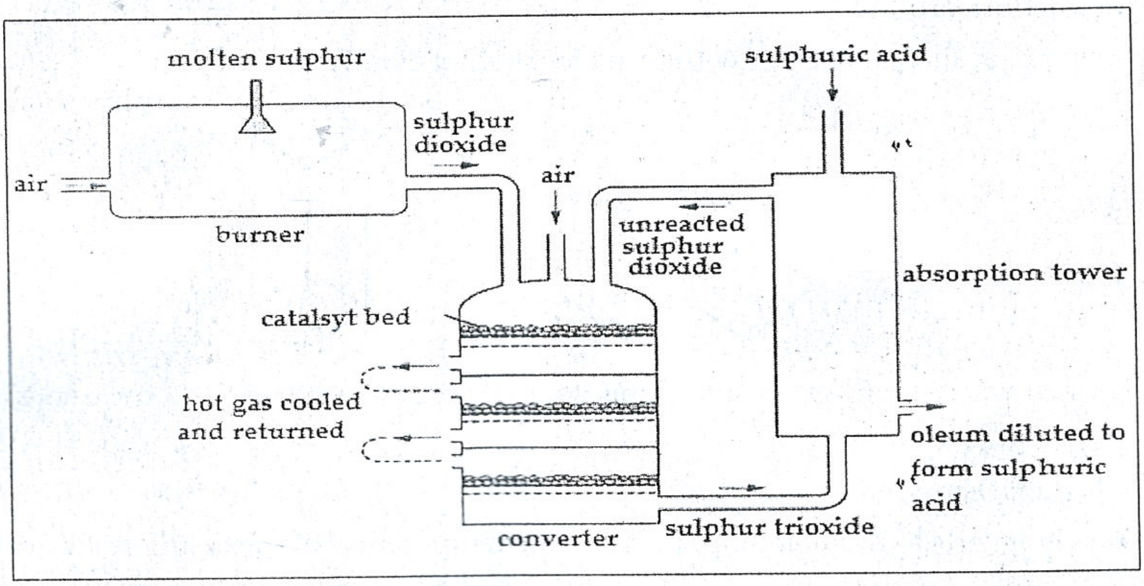
Uses of Sulphur

1. To produce sulphuric acid.
2. To vulcanise rubber. Vulcanisation means adding sulphur to rubber in order to toughen it.
3. To make matches, pesticides, drugs and paper.
4. To make sulphur concrete.
5. Manufacture of gun powder, plastic flowers and medicine
6. To manufacture of insecticides, fungicides, germicides
7. To manufacture fire extinguishers
8. To make artificial hair
9. To make bleaching agent

Sulphuric acid (H_2SO_4)

The major use of sulphur is the production of Sulphuric acid. The sulphuric acid is produced by a process known the contact process.

The Contact Process



The process has the following stages.

Stage 1

- Sulphur is burnt in air

$$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$$

Stage 2 & 3

- Sulphur dioxide is reacted with some more oxygen to produce and is made to pass through a catalyst called vanadium oxide which helps the sulphur dioxide react with oxygen to form sulphur trioxide

$$2SO_{2(g)} + O_{2(g)} \xrightarrow{V_2O_5} 2SO_{3(g)}$$

$2SO_3$ H_2SO_4

Stage 4

- Sulphur trioxide is mixed with concentrated sulphuric acid to produce a substance known as oleum ($H_2S_2O_7$)

$2H_2SO_4$

$$H_2SO_{4(aq)} + SO_{3(g)} \rightarrow H_2S_2O_{7(l)}$$

Stage 5

- The oleum is then added to water to produce sulphuric acid of the required concentration;

$$H_2S_2O_{7(l)} + H_2O_{(l)} \rightarrow 2H_2SO_{4(l)}$$

H_2SO_4 + $H_2S_2O_7$

SO_3

Uses of Sulphuric Acid

1. to manufacture fertilisers e.g. sulphate fertilisers
2. to make paint and dye stuff
3. to make fibres; eg in the manufacture of nylon.
4. as acid in car batteries where it acts as an electrolyte
5. to make detergents and soaps.
6. It is a dehydrating agent. Dehydrating agent – sulphuric acid dehydrates sugar into water and carbon, and also will dehydrate copper sulphate

- ✓ In the chemical industry , sulphuric acid is used to dry certain gas mixtures (such as N_2 and CO_2) for analysis.
- ✓ Ammonia gas is not able to be dehydrated by sulphuric acid, **explain why? (exercise)**

7. Used in cleaning metals.

3. Phosphorus

Sources of phosphorus

- Found in earth's crust as calcium phosphate, bone meal, composited farm manure, banana peels, eggs and in phosphate rocks.

Physical properties

1. Non-metallic group 5 element
2. It is a Biogenic element (produced by living organisms or biological processes)
3. Exist in three forms (allotropes) which are white phosphorus, red phosphorus and black phosphorus
4. Changes directly from solid to liquid when exposed to light
5. Insoluble in water
6. White phosphorus gives a beautiful greenish glow

Chemical properties

1. React with oxygen to produce phosphorus pentoxide

$$P_{4(S)} + 5O_{2(g)} \rightarrow P_4O_{10(S)}$$

2. React with group seven elements

$$P_{4(S)} + 6Cl_{2(g)} \rightarrow 4PCl_{3(l)}$$

Uses of Phosphorus

1. Manufacture of tooth paste, baking powder and matches
2. Used to produce detergents

$2H_2SO_4$

3. Used to produce phosphate salts which are used to manufacture food products such as biscuits.
4. It is a minerals required by livestock and poultry.
5. Complexes with calcium to give rigidity to bones.
6. Helps in energy and protein metabolism.
7. Almost every biochemical reaction that occurs in muscle, blood and other soft tissues involves phosphorus.
8. Affects protein synthesis, lean deposition in growing animals.

Phosphoric acid (H_3PO_4)

- Produced by either a Wet process or Thermal process

Wet process

- Phosphate rock is reacted with concentrated sulphuric acid to produce phosphoric acid and calcium sulphate as follows

$Ca_3(PO_4)_{2(S)} + H_2SO_{4(l)} \rightarrow H_3PO_{4(l)} + CaSO_{4(S)}$

- Phosphorus produced in this process is less pure.

Thermal process

- White phosphorus is burnt in air at about 1527°C to 2727°C to produce phosphorus pentoxide as follows

$P_{4(l)} + 5O_2 \rightarrow 2P_2O_{5(g)}$

- Phosphorus pentoxide is then reacted with water to produce phosphoric acid

$P_2O_{5(g)} + 3\,H_2O_{(l)} \rightarrow 2H_3PO_{4(l)}$

Uses of phosphoric acid

1. Manufacture of fertilizers
2. Removing rust from surfaces of metals
3. Use as a pH adjuster in cosmetics and other skin care products
4. Catalyst on the production of ethanol from ethane
   - ✓ Used as a food additive;
     - ➢ Enhance flavour
     - ➢ Enhance Freshness

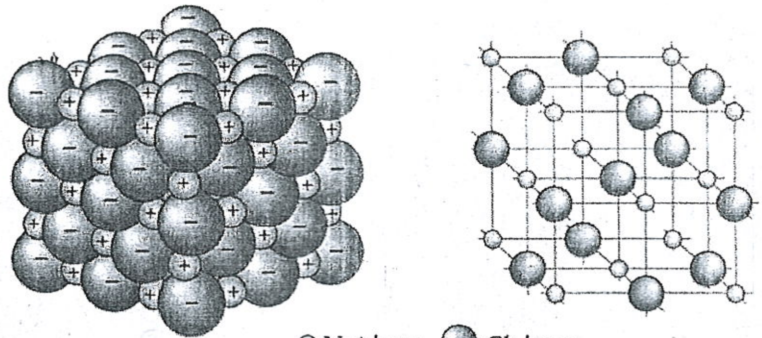
Topic 3: Chemical bonding and properties of matter

Ionic and Covalent Compounds

Ionic Compounds

- Ionic compounds are usually formed when metals bond to non-metals forming an ionic bond.
- Ionic bonds are the electrostatic forces of attraction between oppositely charged ions.
- The oppositely charged ions are arranged in a regular way to form giant ionic lattices or crystal.
- The illustration below shows part of a sodium chloride (NaCl) ionic lattice.

Van der waal's forces

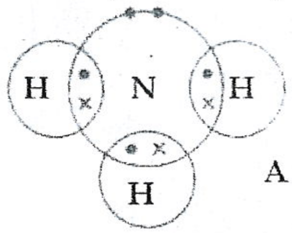


Na^+ ions Cl^- ions

q

1. **Covalent Compounds**

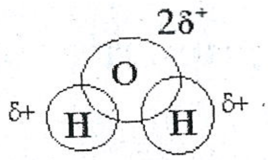
- A covalent compound is a compound in which the atoms that are bonded share electrons.
- Covalent compounds are formed when two non-metals bond to each other.
- In covalent compounds, there are strong covalent bonds between atoms in the molecules. Below is an example of a simple covalent structure.

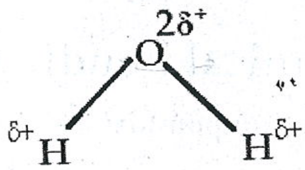


A molecule of ammonia

93
92
84
76
67
82

- In addition to the covalent bonds, there are weak forces of attraction between the molecules called Van der Waal's forces.
- Covalent compounds are mostly gaseous.
- Some have slight negative or positive charge at opposite ends of a covalent bond which gives them molecular polarity. Eg a water molecule shown below.





Properties of Ionic and Covalent Compounds

1. **Solubility**

Experiment 1

Aim: To Investigate the solubility of ionic and covalent compounds in solid state.

Materials: Sugar, NaCl, candle wax, water, 3 beakers, stirring rod.

Procedure:

1. Pour some water into one of the three beakers.
2. Place some sugar into the beaker with water.
3. Stir the mixture.
4. Observe and record what happens to solid
5. Repeat steps 1 to 4 above with NaCl and candle wax respectively.

Observation

- Sugar and NaCl dissolve in water while candle wax does not.

Conclusion

Ionic substances are soluble in water; some covalent substances are also soluble in water while other covalent substances are not soluble.

2. **Melting and Boiling Points**

Experiment 2:

Aim: To Compare the strength of electrostatic forces in covalent and ionic compounds

Materials: Sugar, common salt, source of heat, 2 crucibles, containers, beam balance, stop clock

Procedure:

1. Weigh 40g of common salt and sugar
2. Place them in separate crucibles
3. Heat the weighed amounts using a heat source
4. Record the time taken for the common salt and sugar to melt.

Observation

Sugar will melt faster than the common salt.

Conclusion

- The electrostatic forces in ionic substances are very strong compared to the electrostatic forces in covalent substances.
- This is because in ionic compounds there is a complete transfer of electrons. This makes ionic compounds to have very high melting points and boiling points compared to covalent compounds.

3. **Conductivity**

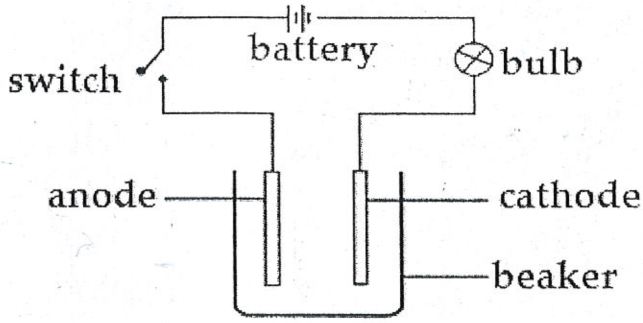
Experiment 3

Aim: to investigate conductivity of ionic and covalent compounds

Materials: table salt, sugar, distilled water, stirring rod, 2 cells, bulb, connecting wires

Procedure:

a. **Connect the circuit as shown below**



b. **Place** 5g of table un-dissolved salt in the beaker such that the salt touches the electrodes
c. Close the switch.
d. Observe and record what happens to the bulb.
e. Add $50cm^3$ of distilled water into the beaker and stir to dissolve the salt.
f. Observe and record what happens to the bulb.
g. Repeat the experiment using sugar

Results

Substance	Observation on bulb
Solid salt	No light
Dissolved salt	Gives light
Solid sugar	No light
Dissolved sugar	No light

Conclusion

- Ionic compound conduct in molten state or in aqueous solution whereas covalent compounds do not conduct at all.

Define dative bond these are

Summary of Properties of Ionic Compounds

Due to strong forces of attraction, all the ionic compounds exist in a solid state.

Some of the physical properties of ionic compounds are as follows:

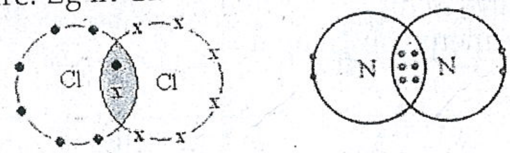
1. Generally, ionic compounds are soluble in water and in many polar solvents. Ionic compounds are insoluble in the organic compounds.
2. Ionic compounds have high melting and boiling points. This is because ionic bonds are very strong and a lot of energy is needed to break them.
3. Ionic compounds are strong electrolytes. In solid state they do not conduct electricity, but in molten state and aqueous solution they conduct electricity. Although ions are charged particles, ionic compounds can only conduct electricity if their ions are free to move.
4. Ionic compounds are not volatile. This means that they do not vapourise.

Summary of Properties of Covalent Compounds

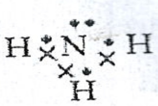
1. Covalent compounds are not usually very soluble in water but many are soluble in non-polar liquids or organic solvents like ether, alcohols, benzene, propanone, etc.
2. Covalent compounds generally have much lower melting and boiling points than ionic compounds.
3. Covalent compounds do not conduct electricity whether in molten state or aqueous solution because they do not have ions and free mobile electrons.
4. Covalent compounds are volatile (they easily change to vapour). Simple molecules are very volatile while giant molecules are non-volatile. Types of covalent compound

Pure Covalent Bond and a Dative Covalent Bond

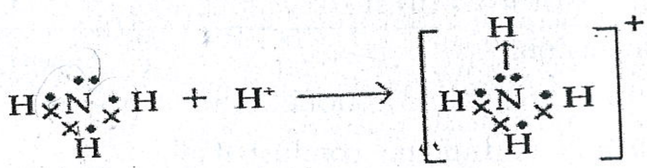
- A pure covalent bond is formed when each atom supplies one electron to the bond for the two atoms to share. Eg in Cl_2 and in N_2



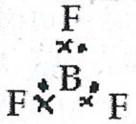
- A dative covalent bond (co-ordinate bond) is a covalent bond in which both electrons which are shared come from the same atom. Dative means to donate.
- For example, ammonia can donate the non-bonding pair and another atom, ion or molecule is able to accept it.



can donate the electron pair to H^-



- This donated ammonia bond is identical to any other non-donated bond. This can be written as $H_3N \rightarrow H$ to acknowledge the dative bond.
- In the same way BF_3 (boron trifluoride) and ammonia can undergo dative bonding. BF_3 is electron deficient like most boron compounds



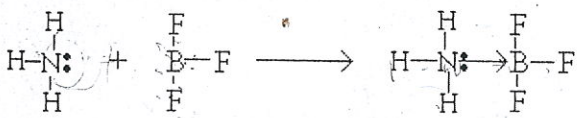
flourine has a full valence shell
bolon has only 6 electrons

$H_3N \rightarrow H$

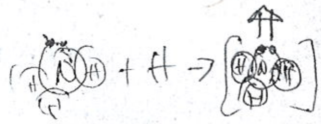
$H_3N \rightarrow H$

- As boron does not have a full outer shell, it is able to accept an electron pair from ammonia as follows.

dative bond



Ammonia-boron trifluoride
(adddition compound)



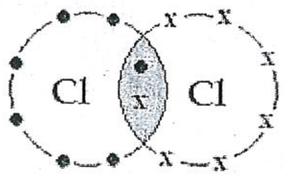
Polar and Non Polar Covalent Bond

Electronegativity

- The ability of an atom to attraction electrons from another atom in a covalent bond

Non Polar Covalent Bond

- Non polar covalent bond is a bond between two non-metals that have the same electronegativity and therefore have equal sharing of the bonding electron pair. Examples are, (H_2), oxygen molecule (O_2) and chlorine molecule shown below.

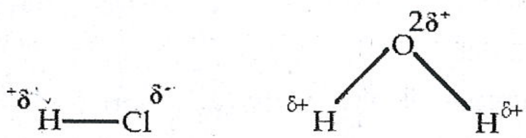


bonding

O_2

Polar Covalent Bonding

- A polar covalent bond is a bond between two non-metal atoms that have different electronegativity and therefore have unequal sharing of the bonding electron pair.
- Examples are water (H_2O), ammonia (NH_3), hydrogen chloride (HCl), alcohols, carboxylic acids, etc. (Nitrogen)
- In non-polar covalent bonding, the result is a bond where the electron pair spend more time near the more electronegative atom. This atom then obtains partial negative charge while the less electronegative atom has a partial positive charge.



H_2O

- The greek letter delta (δ) indicates 'partially'. A polar molecule always contains polar bonds, but some molecules with polar bonds are non polar eg. carbondioxide. Partial charges are denoted as (δ^+) (delta plus) and (δ^-) (delta minus).

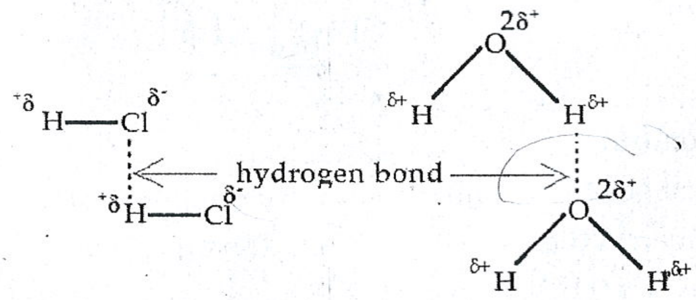
Intermolecular Forces (IMF)

- Intermolecular forces are forces of attraction between molecules. The strength of the intermolecular forces increases as the size of the molecules increase.

Types of Intermolecular Forces

1. **Van der Waal's forces (or London forces)**
- These are weak attractive forces between non-polar molecules.
- Non-polar molecules can exist in liquid and solid phases because van der Waals forces keep the molecules attracted to each other
- Exist between CO_2, CH_4, CCl_4, CF_4, diatomics and monoatomics.
- This attractive force is due to the electrostatic attraction of the electrons of one molecule or atom for the nuclei of another. .

2. **Hydrogen bonding** √
- Strong attraction between special polar molecules
- Hydrogen bonds occur between molecules that have a permanent dipole resulting from hydrogen being covalently bonded to more electronegative atoms, either fluorine, oxygen or nitrogen.
- Occurs **only** between H of one molecule and N, O, F of another as shown below



Allotropy

- This is when an element exist in two or more different forms in the same physical state.
- The different forms are known as allotropes.
- Allotropes are different forms of the same chemical element that exhibit different physical and chemical properties.
- Examples of elements that have allotropes include carbon, oxygen, phosphorus and sulphur.

this is When an element existed in between two or more different forms in the same physical state diatomic | monoatomic

Define Van der waal's force
→ This is a weak attraction force between non-polar molecule

1. **Allotropes of Oxygen**

- Oxygen has different allotropes which are dioxygen and ozone

a. **Dioxygen**

Dioxygen (O_2) also known as diatomic oxygen is the most common form of elemental oxygen. It is generally known as oxygen, but may be called dioxygen or molecular oxygen to distinguish it from the element itself.

b. **Ozone**

Another common allotrope of oxygen is triatomic oxygen or ozone (O_3) It is thermodynamically unstable and is highly reactive form of oxygen.

2. **Allotropes of Sulphur**

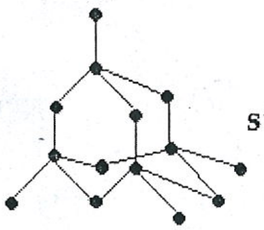
- There are a large number of allotropes of sulphur. The most commonly found in nature are rhombic sulphur and monoclinic sulphur.

3. **Allotropes of Carbon**

- Carbon has two natural crystalline allotropic forms, graphite and diamond.

Diamond

- Diamond is a giant structure of carbon atoms joined together by covalent bonds.
- Each carbon in a diamond crystal is bonded to four other carbon atoms making all the four outer electrons of each carbon atom 'localised' between the atoms in covalent bonding, therefore, the movement of electrons is restricted.
- The carbon atoms are arranged in a tetrahedron shape with bond angle of 109°.



16

structure of diamond

- **Physical properties of Diamond**

1. The hardest substance with boiling point of $3550^{0}C$. The hardness is due many strong covalent bonds holding the structure together..
2. Nonconductors (insulator). None of the electrons are free to move.
3. Insoluble in water.
4. Forms a colourless transparent crystal which sparkles in light.
5. Diamond has very *high* melting point.

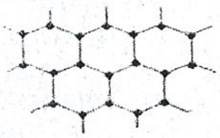
Uses of Diamond

1. Used in cutting instruments like glass cutters, marble cutters and in rock drilling equipment.

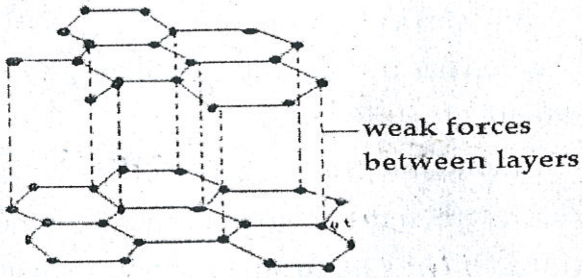
2. They are used for making jewelry.
3. Used as micro bearings in many small mechanical devices and laboratory instruments.
4. Diamond form the coating of wear resistant parts which are used to protect the surface of mechanical parts prone to wear and tear.

Graphite

- Graphite forms a giant structure made of flat sheets of carbon atoms.
- Each carbon atom uses only three of its four outer energy level electrons in covalently bonding to three other carbon atoms in a plane.
- Each carbon atom contributes one electron to a delocalized system of electrons that is also part of the chemical bonding.
- The delocalized electrons are free to move throughout the plane.



one layer of graphite



how the layers fit together

Physical properties of Graphite

- Graphite is a soft, slippery, grayish-black substance. Graphite powder is used as a lubricant.

1. They are good conductors of heat and electricity. Have some free electrons in the crystal.
2. They are brittle. Graphite rods used for electrolysis easily break when dropped.
3. They are insoluble in water.
4. They have high melting point due to many strong covalent bonds.

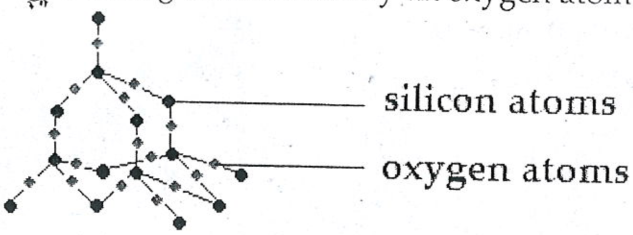
Uses of Graphite

1. Making lead pencils of different hardness, by mixing it with different proportions of clay.
2. Due to its slippery nature, graphite is used as a dry lubricant in machine parts.
3. Being resistant to chemicals and having high melting point and also because it is a good conductor of heat, graphite is used to make crucibles.
4. It is used to make carbon electrodes in dry cell and electric cars.

5. Graphite has the ability to absorb fast moving neutrons thus, it is used in nuclear reactors to control the speed of the nuclear fission reactions.
6. It is used as anodes in some aqueous electrolytic processes such as in the production of halogens (chlorine and fluorine).
7. It is a good candidate for a refractory; Making the inside wall of a furnace.

Silicon Dioxide

- Silicon dioxide also known as Silicon (IV) Oxide or Silica is an oxide of silicon with chemical formula (SiO_2).
- Crystalline silicon has the same structure as diamond but each silicon atom is bridged to a neighbouring silicon atom by an oxygen atom as shown below.



structure of silicon dioxide

➢ This is just a tiny part of a giant structure extending on all three dimensions.

Physical properties of Silicon dioxide

➢ Silicon dioxide has physical properties which are similar to those of diamond since their structures are similar. Some of the properties of silicon dioxide are as follows:

1. It has high melting point because of the very strong silicon-oxygen covalent bonds.
2. It is hard. This is due to the need to break the very strong covalent bonds.
3. Does not conduct electricity. All the valence electrons are localized, not free to move.
4. It is insoluble in water and organic solvents.

Metals

➢ The way that metal atoms are arranged to make a crystal lattice gives metals particular properties. Most of the uses of metals depend on their properties.

Physical Properties of Metals and Uses

1. They have high melting and boiling points.

➢ Used in making cooking utensils, filament of a light bulb

2. They are good conductors ofheat and electricity

➢ Used in electrical wiring, cooking materials

3. They are also malleable, which means they can be beaten or pressed into thin sheets.

➢ Helps in making objects of different shapes

4. They are ductile hence used in electric cables

5. They are hard and strong
➢ Used in construction of bridges and buildings and car bodies
6. They are shiny hence used in jewelries, ornaments and door handles
7. Metals are sonorous(make a ringing sound)
➢ Used in bell making

Alloys

➢ A mixture of two or more metals which is formed by thoroughly mixing molten metals.
➢ Metals are not made to chemically bond with one another but are simply mixed together to achieve certain properties.
➢ In some alloys a metal is mixed with a non-metal

Properties and uses of some alloys

Alloy	Properties	Uses
Brass	Alloy of copper and zinc, it is harder, sonorous and does not corrode	Making propellers of ship, water taps, musical instruments, jewelries door, screws, ornaments, furniture
Bronze	Alloy of copper and tin, it does not corrode and is sonorous	Making coins, medals, statues, bells springs,
Stainless steel	An alloy of chromium, iron and steel. It is tough and does not corrode	Cutlery, kitchen sinks, surgical instruments, car parts.
Solder	An alloy of lead and tin. Low melting point	Joining wires and pipes
Cupronickel	An alloy of copper and nickel. Hard with an attractive silver colour	Turbine blades, coinage metal
Duralium	An alloy of aluminium and copper. Light and stronger than alluminium	Aircraft and bicycle parts

Topic 4: Stoichiometry

- **Stoichiometry** is the relationship between quantities of substances that take part in a chemical reaction or form a compound.

Writing chemical formulae of compounds

- Using charges on the cations and anions of both monoatomic and polyatomic ions, we can determine the formula of a compound that would be formed between two ions.

Rules

1. The positive ion is place first in the formula, the negative ion is second. eg, CuO
2. Positive and negative ions are combined so that the total number of positive charges is balanced by the total number of negative charges. eg, CuS, $CuCl_2$, and Al_2O_3.
3. For polyatomic ions, If more than one of the ions is used to balance the charge of a compound, then it is placed in brackets with the required number written as a subscript after the brackets. Eg. $Mg(NO_3)_2$

it is a mixture of two or more of metal which is formed by mixing thoroughly m

Steps

a. Write the chemical symbol of each element

eg Mg and Cl

b. Write the charge on each symbol on top and to the right of each symbol

Mg^{+2} Cl^{-1}

c. Exchange the charge numbers and write them below each symbol

$MgCl_2$

$Na + SO_4^{2-}$

Na_2SO_4

Practice

Write down the formula of the compounds formed between

1. Na and Cl
2. Na and O
3. Al and S
4. Ca and O
5. Na and SO_4^{2-}
6. Mg and NO_3^-
7. K and CO_3^{2-}
8. NH_4^+ and SO_4^{2-}
9. Ca and OH^-
10. Al and OH^-

Writing chemical equations

Steps in writing chemical equations

1. Identify reactants and products of a reaction

➢ **Reactants:** carbon and oxygen **Product:** carbon dioxide

2. Write down the word equation

Carbon + Oxygen → carbon dioxide

3. Write down the correct chemical symbols/formulae of the reactants and products

$C + O_2 \rightarrow CO_2$

Law of conservation of matter

- It states that matter is never created nor destroyed but only rearrange
- To conserve matter, equations have to be balanced.

Balancing chemical equations

- A chemical equation is balanced if numbers of each kind of atom are equal on both sides of the equation.
- Change the coefficients only and not the subscripts.
- The final coefficient of each chemical formula in the balanced chemical equation must be a whole number.

Steps to Balancing Equations

1. Write the correct formula for the reactants and the products. **Do Not Change the Formulas!**
2. Find the number of atoms for each element on both sides of the equation.
3. Change the coefficients to balance number of each kind of atoms
4. Check your answer to see if:
   - ➢ The numbers of atoms on both sides of the equation are now balanced.
   - ➢ The coefficients are in the lowest possible whole number ratios. (reduced)

Example:

Write down the balanced chemical equations of the following reactions

1. Reaction of magnesium with chlorine to produce magnesium chloride

magnesium + chlorine ⇒ magnesium chloride

$Mg + Cl_2 \Rightarrow MgCl_2$

Balancing: $Mg + Cl_2 \Rightarrow MgCl_2$

2. Reaction of Hydrogen and oxygen to produce water

Hydrogen + oxygen ⇒ water

$H_2 + O_2 \Rightarrow H_2O$

Balancing $2H_2 + O_2 \Rightarrow 2H_2O$

Practice work

Balance the following equations

a. $Na + Cl_2 \rightarrow NaCl$

b. $NO + O_2 \rightarrow NO_2$

c. $CH_4 + O_2 \rightarrow CO_2 + H_2O$

d. $Al + O_2 \rightarrow Al_2O_3$

e. $Na + H_2O \rightarrow NaOH + H_2$

b. Finding the RAM or the RFM of the substance.
c. Expressing the mass in grams.

Example: find the molar mass of ethanol

Formula: C_2H_5OH

RFM: $(2 \times 12) + (1 \times 6) + (1 \times 16) = 46$

Molar Mass = 44g

Practice

Calculate the molar masses of each of the following;

1. CO_2 2. NaCl 3. HCl 4. NaOH 5. $CaCO_3$

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

Mole Calculations

The following equations are used to solve problems involving moles.

a. Number of moles of atoms (N) $= \frac{\text{mass of an element in grams (m)}}{\text{relative atomic mass(RAM)}}$

$$N = \frac{m}{RAM} \rightarrow N = \frac{m}{\text{Molar Mass}}$$

Example: How many moles of aluminium atoms are in 5.4g of aluminium?

$$N = \frac{m}{RAM} = \frac{5.4}{27} = 0.2\text{moles}$$

b. Number of moles of molecules or ions $= \frac{\text{mass of substance in grams (m)}}{\text{Relative formula mass (RFM)}}$

$$N = \frac{m}{RFM} \rightarrow \frac{m}{\text{Molar Mass}}$$

Example

a. Work out number of moles in 1.6g of sodium hydroxide, NaOH.

$$N = \frac{m}{RFM} = \frac{1.6}{40} = 0.04\,moles$$

b. What is the mass of 0.02moles of carbon dioxide, CO_2?

$$N = \frac{m}{RFM} \rightarrow m = N \times RFM = 0.02 \times 44 = 0.88g$$

c. Number of moles of particles(N) $= \frac{\text{number of particles (P)}}{\text{avogadro constant (Av)}}$

$$N = \frac{P}{Av}$$

Example

a. Find number of atoms in 0.001moles sodium.

$P = N \times Av = 0.001 \times 6.02 \times 10^{23} = 6.02 \times 10^{20}$

b. Find number of moles in 3.01×10^{24} particles of oxygen.

$$N = \frac{P}{Av} = \frac{3.01 \times 10^{24}}{6.02 \times 10^{23}} = 5\,moles$$

6.02

d. From the above equations we also equate the RHS since LHS is the same to get;

$$\frac{P}{Av} = \frac{m}{RAM} \quad \text{and} \quad \frac{P}{Av} = \frac{m}{RFM}$$

Example

a. Find number of atoms contained in 8g of calcium.

$$\frac{P}{Av} = \frac{m}{RAM} \rightarrow P = \frac{m}{RAM} \times Av = \frac{8}{40} \times 6.02 \times 10^{23} = 1.204 \times 10^{23}$$

b. What mass of sulphur dioxide contain 1.505×10^{23} particle.

$$\frac{P}{Av} = \frac{m}{RAM} \rightarrow m = \frac{P \times RFM}{Av} = \frac{1.505 \times 10^{23} \times 64}{6.02 \times 10^{23}} = 16g$$

Practice

Work out

1. Number of mole in 2g of magnesium.
2. The mass of gold (Au) which contains the same number of atoms as 3g of carbon.
3. The number of moles in 10g of sulphur dioxide, SO_2.
4. Number of moles in 175.5g of sodium chloride, NaCl.
5. Number of atoms in 0.78g of potassium.
6. Number of molecules in 32g of sodium hydroxide, NaOH.
7. The mass of CuSO4 containing 0.2 moles.
8. The of mass of sodium in 6.0 x 10^{22} atoms

$\frac{P}{AV} = \frac{m}{RAM}$

$P = N \times$

$\frac{P}{AV} = \frac{m}{RAM}$

Molar volume of gases

- Volume occupied by 1 mole of any gas occupies at standard temperature and pressure (stp) and it is **22.4dm³**.
- Standard temperature is 0°C. Standard pressure is 1 atm.
- There is also molar volume of any gas at room temperature and pressure (rtp) and it is **24dm³**.
- Room/Normal temperature and pressure is 25°C and 1 atm respectively.

Example

Work out

\# of moles of particles = # of particles / avogadro constant

if 22.4dm³ = 2g

100dm³ = ?

a. The mass of 100dm³ of hydrogen gas, H_2 at stp.

1mole of H_2 at stp = **22.4dm³ = 2g.**

i.e. 22.4dm³ = 2g

therefore 100dm³ = x

by simple proportion $x = \frac{100dm^3 x\ 2g}{22.4dm^3} = 8.9g$

Des reflac reflec e

b. The volume occupied by 12g of oxygen gas at stp.

Molar mass for oxygen = 32g

Therefore 32g = 22.4dm³

12g = x

∴ $x = \frac{22.4dm^3 x\ 12g}{32g} = 8.4dm^3$

c. Number of moles in 4.8 dm³ of nitrogen gas, N_2 at rtp.

At rtp: 24dm³ = 1mole

∴ 4.8dm³ = x

$x = \frac{4.8dm^3\ x\ 1mole}{24dm^3}$ $0.2moles$

Practice

Work out

a. The mass of 2dm³ of suphur dioxide at stp.
b. The volume occupied by 196g of HCl gas at stp.
c. Number of moles of 72dm³ of NH_3 gyas at rtp.
d. Volume of 0.001 moles of CO_2 gas at rtp.

Water of crystallisation

- Amount of water incorporated into the structure of a substance as it crystallises expressed as a percentage.

i.e. percentage water = $\frac{mass\ of\ water}{mass\ of\ the\ salt\ hydrate}$ x **100%**

Example

Calculate the percentage of water crystallization in sodium carbonate decahydrate ($Na_2CO_3.10H_2O$). RAM of Na = 23, C = 12, O = 16 and H = 1).

percentage water = $\frac{mass\ of\ water}{mass\ of\ the\ salt\ hydrate}$ x **100%**

mass of water = 10 x 18 = 180g

mass of $Na_2CO_3.10H_2O$ = (2 x 23) + (1 x 12) +(3 x 16) + (180) = 286g

therefore %water = $\frac{180g}{286}$ x 100% = 62.95

Practice

a. Work out the percentage of water in hydrated copper sulphate pentahydrate ($CuSO_4.5H_2O$). (RAM: Cu = 64, H = 1, S = 32, O = 16).
b. Calculate the percentage of water of crystallization in magnesium sulphate heptahydrate, $MgSO_4 \cdot 7H_2O$. How much magnesium sulphate would be in 12g of $MgSO_4 \cdot 7H_2O$? (RAM: Mg = 24, H = 1, S = 32, O = 16)

Experiments

a. Heating sugar

Aim: To determine percentage of water in sugar

Materials: sugar, a beaker, tripod stand, wire gauze, a burner, and a balance.

Procedure

a. Weigh the empty beaker and record the mass in the table below.
b. With the beaker still on the balance, add 10 g of sugar.
c. Record the mass of sugar in the table below.
d. Heat the sugar in the beaker until a dry, black solid (carbon) is formed.
e. Weigh the beaker + black substance and record the mass in the table of results.
f. Calculate the mass of carbon (mass of beaker + carbon – mass of empty beaker).
g. Record mass of carbon in the table.
h. Calculate mass of water lost ie (mass of sugar – mass of carbon)

Table of results

Item	Mass (g)
Empty beaker	
sugar	10
Beaker + carbon	
Carbon	
Lost water	

i. Calculate the percentage composition by mass of water in sugar.

b. Heating hydrous copper sulphate

Aim: To determine the percentage of water in hydrated copper sulphate

Materials: Hydrated copper sulphate, evaporating basin, burner,balance, tripod stand and wire gauze.

Procedure

a. Weigh the empty basin and record its mass ~~in~~ in the table of results.
b. Add 10g of hydrated copper sulphate.
c. Record the mass of the basin plus the hydrated copper sulphate.
e. Heat copper sulphate gently until it turns into a white powder.
f. Weigh the basin plus white powder (anhydrous copper sulphate) and record the mass in the table results.
g. Work out the mass of the white powder (i.e. (f – a)).
h. Find the mass of water lost ie. (b - g)

Table of Results

Item	Mass (g)
Empty basin	
hydrated copper sulphate	10
Basin + anhydrous copper sulphate	
Anhydrous copper sulphate	
Water lost	

i. Calculate the percentage of water in the hydrated copper sulphate.

Empirical formulae

- Simplest formula that shows the lowest ratio of the atoms that make up a compound. Eg

Compound	Molecular Formula	Empirical Formula
hexene	C_6H_{12}	CH_2
glucose	$C_6H_{12}O_6$	CH_2O
Ethanoic acid	CH_3COOH	CH_2O

Working out empirical Formula

Steps

a. Determine the percentage by mass in grams of each element.
b. Calculate the atomic proportion (number of moles) of each element.
c. Work out the simplest more ratios. (Divide the atomic proportion by the lowest value among them).
d. Express the formula using the simplest atomic ratios.

Example

Work out the empirical formula of a compound that has the following percentage composition by mass of elements : C=40%, H=6.67%, and O=53.33%. (RAM C=12, H=1, 0=16)

Solution

Element	% composition	Atomic proportion	Simplest ratio
C	40	$\frac{40}{12} = 3.33$	$\frac{3.33}{3.33} = 1$
H	6.67	$\frac{6.67}{1} = 6.67$	$\frac{6.67}{3.33} = 2$
O	53.33	$\frac{53.33}{16}$ 3.33	$\frac{3.33}{3.33} = 1$

Empirical formula is CH_2O

Working out molecular formulae

- Molecular formula shows the actual numbers of the atoms making up a compound.
- To find the molecular formula of a compound we multiply number of each kind of atoms in the empirical formula by a number N obtained from the ratio of molar mass to empirical formula mass.

Example

In an experiment an unknown organic compound was found to contain 0.12g of carbon and 0.02g of hydrogen. If the RFM of the compound is 56, calculate the molecular formula. 12 + 2

Solution

Element	Composition (g)	Atomic proportion	Simplest ratio
C	0.12	$\frac{0.12}{12} = 0.01$	$\frac{0.01}{0.01} = 1$
H	0.02	$\frac{0.02}{1} = 0.02$	$\frac{0.02}{0.01} = 2$

Empirical formula = CH_2

$$N = \frac{Molar\ mass}{Empirical\ formula\ mass} = \frac{56}{14} = 4$$

C_4H CH_2

Molecular formula = $(CH_2)_N = (CH_2)_4 = C_4H_8$

$4(CH_2)$ 12+2 14

Practice

1. Work out the empirical formula of the compound with the following percentage composition: P = 20.19%, O = 10.42%, Cl = 69.39%.
2. A hydrocarbon has a percentage composition by mass of 85.7% of carbon. Calculate the empirical of the hydrocarbon. (RAM; C = 12, H = 1)
3. 1.2g of magnesium combines with 3.55 g of chlorine. If the formula mass is 94g, work out the molecular formula of the compound. (RAM; Mg = 24, Cl = 35.5).
4. An alkane has percentage composition is 84.2 % carbon and 15.8% hydrogen. Its formula mass is 114g. What is its molecular formula? (RAM; C = 12, H = 1)
5. In an experiment 2.6g of chromium were heated in an excess of chlorine gas. It was found that 7.93g of chromium chloride was formed. If RFM of chromium chloride is 158, calculate molecular formula of chromium chloride.

Concentration of a solution

- Amount of solute dissolved in a specific volume of solvent.
- Amount of a Substance in a unit Amount of Solution

Way of expressing the concentration of a solution

1. **Mass per unit volume as;**

   Concentration = $\frac{mass\ of\ solute}{volume\ of\ solution}$

   Units are g/litre, g/cm^3 or g/dm^3

2. **Mass per unit mass as:**

   Concentration = $\frac{mass\ of\ solute}{mass\ of\ solution}$

   Units: grams of solute/100g of solution

3. Percentage concentration

   a. % W/W = $\frac{mass\ of\ solute}{mass\ of\ solution}$ X 100

   b. % V/V = $\frac{Volume\ of\ solute}{volume\ of\ solution}$ X 100

4. *Molarity (M)*, or ***molar concentration***

   Defined as the number of moles of solute per liter of solution

   Molarity = $\frac{moles\ of\ solute}{litres\ of\ solution}$

   Units: moles/l, $moles/dm^3$

   When one mole of solute is dissolved and made into $1dm^3$ of a solution, the solution is called a **molar solution.**

   ie $\frac{1mole}{1dm^3}$ **= 1 molar = 1M**

Other important equations

i. Concentration (g/dm^3) = Molarity X RFM

ii. Mass (g) = Molarity X Volume X RFM

Worked Examples

1. Calculate the concentration in (g/cm^3) of solution containing;

   a. 9.8g of sulphuric acid dissolved in water and made into $500cm^3$.

      C = m/V = $9.8g/500cm^3$

      = $0.02g/cm^3$

   b. 0.2 mole of sodium hydroxide solution dissolve in water and made up to $100cm^3$

      M = N/V = 0.2mol ÷ $\frac{100cm^3}{1000cm^3}$ = 2M

      C = M x RFM = $2mol/cm^3$ x 40g/mol = $80g/cm^3$

2. Calculate the mass of potassium nitrate (KNO_3) which needs to be used to prepare $200cm^3$ of a $2mol/dm^3$.

   - m = M x V x RFM = $2mol/dm^3$ X $0.2dm^3$ X 101g/mol = 40.4g

$M = \frac{\#\text{ of mole}}{\text{volume}}$

3. How many moles are contained in 250cm³ of of 2M HCO_3

   $N = V \times M = \frac{250cm^3}{1000cm^3} \times dm^3 \times 2mol\ dm^{-3}$

   = 0.5moles

4. Calculate the molarity of a solution in which 5.85g NaCl is dissolved in 4dm³.

   - N = m/RFM = 5.85/ 58.5 = 0.1mols

     M = N/V = 0.1mole/4dm³ = 0.025M

5. How many moles are in 3000cm³ 2M nitric acid solution?

   $N = M \times V = 2M \times \frac{3000cm^3}{1000} = 6moles$

6. What volume of pure ethanol must be used to make 500ml of a 70% ethanol solution?

   - 70% ethanol means 70ml per 100ml solution

     Required volume $= \frac{70ml}{100ml} \times 500ml = 340ml\ of\ ethanol$

Practice Questions

1. What is the molarity of a solution containing 0.25moles in 10cm³?
2. How many moles are in 500cm³ of 0.25M sodium hydroxide solution?
3. Calculate the concentration in g/dm³ 0.025 moles of anhydrous sodium carbonate(Na_2CO_3) dissolved in 2000cm³.
4. Calculate the mass in grams used in making 80cm³ of 0.5M sodium hydroxide.
5. Calculate the molarity of a solution containing 6.3g HNO_3 in 250cm³.

Standard Solution

- A solution whose concentration is Known
- Standard solution can easily be used in a laboratory for a specific purpose
- It is also used to determine the concentration of other unknown solutions.

Preparing Standard Solutions

1. **Dissolution of solutes (i.e. dissolving Solutes)**

- A solid solute of a measured mass is dissolved into a solution of known volume.
- To prepare a standard solution by solute dissolution we need to know the concentration and the volume of the solution to be prepared.

Procedure

a. Calculate the required mass of the solute using the equation:

   Mass = molality x volume x RFM

b. Weigh out the exact mass calculated using a beam/electronic balance.

c. Dissolve the solute completely in a beaker.

d. Transfer the solution quantitatively into a volumetric flask ie nothing must be left in the beaker.

e. Rinse the beaker thoroughly into the flask.

f. Fill the flask with more distilled water up to the up to the mark, put the stopper and shake to mix thoroughly.

Example

- Describe how you can prepare 250ml of 2M Potassium carbonate (K_2CO_3) solution. (RAM; K = 39, C =12, O = 16)

  Working out

- Calculate the required mass of $K_2 CO_3$ as follows;

  M = M x V RFM = 2M x O.25dm³ x 138g/mol = 69g

- Weigh out 69g of K_2CO_3 in a beaker using.
- Add just enough distilled water into the beaker to dissolve the pellets.
- Transfer the solution into the volumetric flask quantitatively.
- Rinse the beaker with distilled water three times adding into the volumetric flask.
- Add distilled water up to the 250ml mark and shake well to mix thoroughly.

  Exercise

1. Describe how you can prepare 200ml of 0.2M NaCl solution.
2. Describe how 250cm³ of a 1M copper sulphate solution could be prepared using hydrated copper sulphate crystals ($CuSO_4 - 5H_2O$).

2. **Dilution of stock solutions**

- Stock solutions are solutions kept so that other standard solutions can be prepared from them.
- These standard solutions are highly concentrated
- Dilution means making a concentrated solution weaker by adding distilled water to them.
- Dilution equation
- Used to solve for the volume of the stock solution required to prepare a solution of a certain concentration. It is given by:

  $C_1 V_1 = C_2 V_2$

  Where

  ✓ C_1 is concentration of the stock solution

  ✓ V_1 = is volume to be transferred from the stock solution (ml or dm3)

  ✓ C_2 is concentration of the solution to be prepared

Volume = cm³ or dm³

✓ V_2 = is volume of the solution to be prepared.

Example

Work out the volume of 0.8M NaOH required to prepare 250ml solution of 0.2M NaOH .

Solution

$C_1 V_1 = C_2 V_2 \Rightarrow V_1 = \frac{C_2V_2}{C_1} = \frac{0.2M \times 250ml}{0.8M} = 62.5ml$

Steps to be followed

a. Calculate the volume, V_1 to be transferred from the stock solution using the equation.

$$V_1 = \frac{C_2V_2}{C_1}$$

b. Measure the calculated volume of the stock solution using a measuring cylinder.

c. Transfer it into an appropriate volumetric flask.

d. Rinse the measuring cylinder with some distilled water and transfer the water into the volumetric flask.

e. Add more distilled water to the volumetric flask up to the mark,

f. Put the stopper and shake well to mix.

Example

- Describe how you can prepare 250 cm³ of 0.2 M hydrochloric acid from 2M hydrochloric acid solution.

Working out

- Calculate the volume (V_1) of the 2M hydrochloric solution to be taken as follows:

$V_1 = \frac{C_2V_2}{C_1} = \frac{0.2M \times 250ml}{2M} = 25cm^3$

- Pour out 25cm³ of the 2M hydrochloric acid and transfer it into a 250cm³ volumetric flask.
- Add distilled water up to the 250cm³ mark and shake to mix.

Exercise

1. Describe how you can prepare 500ml of 0.555M HCl from 11.3M HCl.
2. The volume of sodium hydroxide solution (NaOH) of concentration 20g/l is increased from 60cm³ to 600cm³ by adding distilled water. Calculate the concentration of the new solution g/l.
3. 60cm³ of a solution whose concentration of 15 g/cm³ was diluted with distilled water by raising its volume to 80 cm³. Calculate the concentration of the new solution.

1. 61 73.4 1.3/15

1. 0.2

4. Describe how you can prepare 500ml of 1.75M solution of sulphuric acid using 8.16M solution of sulphuric acid, H_2SO_4. (RAM: H=1, S=32,O=16).

Determination of Concentration of a Solution by Titration

Titration

- Titration is the gradual addition of one liquid to another until the reaction is completed.
- The best known example of this is neutralisation reaction in which an acid is slowly added to a base.
- A simple acid-base titration involves the addition of an acid solution to an alkali (a base) solution until the reaction is complete.
- When enough acid has been added to the base to neutralise it we say that we had reached the end point or neutral point.
- To know exactly when the end point has been reached, an indicator is used. A commonly used indicator is phenolphthalein.
- An indicator is a substance used to show whether a substance is acidic or alkaline (basic).
- During titration there are two different solutions;

  A titrant : is a standard solution

  The analyte: the solution whose concentration is to be determined.

Quantities in a titration process

- ✓ Molarity of an acid M_1
- ✓ Volume of an acid V_1
- ✓ Number of moles of an acid in a balanced chemical equation, N_1
- ✓ Molarity of a base M_2
- ✓ Volume of a base V_2
- ✓ Number of moles of a base in a balanced chemical equation, N_2

- The equation below uses the above quantities to determine the molarity and volume of solutions:

$$\frac{M_1V_1}{N_1} = \frac{M_2V_2}{N_2}$$

Example

25.0cm3 of 0.1M NaOH were neutralised by 20.0cm^3 of HCl of molarity M_2 from the burette. Calculate the molarity of the acid (M_2).

Solution:

$$HCl + NaOH \rightarrow NaCl + H2O$$

$$\frac{M_1V_1}{N_1} = \frac{M_2V_2}{N_2} \Rightarrow \frac{0.1M \times 25.0}{1} = \frac{M_2 \times 20.0}{1} = 0.125M$$

Exercise

1. What volume in cm^3 of 0.3M HCl will neutralise $50.5cm^3$ of 0.1M $Ca(OH)_2$?
2. The molarity of a solution of HCl is found by titrating it against 0.25M Na_2CO_3 solution. The result was that $25cm^3$ of 0.25M Na_2CO_3 solution required $20.8cm^3$ of the acid of unknown molarity M_2. Work the concentration of the acid in
   (a) moles/dm^3
   (b) grams/dm^3.
3. What volume in cm^3 of 0.3M HCl will neutralise $10cm_3$ of $Al(OH)^3$? The balanced equation for the reaction is:

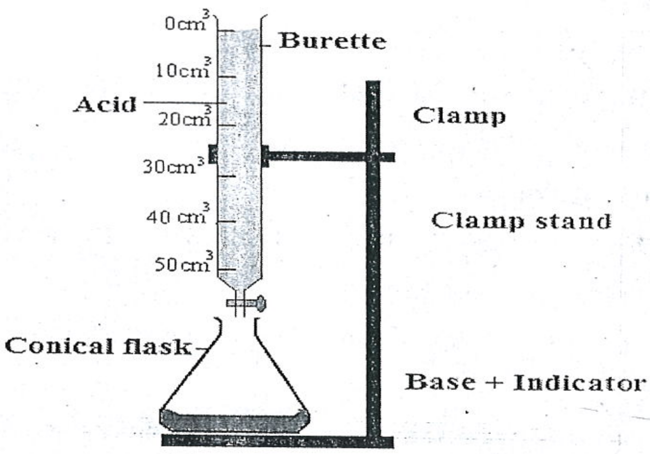
   $$3HCl + Al(OH)_3 \rightarrow AlCl_3 + 3H_2O$$
4. $22.4cm^3$ of a solution containing $0.10moldm^{-3}$ of sulphuric acid just neutralised $25.0cm^3$ of a sodium hydroxide solution. What is the concentration of this sodium hydroxide solution?

Apparatus and materials for the titration process

- a burette, clamp and clamp stand, measuring cylinder, filter funnel, conical flask, phenolphthalein indicator, standard solution and a solution of unknown concentration.

Titration procedure

a. Arrange the apparatus as shown figure below;



b. Put a known volume of the base into a conical flask.
c. Adding a few drops of indicator into the flask and note the colour change.
d. Pour the acid into the burette.

e. Drip the acid from the burette into the conical flask slowly, while swirling the flask continuously.
f. Stop when there is a colour change of the indicator. This is the end point.
g. Record the new level of the solution in the burette.
h. Find the volume of the solution used. (Volume used = final volume – initial volume),
i. Calculate the concentration of the solution whose concentration is not known.

Sources of Error in Titration

1. Misjudging the colour of the indicator near the end point. Different people have different sensitivity to colours.
2. Misreading the volume
3. Using contaminated solutions
4. Using wrong amount of indicator.
5. Not filling the burette properly
6. Titrating at wrong temperature
7. Vigorous swirling making the solution to splash from the titration conical flask leading to losing the solution.
8. Leaking burette.

Determination of Limiting Reagent and Excess Reagent

Limiting Reagent

- Is the reactants in a chemical reaction that limit the amount of production that can be formed.
- Used up in a reaction implying that the reaction will automatically stop when all the limiting reactant is consumed.

Excess Reagent/ Reactant:

- Is the reactant in a chemical reaction that remains when a reaction after the limiting reactant is completely consumed.
- This means that the excess remains because there is nothing with which it can react.

First approach

- Calculate moles or grams of one reactant A needed to completely consume a second reactant B. if A available is more than required, then it is the excess reagent, if not, then it is the limiting reagent.

Example

Determine the limiting reagent and excess reagent for the synthesis of urea.

$2NH_3 + CO_2 \rightarrow (NH_2)_2CO + H_2O$

Given 7.481s NH_3 and 7.992g CO_2

Solution

Starting with NH_3 we calculate the minimum grams of CO_2 needed to complete 7.481g of NH_3.

From the balanced equation

$2NH_3 + CO_2 \rightarrow (NH_2)_2CO + H_2O$

34g 44g

7.481 X

$$X = \frac{7.481 \times 44}{34} = 9.68$$

- Because the 7.992 CO_2 available is less than the 9.666g needed, CO_2 is the limiting reagent and NH_3 is the excess reagent.

Second approach

- Calculate moles or grams of the product that can be obtained when each reagent is completely used up.
- The actual limiting reagent is the reactant producing the least amount of product.

Example

Determine the limiting reagent and excess reagent for the synthesis of urea

$$2NH_3 + CO_2 \longrightarrow (NH_2)_2CO_2 + H_2O$$

Given 7.481s NH_3 and 7.992 CO_2

Solution

From the balanced equation

$2NH_3 + CO_2 \rightarrow (NH_2)_2CO + H_2O$

34g 44g 60g 18g

7.481 7.992 x

34g of ammonia produce 60g of urea. 7.481g of ammonia produces xg of urea

Using simple proportion we have

$$x = \frac{7.481 \times 60}{34} = 13.2g$$

In the same way 44g of CO_2 produce 60g of urea. 7.992g of CO_2 produces xg of urea

Using simple proportion we have

$$x = \frac{7.992 \times 60}{44} = 10.89g$$

Because 7.992g CO_2 generates less Urea that 7.481g NH_3 is the limiting reagent and NH_3 is the excess reagent.

Exercise

1. What mass of hydrogen gas at STP is produced from the reaction of 50.0g of Mg and 75.0 grams of HCl? How much of the excess reagent is left over (in grams)?
2. What mass of calcium sulfate can be produced from the reaction of 1000 g calcium phosphate with 980 g H_2SO_4? Identify the limiting reagent.
3. A 2.00 g sample of ammonia is mixed with 4.00 g of oxygen. Which is the limiting reactant and how much excess reactant remains after the reaction has stopped?
4. If 192 grams Ca mixed with 56 grams N_2, which is the limiting reactant?
5. When 4.00 mol H_2 is mixed with 2.00 mol Cl_2,how many moles of HCl can form? Identify the limiting reactant.
6. Find the limiting reagent when 1.22g O_2 reacts with 1.05g H_2 to produce H_2O.

Theoretical Yields and Percentage Yields

Theoretical yield

- Maximum amount of product expected when the limiting reagent is completely used up.
- Predicated by the balance equation.

The actual yield

- The amount of product actually obtained for a reaction.
- Usually it is less than the theoretical yield.

Percentage yield

Is the ratio of actual yield to theoretical yield, multiplied by 100: or is the actual yield expressed as a percentage of the theoretical yield

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

Percentage yield is used to indicate how much desired production is obtained from a reaction.

Consider the preparation of nitrobenzene $C_6H_5NO_2$, by the reaction of a limited amount of benzene C_6H_6, with excess nitric acid, HNO_3. The balanced equation for the reaction may be written as

Example

A 15.6gram sample of benzene (C_6H_6) is mixed with excess nitric acid HNO_3. We isolate 18.0grams of nitrobenzene, $C_6H_5NO_2$. What is the percentage yield of $C_6H_5NO_2$ in this reaction?

Solution

C_6H_6	+	HNO_3	$\rightarrow$	C_6H_5NO	+	H_2O
1mol		1mol		1mol		1mol
78.g		63.0g		123.0g		18.0g

Theoretical yield of $C_6H_5NO_2$

Yield = $\frac{15.6 \times 123}{78}$ = 24.6g

Percentage yield = $\frac{\text{Actual yield of product}}{\text{Theoretical yield of product}}$ x 100%

= $\frac{18.0g}{24.6g}$ X 100% = **73.2 percentage yield**

Exercise

1. When copper is heated with an excess of sulfur, copper(I)sulfide is formed. In a given experiment, 1.50 g copper was heated with excess sulfur to yield 1.76 g copper(I) sulfide. What is the theoretical yield? What is the percent yield?
2. Calcium carbonate is decomposed by heating, as shown in the following equation. $CaCO_3 \rightarrow CaO + CO_2$
   a. What is the theoretical yield of CaO if 24.8 g $CaCO_3$ is heated?
   b. What is the percent yield if 13.1 g CaO is produced?
3. When 84.8 g of iron (III) oxide reacts with an excess of carbon monoxide, 54.3 g of iron is produced. What is the percent yield of this reaction?
   $Fe_2O_{3(s)}$ + $CO_{(g)}$ $\rightarrow$ $Fe_{(s)} + CO_{2(g)}$
4. What is the % yield of NH_3 if 40.5 g NH_3 is produced from 20.0 mol H_2 and excess N_2?
5. What is the % yield of ferrous sulfide (FeS) if 3.00 moles of Fe reacts with excess sulfur to produce 220 grams of ferrous sulfide?
   $Fe + S \rightarrow FeS$

$Cu^{2+} + S_{(s)} \rightarrow CuS_{(s)}$

Topic 5: Heats of Reactions

- During chemical reactions, heat energy is absorbed from the surrounding to break the existing bonds and heat energy is released to the surrounding when new bonds are formed.
- The difference between the heat energy taken in to break the existing bonds and the heat energy given out when the new bonds are formed during chemical reactions is called **heat of reaction or Enthalpy change denoted ΔH.**

  ΔH = heat energy taken in - heat energy given out.

Exothermic and Endothermic Reactions

- There are two types of heat of reactions

1. Exothermic Reaction

- It is a reaction that produces and releases or gives out heat to the surrounding.
- During such reactions, temperature rises which cause the container to heat up.
- The enthalpy change is negative ($-\Delta H$) indicating that heat content decreases from reactants to products during the reaction.
- Examples include

$\Delta H =$

1. Neutralization reaction such as:

   $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaOH_{(aq)} + H_2O_{(l)}$.
2. Mixing of concentrated sulphuric acid and water.
3. Combustion reaction such as:

   $C_{(s)} + O_{2\,(g)} \rightarrow CO_{2(g)}$.
4. Metal and acid reactions for example:

- $Mg_{(s)} 2HCl_{(aq)} \rightarrow MgCl2_{(aq)} + H_{2(g)}$

Experiment

Aim: determining temperature changes in exothermic reaction.

Materials: Sodium hydroxide pellets (NaOH), 0.5M Hydrochloric acid (HCl), 100ml beaker, and distilled water.

Procedure

- Add 50ml of water into the beaker
- Note and record the initial temperature
- Using a hand, feel the beaker, and determine whether it is warm or cold.
- Add a full spatula of sodium hydroxide pellets into the water.
- Note and record the temperature
- Using a hand, feel the beaker, and determine whether it is warm or cold.

- Record your results in the table below.

	Water only	Water +NaOH
Temperature		
Cold/warm		

- What is the conclusion of the experiment?

2. **Endothermic Reaction**

- A reaction which takes in heat energy from the surroundings.
- The value of the heat of reaction for an endothermic reaction has a positive sign ($+\Delta H$).
- Temperature decreases during such reactions and the container feels colder.
- Examples of endothermic reactions;

1. Photosynthesis.

$$\text{Sunlight} + 6CO_{2(g)} + H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + 6O_{2(g)}$$

2. The dissolving of ammonium nitrate in water.

$$NH_4NO_{3(s)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + NO^-{}_{3(aq)}$$

3. The evaporation of water,

$$H_2O_{(l)} + \text{Heat} \rightarrow H_2O_{(g)}$$

Experiment

Aim: Determination of temperature changes in an endothermic reaction.

Materials: Ammonium chloride, distilled water, 100ml beaker, thermometer, and stirrer.

Procedure

- Pour 50ml of distilled water into the beaker
- Measure the initial temperate using the thermometer.
- Add two full spatula of ammonium nitrate to the 50ml of distilled water in the beaker, and stirrer until all the nitrates are dissolved.
- Using the thermometer, measure the final temperature
- Present your results in the table below.

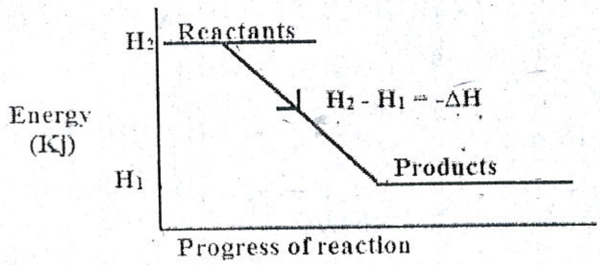
	Water	Solution ($NH_4NO_{3(s)} + H_2O_{(l)}$)	Change in temperature ^0C (final - initial)
Temperature			

- What is the conclusion of the experiment?

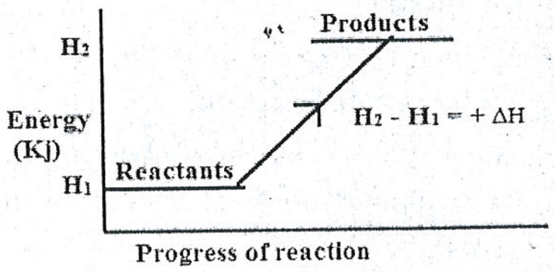
Heat Energy Diagrams

- Diagrams that plot heat energy of reactants and products against time as the reaction progress.

- For an exothermic reaction, the reaction moves from a high energy level to a low energy level while for an endothermic reaction, the reaction moves from a lower energy level to a higher energy level.



Exothermic Reaction



Endothermic Reaction

- The heat of reaction equals the change in energy given by:

$$\Delta H = \text{Final heat} - \text{Initial heat}$$

For exothermic reaction,

$$\Delta H = H_1 - H_2 = -\Delta H$$

- Therefore, energy change for exothermic reactions is negative

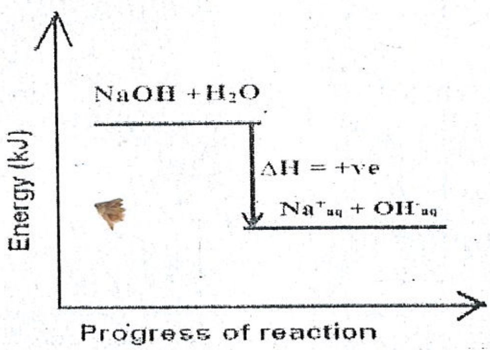
For an endothermic reaction,

$$\Delta H = H_2 - H_1 = +\Delta H$$

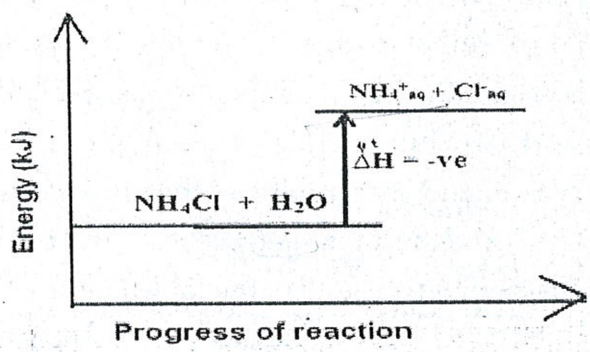
- Therefore, energy change for endothermic reactions is positive

Example

- Below are energy level diagrams for the dissolving of sodium hydroxide in water (a) and dissolving of ammonium chloride in water (b).



a



b

Heats of Reaction in Equations

- There are two ways of writing chemical equations which include the Heat of reaction, ΔH.

1) The chemical equation is written out in full, and the heat of reaction (ΔH), is written separately using a comma at the end of the equation.

- For exothermic reaction;

  $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}, \quad \Delta H = -18KJ$

- For endothermic reactions

  $H_2O_{(l)} \rightarrow H_2O_{(g)}, \ \Delta H = +5kJ$

2) The heat of reaction is written as a part of the reaction.

   - For exothermic it is shown as one of the products.

     $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)} + 18KJ.$

   - For endothermic reactions, it appears as one of the reactants

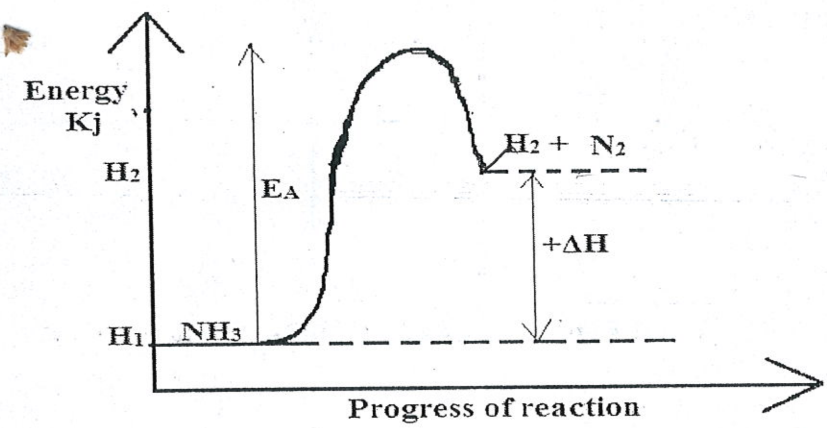
     $5kJ + H_2O_{(l)} \rightarrow H_2O_{(g)}$

Bond Energies

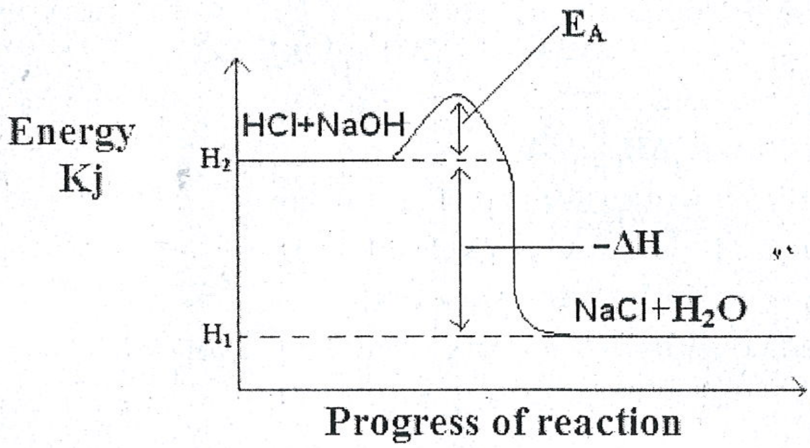
- Bond energy is the energy required to break or make one mole of bonds. It is measured in kJ/mole.
- High bond energy means that the bond is strong; hence more energy is needed to break or form it.
- The amount of energy required to break a bond is the same as the amount of heat energy given of when the same bond is formed.

Bond Making and Bond Breaking

- Bond breaking requires use of heat energy from the surrounding, therefore the process is endothermic.
- Bond making releases heat energy to the surrounding, therefore the process is exothermic.
- A reaction is endothermic when more heat energy is absorbed to break the bonds than heat energy released during bond making.



- A reaction is exothermic when less heat energy is absorbed to break the bonds and more heat energy is released during bond making.



✓ A certain amount of energy is absorbed to start the reaction which breaks the bonds of the reactants. This amount of energy is called **activation energy**, E_A.

Working Out Heat of reaction using Bond Energies

- In a reaction both bond breaking and bond making take place simultaneously.
- To determine the overall energy change (heat of reaction), we calculate the difference between energy used in bond breaking and energy released during bond making for a particular reaction. Ie

  Heat of reaction (ΔH) = energy in – energy out
- Below is a table showing some bonds and their bond energies,

Bond	Bond Energy Kj/mol
C – C	346
C = C	612
C – O	358
O = O	498
O – H	464
H – H	436
C – H	435

Bond	Bond Energy KJ/mol
N – H	391
Cl – Cl	242
H – Cl	431
C = O	803
C – Cl	339
N ≡ N	946

Example 1

Calculate the heat of reaction for the reaction below. Name the type of reaction

$$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$$

Solution

Energy in to break bonds.

$1 \times 436 = 436kJ$

$1 \times 242 = 242kJ$

Total energy = 678kJ

Energy from bond formation = 2 x 431kJ = 862kJ

Heat of reaction (ΔH) = energy in – energy out = (678 – 862) kJ $\Delta H = -184kJ$

✓ The overall reaction takes in 184kJ less energy than it gives out.

✓ Therefore the reaction is exothermic reaction.

Example 2

Calculate the heat of reaction for the complete combustion of ethanol C_2H_5OH.

Solution

First, write down a balanced chemical equation for the reaction

$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

Energy absorbed

In C_2H_5OH. = 5C–H + O–H + C–O = 5(435) + 464 + 358 = 2997

In O_2 =3[O=O] = 3(497) =1491

Total energy absorbed = 4488Kj

Energy released

For CO_2 =2[2(C=O)]=2[2x803] = 3212

For H_2O =3[2(H–O)] =3[2x464] = 2784

Total energy absorbed =5996KJ

ΔH= Energy absorbed – energy released =4488 – 5996 = –1508KJ

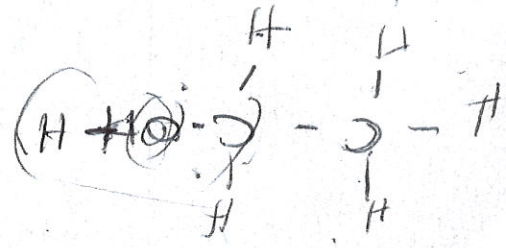
✓ Therefore the reaction is exothermic reaction.

Practice work

For the reaction represented by the chemical equations below, determine the heat of reaction and identify the type of reaction in terms of enthalpy change

1. $CH_4 + O_2 \rightarrow CO_2 + H_2O$
2. $H_2O \rightarrow H_2 + O_2$
3. $CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$
4. $NH_3 \rightarrow N_2 + H_2$

[Hint balance the equations first]



Topic 6: Alkanols/ Alcohols

- These are organic compounds called Oxycarbons as they contain an oxygen atom apart from the carbon and the hydrogen atoms.
- They are also called alcohols.

Functional group of alkanols

- The functional group of the alkanols is a hydroxyl group, -OH.

Common nomenclature

- All names of the members of alkanols end with a suffix –anol. The table below shows the names for the first ten Alkanols.

No of c-atoms	Prefix	Suffix	Name
1	Meth-	-anol	Methanol
2	Eth-	-anol	Ethanol
3	Prop-	-anol	Propanol
4	But-	-anol	Butanol
5	Pent-	-anol	Pentanol
6	Hex-	-anol	Hexanol
7	Hept-	-anol	Heptanol
8	Oct-	-anol	Octanol
9	Non-	-anol	Nonanol
10	Dec-	-anol	Decanol

General formula and Molecular formulae

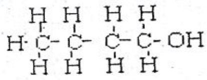
- The general formula for alkanols is $C_nH_{2n+1}OH$. This formula is used to work out molecular formula for alkanols.
- A molecular formula is a formula which shows the number and the kinds of atoms in a molecule.
- For instance, when n = 4 we have:

  $C_nH_{2n+1}OH \quad = \quad C_4H_{2(4)+1}OH \quad = \quad C_4H_9OH$

- The molecular formular can be used to work out the Structural, condensed and skeletal formulae of all the alkanols

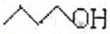
Structural formula

- Shows the arrangement of atoms and their bonds. For we C_4H_9OH have;



Skeletal formula

- A short hand representation of bonding and some detail about its molecular geometry. It is sometimes called line-angle formula. Eg

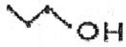
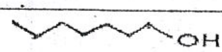
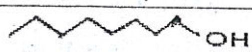
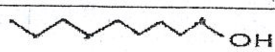
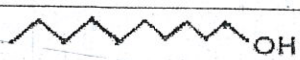
   (Butanol).

Condensed formula

- Atoms are shown in the way as they appear in the structure; eg for butanol we have;

  $CH_3CH_2CH_2CH_2OH$

- The table below shows the molecular, structural, condensed and skeletal formulae for the first ten alkanols:

N	Molecular Formula	Structural	Condensed	Skeletal
1	CH_4	H H-C-OH H	CH_3OH	
2	C_2H_6	H H H-C-C-OH H H	CH_3CH_2OH	
3	C_3H_8	H H H H-C-C-C-OH H H H	$CH_3CH_2CH_2OH$	
4	C_4H_{10}	H H H H H-C-C-C-C-OH H H H H	$CH_3(CH_2)_3OH$	
5	C_5H_{12}	H H H H H H-C-C-C-C-C-OH H H H H H	$CH_3(CH_2)_4OH$	
6	C_6H_{14}	H H H H H H H-C-C-C-C-C-C-OH H H H H H H	$CH_3(CH_2)_5OH$	
7	C_7H_{16}	H H H H H H H H-C-C-C-C-C-C-C-OH H H H H H H H	$CH_3(CH_2)_6OH$	
8	C_8H_{18}	H H H H H H H H H-C-C-C-C-C-C-C-C-OH H H H H H H H H	$CH_3(CH_2)_7OH$	
9	C_9H_{20}	H H H H H H H H H H-C-C-C-C-C-C-C-C-C-OH H H H H H H H H H	$CH_3(CH_2)_8OH$	
10	$C_{10}H_{22}$	H H H H H H H H H H H-C-C-C-C-C-C-C-C-C-C-OH H H H H H H H H H H	$CH_3(CH_2)_9OH$	

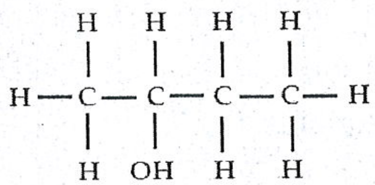
IUPAC Naming

- Involves showing the position of the functional group in the structure eg

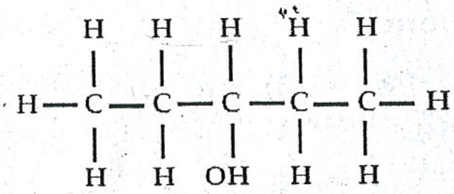
~~international union of pure applie uprite chemistry~~

IUPAC

international union of pure and upright chemistry



Butan-2-ol



Pentan-3-ol

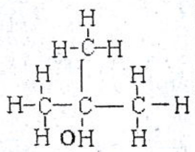
Writing and Naming branched chain alkanols

Rules

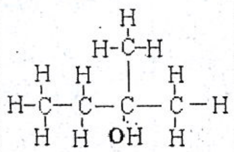
- Select the longest chain called the parent chain (longest continuous chain of carbon atoms) that MUST include the functional group.
- Number the parent chain from the end nearest the functional group.
- Identify and name alkyl groups attached to the longest chain.
- Name the compound, listing the alkyl groups in alphabetical order.
- Prefixes di, tri, tetra are used to indicate presence of 2, 3 and 4 alkyl groups respectively

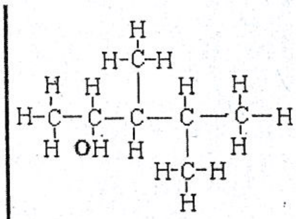
Example

Name the following compounds

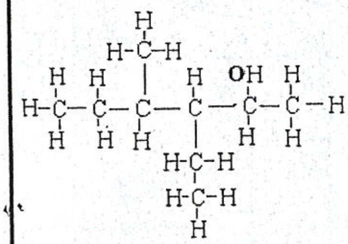


2methylproan-2-ol



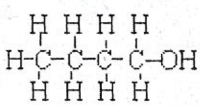


3,4dimethylpentan-2-ol

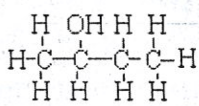


Types of Alkanols

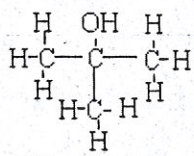
1. **Primary (1°) Alkanols**
   - The 'C' with the –OH group is bonded to one other C.
2. Secondary (2°) alkanols
   - The 'C' with the –OH group is bonded to two other carbon atoms.
3. In tertiary, (3°) alkanols
   - The 'C' to with the –OH group is bonded to three other C atoms.



Primary



Secondary



Tertiary

Ethanol

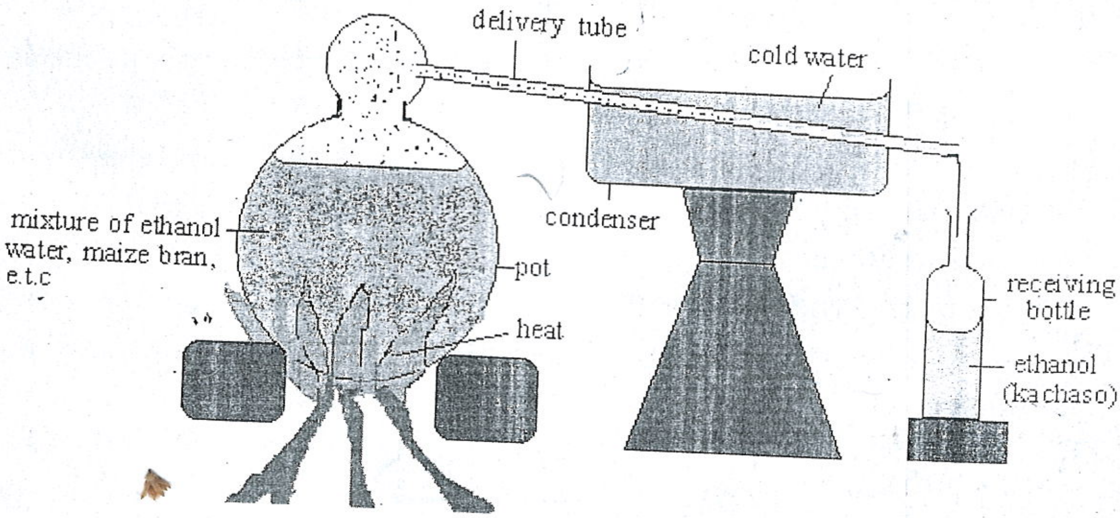
- Ethanol, C_2H_5OH, is the second member of the alkanol series.
- It is a clear colourless liquid with a pleasant smell.
- Ethanol is the most commonly used alkanol in the world over.
- Ethanol is very soluble in water.
- It burns readily in the presence of oxygen with a clean hot flame.

Manufacture of ethanol

There are two major ways of manufacturing ethanol. These ways are discussed below;

1. **Indigenous ways of preparing ethanol**

- Maize bran (madeya) with sugar or sugarcane or juices of fruits such as mangoes, *masuku*.
- The mixture is left for a few days (three to five days) for the fermentation to take place
- When the fermentation process stops, ethanol is then distilled from the mixture as shown below.



- When the mixture is heated, ethanol boils faster than the water because of its boiling point is lower.
- The gaseous ethanol rises up in the pot and passes through the delivery tube.
- Cold water condenses ethanol gas to liquid which is collected in the receiving bottle.

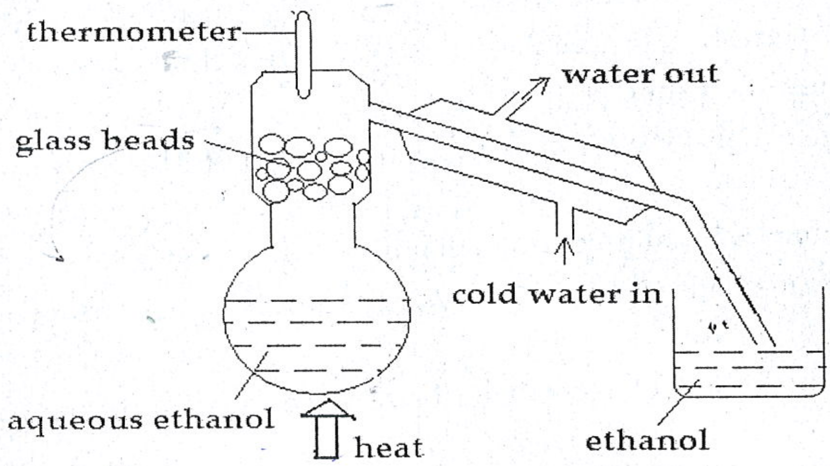
2. **Preparation of ethanol through fermentation of sugar by yeast**

- Fermentation is a process in which sugars (i.e. glucose) are broken down (decomposed) in presence of enzymes (yeast) to form ethanol and carbon dioxide.

- Yeast is added to sugar solution and the mixture is kept at room temperature.
- The reaction produces Ethanol and CO_2 as shown below

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2.$$

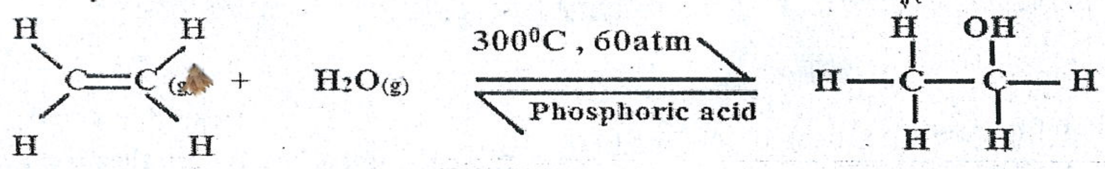
- Ethanol produced in this reaction contains some water. Pure ethanol is obtained from the mixture by fractional distillation as shown below



- When heat is supplied to the mixture, the liquid evaporate and rises into the condenser.
- The temperature in the condenser is regulated to about 78ºc (boiling point of Ethanol). This makes all the other substances with their boiling points above 78ºc to condense and fall back into the flask.
- Only ethanol vapours passes on to the delivery tube where it condenses and is collected as liquid ethanol.

3. **Hydration of Ethene**

- Ethane is reacted with water in the presence of a catalyst such as of phosphoric (V) acid (absorbed on silica pellets) at a temperature of 300°C and a pressure of 60 atmospheres.



Ethene

Ethanol

- The reaction is exothermic and reversible.
- the conditions have been chosen to ensure the highest possible yield of ethanol.

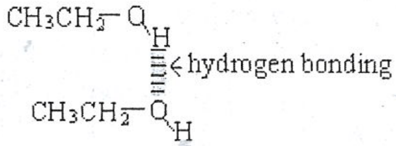
Physical Properties of alkanols

1. **Polarity**

The presence on the O–H bond makes alkanols polar.

2. **State at room temperature**

- The presence the hydroxyl group (-OH) in the structure of Alkanols raises the forces of attraction between the molecules of Alkanols.



- This makes small Alkanols to be liquids at room temperature. Large Alkanols are solids at room temperature.

3. **Solubility**

- The law of solubility states that like dissolve like. Alkanols are soluble in water because of the presence of -OH.
- However, their solubility decreases moving down the series. This is the case because the relative size of –OH decreases moving down the series.

4. **Melting points and boiling points**

- Alkanols have higher melting and boiling points as compared to hydrocarbons (alkanes and alkenes).
- Moving down the series, the molecular size of alkanols increase. As a result, the intermolecular forces increase as well. This leads to the increase in boiling and melting points of alkanols.

5. **Densities**

- The density of alkanols increases moving down the series.

6. **Volatility**

- Alkanols are volatile (easily vaporises). Volatility decreases down the series due to the increase in strength of the intermolecular forces.

7. **Viscosity**

- Alkanols have high viscosity than water and hydrocarbons.
- As we Move down the series, the molecular size of alkanols increases. As a result, the intermolecular forces increase as well. This leads to the increase in the viscosity of alkanols.

8. **Conductivity**

- Alkanols do not ionise in water as such do not conduct an electric current.

Chemical Properties of alkanols

1. **Reaction of alkanols with alkali metals.**

Alkanols react with alkali metals to produce organic compounds known as alkoxide.

For example, ethanol reacts with sodium to produce sodium ethoxide as represented by the following equation.

$2CH_3CH_2OH + 2Na \rightarrow 2CH_3CH_2ONa + H_2$.

Experiment;

Aim: Comparing reaction rates of ethanol with sodium and water with sodium.

Material: Ethanol, Water, sodium metal, phenolphthalein indicator, 2 test tubes.

Procedure

a. Pour 3 ml of ethanol into a test tube.
b. Add 3 drops of phenolphthalein indicator into the test tube.
c. Put a small piece of freshly cut sodium metal into the test tube.
d. Record observations in the table of results below.
e. Repeat steps (i) and (iv) with water.

Test liquid	Observation
water	
Ethanol	

Observations

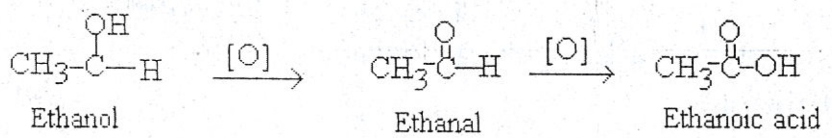
- In both reactions the solution turned pink indicating presence of a basic solution.
- The reaction of sodium with water was more violent than the reaction between sodium and ethanol.

Conclusion

Sodium reacts with ethanol the same way it reacts with water; reaction of sodium with ethanol is more gentle than the reaction between sodium and water.

2. **Oxidation of alkanols to alkanals and alkanones and alkanoic acids**

- A primary alkanol can be oxidized to alkanal or carboxylic depending on the experiment conditions.
- The oxidizing agent used in these reactions is usually a solution of sodium or potasssium dichromate(VI) acidified with dilute sulphuric acid.
- Secondary alkanols can be oxidized to corresponding alkanones.
- For example, ethanol can be oxidized first to ethanal and then to carboxylic acid as follows;



- The oxidation of ethanol to ethanal is by losing a hydrogen atom, H, while the oxidation of ethanol to ethanoic acid is by gaining an oxygen atom, O.

3. **Dehydration of ethanol to ethene.**

Ethanol is dehydrated to ethene when it is heated at about 85°C with sulphuric acid or phosphoric acid as the following reaction equation shows;

$$CH_3CH_2OH \rightleftharpoons CH_2{=}CH_2 + H_2O.$$

4. **Combustion**

- Ethanol for instance burns with a pale blue frame and products of complete combustion are carbon dioxide and water.

  $2CH_3CH_2OH_{(l)} + 6O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(l)}$.

Testing For Alkanols

1. **Solubility Test**

- Fill a test tube with 2cm³ of water and Add 5 drops of the test liquid.
- If only one layer forms, the test liquid may be a small alkanol.

2. **Sodium Test**

- Put 10 drops of the test liquid into a test tube
- Drop a small piece of sodium metal into the test tube.
- Add two drops of phenolphthalein.

Results and conclusion

- If phenolphthalein indicator turns pink, then the test liquid is an alkanol. If phenolphthalein does not turn pink/scarlet, then the test liquid is not an alkanol.
- Bubbles of a gas are formed that give a 'pop' sound when ignited (brought near a flame). The gas is hydrogen.

Uses of ethanol

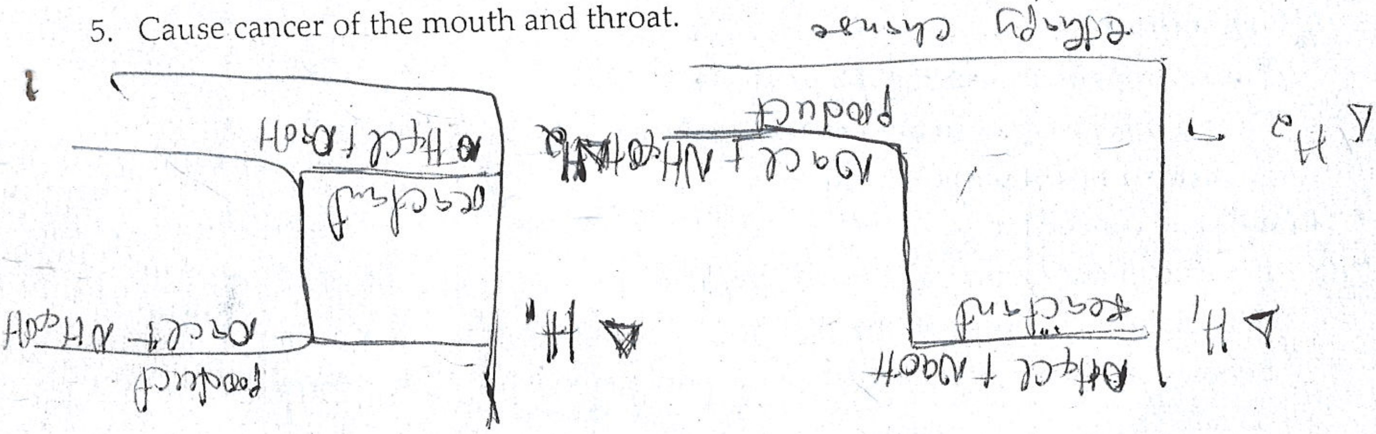
as a solvent, as constituent of alcoholic beverages and medicine, as an antiseptic, as a raw material in industry and as fuel for vehicles

1. Methylated spirit is ethanol with small amount of poisonous substances added to stop people from drinking it.
2. Used in food flavourings.
3. Very good solvent and used extensively as a solvent for paints, glue, aftershave and many other everyday products.
4. Ethanol is also used as a source of energy i.e. for heating, cooking and lighting.
5. It is also used as a raw material for making other substances, such as synthetic rubbers and flavourings
6. As a constituent of alcoholic drinks. Ethanol is present in alcoholic drinks such as Malawi gin.

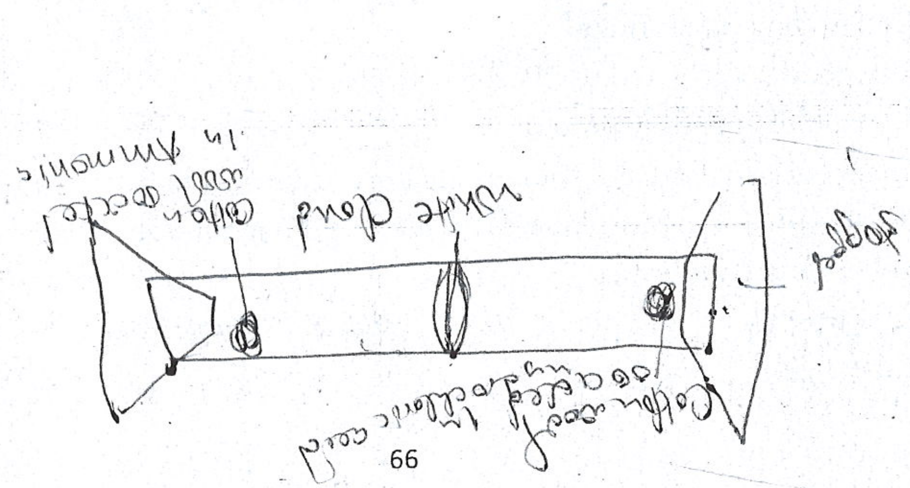
7. As a raw material in industries. Ethanol is used as raw material in the production of vinegar and yeast.
8. As medicine. Ethanol is used in the processing of antibiotics, vaccines, tablets, pills and vitamins.
9. As Antiseptic (a substance that destroys microorganisms that carry diseases). Many cleaning products contain high volumes of industrial alcohol.

Effects of excessive drinking of alcohols on human health

1. Causes accidents such as car crush, burns and drowning which may lead to injuries and loss of life
2. Increases job injuries and loss of productivity
3. Suicide, sexual assault and domestic violence
4. Cause high blood pressure, stroke, and other heart related problems.
5. Cause cancer of the mouth and throat.



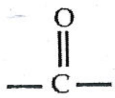
$NH_4Cl + NaOH \rightarrow NaCl + NH_4OH$



Topic 7: Alkanals and Alkanones

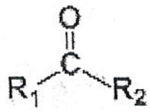
Functional group of alkanals and alkanones

- Alkanals and alkanones are organic compounds with functional group C=O called the carbonyl group,.

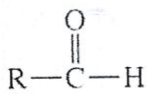


Structural difference between Alkanals and alkanones

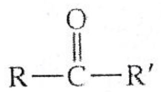
- They have the general formular as shown below where R is a hydrogen, H, or an alkyl group such as CH_3- or CH_3CH_2-



- Alkanals are also called aldehydes while alkanones are also called ketones.
- **Alkanals** have at least one hydrogen attached to the **carbonyl group**.
- Alkanones have two carbons attached to the **carbonyl groupIn** alkanals, ie both R_1 and R_2 are alkyl groups.

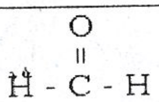
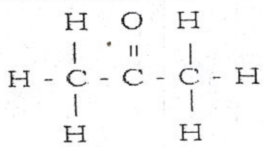
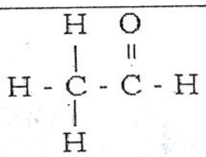
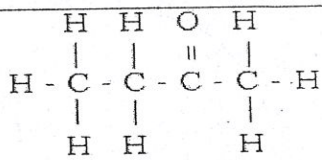
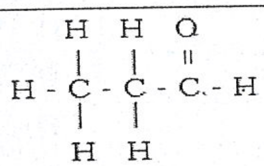
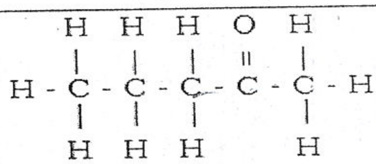


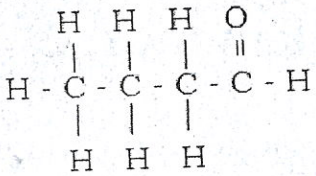
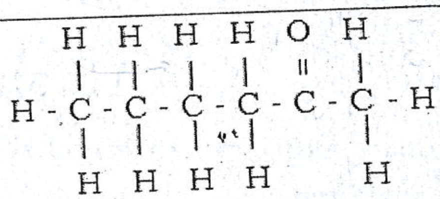
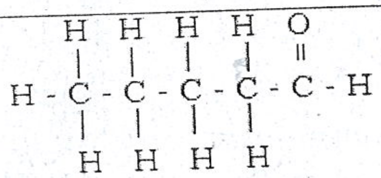
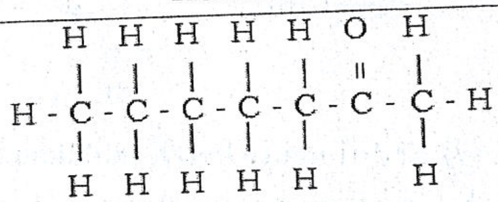
alkanal



alkanone

- Ketones do not have a hydrogen atom attached to the carbonyl group.

Alkanal structure	Alkanone structure
 Methanal	 Propanone
 Ethanol	 Butanone
 Propanal	 Pentanone

 Butanal	 Hexanone
 Pentanal	 Heptanone

Naming Alkanals and alkanones

- Replace –e with –al in the hydrocarbon to name the corresponding alkanal and –e with –one for alkanones.
- Below are names of first five alkanals and alkanones

Alkanals	Alkanones
Methanal	Propanone
Ethanal	Butanone
Butanal	Pentanone
~~Butanal~~ pentanal	Hexanone
Pentanal hexanone	Heptanone

$C_nH_{2n+1}COH$

$C_nH_{2n}O$

C_nH_{2n}

C_nH_{2n+2}

Molecular formulae for the first five alkanals and alkanones

1. **Alkanals**

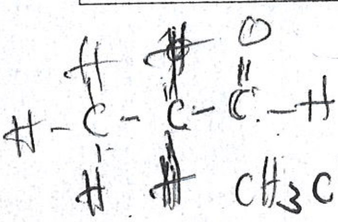
- The general formula of alkanals is $CnH_{2n+1}COH$ where n = 0, 1, 2, 3, ………
- This formula can be used to work the formula of any alkanal given the value of n.

Example:

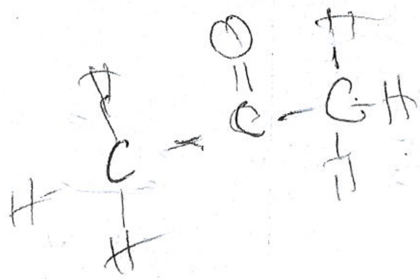
1. When n = 0 we have $C_0H_{2(0)+1}COH \rightarrow HCOH$
2. When n = 3 we have $C_3H_{2(3)+1}CHO \rightarrow C_3H_7COH$

- Table below shows the molecular and condensed formula for the first 5 alkanals.

n	Name	Molecular formular	Condensed formula
1	Methanal	$HCOH$	$HCOH$
2	Ethanal	CH_3COH	CH_3COH
3	Propanal	C_2H_5COH	CH_3CH_2COH
4	Butanal	C_3H_7COH	$CH_3CH_2CH_2COH$
5	Pentanal	C_4H_9COH	$CH_3CH_2CH_2CH_2COH$



CH_3CH_2COH



2. **Alkanones**

- The general formula of alkanones is $C_nH_{2n}O$.
  When n = 3 we get C_3H_6O
  Using the general structure this is usually written as CH_3COCH_3
- The table below shows the molecular and condensed formulae of the first five alkanones.

n	Name	Molecular formula	Condensed formula
3	Propanone	CH_3COCH_3	CH_3COCH_3
4	Butanone	$C_2H_5COCH_3$	$CH_3CH_2COCH_3$
5	Pentanone	$C_3H_7COCH_3$	$CH_3CH_2CH_2COCH_3$
6	Hexanone	$C_4H_9COCH_3$	$CH_3CH_2CH_2CH_2COCH_3$
7	Heptanone	$C_5H_{11}COCH_3$	$CH_3CH_2CH_2CH_2\ CH_2COCH_3$

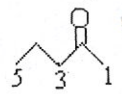
IUPAC Naming

- Includes the position of functional group

Rules for naming straight chain alkanals and alkanones

- Find longest chain that bears carbonyl group C=O.
- Number carbon chain so carbon bearing C=O has lowest possible number, it cannot be #1
- Name the corresponding alkane.
- Modify the name of the alkane by removing the end –e and replace it with –one or al.
- For alkanones, If there are more than 4 carbon atom in the longest carbon chain of the alkanone, a number which indicates the carbon atom to which the oxygen atom, O, is bonded is placed before the suffix -one.

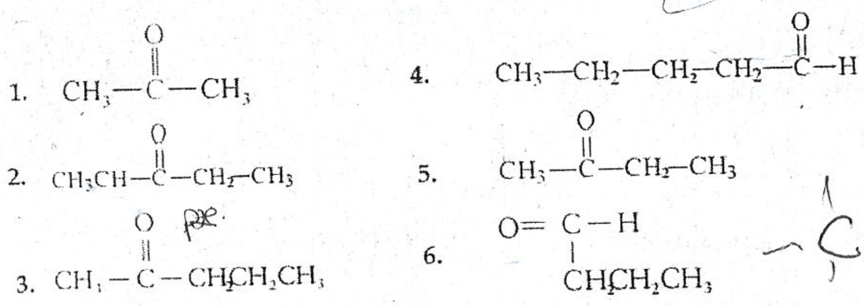
Example

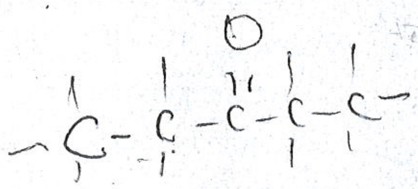


- Number of carbon atoms is 5
- Corresponding alkane is pentane
- Replacing –e with –one, we have pentanone.
- Carbon number to which O is attached is 2
- The full name is pentan -2-one.

Name the following

C_nH_{2n+1}
$C_4H_{2(4)+1}$
C_4H_9



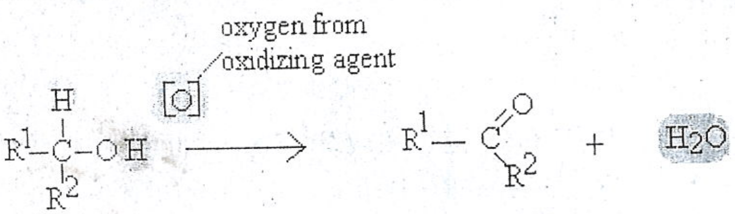


Draw the structure of the following

1. Heptan-2-one 2. Hexan-2-one 3. Heptan-4-one

Sources of alkanals and alkanones

- They are both produced from the oxidation alkanols
- Primary alkanols can be oxidized to corresponding alkanals
- Secondary alkanols can be oxidized to corresponding alkanones.
- The oxidizing agent used in these reactions is usually a solution of sodium or potasssium dichromate(VI) acidified with dilute sulphuric acid.
- The net effect is that an oxygen atom removes a hydrogen atom from the –OH group and another hydrogen from the carbon to which the –OH group is attached as follows;



- If either or both of R¹ and R² are hydrogen, the product is an aldehyde as shown in the eqaution below;

$CH_3CH_2CH_2OH + [O] \rightarrow CH_3CH_2CHO + H_2O$

Propanol **Propanal**

- If both R¹ and R² are alkyl groups the product is ans alkanone as follows;

$CH_3\underset{}{\overset{OH}{C}}HCH_3 + [O] \longrightarrow CH_3\overset{O}{C}CH_3 + H_2O$

Reactions of alkanals and alkanones

1. Adehydes React with oxygen (from an oxidising agent) to produce corresponding alkanoic acids

$2CH_3COH$ + [O] → $2CH_3COOH$

Ethanal + oxygen → ethanoic acid

- Alkanones are not readily oxidizable because they do not contain an oxidizable hydrogen atom bonded to the carbonyl group.

2. In an Addition reaction alkanals and alkanones react with hydrogen to form alcohols.

$$\underset{\text{aldehyde}}{R-\overset{\overset{\large O}{\|}}{C}-H} + H \xrightarrow{Pt} \underset{\text{primary alcohol}}{R-\underset{\underset{H}{|}}{\overset{\overset{OH}{|}}{C}}-H} \qquad \underset{\text{ketone}}{R-\overset{\overset{\large O}{\|}}{C}-R'} + H \xrightarrow{Pt} \underset{\text{secondary alcohol}}{R-\underset{\underset{H}{|}}{\overset{\overset{OH}{|}}{C}}-R'}$$

Examples

$$\underset{\text{propanal}}{CH_3CH_2-\overset{\overset{\large O}{\|}}{C}-H} + H_2 \xrightarrow{Pt} \underset{\text{1-propanol}}{CH_3CH_2\overset{\overset{OH}{|}}{CH_2}}$$

$$\underset{\text{acetone}}{CH_3-\overset{\overset{\large O}{\|}}{C}-CH_3} + H_2 \xrightarrow{Pt} \underset{\text{2-propanol}}{CH_3-\overset{\overset{OH}{|}}{CH}-CH_3}$$

Physical properties of alkanones and alkanals

1. All alkanals and alkanones are liquids except methanal which is a gas.
2. All are colourless except benzaaldehyde which is pale yellow.
3. Aldehydes and ketones are polar compounds as the carbonyl group is polar.
   - Cannot form intermolecular hydrogen bond but can hydrogen bond to water and ethanol.
4. Have higher melting points and high boiling points than the corresponding alkanes but have lower melting points and boiling points than the corresponding alkanols due to Lack of hydrogen..

Trends in melting points and boliling points and solubility

1. Both melting points and boliling points of ~~alakanals~~ alkanals and alkanones increase as the size of the molecules get bigger which ~~inreases~~ increase the size of the intermolecular forces .
2. The small alkanals and alkanones dissolve freely in water because they are able to hydrogen bond with water molecules.but as size of the molecules increases the solubility decreases.

Test for Alkanals and Alkanones

- Alkanals can be distinguished from alkanones using oxidation reaction using Fehling's, Brady's and Tollen's (silver mrror test).

A. The Brady's test

Materials: Two test tubes, Brady's reagent, Subsances to test.

Procedure

1. Add a few drops of Brady's reagent into each of the two test tubes.
2. Add a few drops of the suspected carbonyl compound

Results

If a yellow-orange precipitate forms, then the substance is an alkanal or alkanone. Note that Brady's test only tells whether a substance is an alkanal or alkanone, but it does not distinguish the two.

B. The Tollen's test

Materials: A test tube, Freshly prepared Tollen's reagent, Test substance.

Procedure

1. Pour some of the freshly prepared Tollen's reagent into the test tube
2. Add a few drops of the suspected substance to the Tollen's reagent in the test tube.
3. Warm the mixture gently in a hot water bath for a few minutes.

Results

If the solution turns silver mirror, then it is test substance is an alkanal otherwise it is an alkanone.

C. The Fehling's test

Materials: A test tube, Fehling's reagent, test Substance.

Procedure

1. Pour some of the Fehling's reagent (which is blue in colour) into the ~~the~~ test tube
2. Add a few drops of the suspected substance to the Fehling's reagent in the test tube.
3. heat ~~Warm~~ the mixture gently in a hot water bath for a few minutes.

Results

No change in the colour of the blue solution means that the substance added is an alkananone.

If the blue solution produces a brown or a brick-red precipitate, then the substance added is an alkanal.

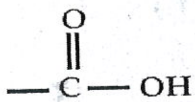
Uses of alkanals and alkanones

1. Synthesing other organic compounds. Eg ethanal is used to make ethanoic acid.
2. Making plastics. Methanal is used to make Bakelite while propanone is used to make Perspex.
3. As solvents. Propanone is used as a solvent in both laboratory and industry.

4. As preservatives. For example, formaline which contains 40% methanal aqueous solution is used to preserve animal specimens
5. As food flavours. Benzaldehyde, which has a characteristic ordour is used to flavour foods.
6. Used in many industries as
    - Food chemicals
        - Natural food additives
        - Artificial additives
    - Fragrance chemicals
    - Medicines
    - Agricultural chemicals

Topic 8: Alkanoic acids

- These are also called Carboxylic **acids**
- have the functional group called carboxyl group: **-COOH** which has the following structure



- The carboxyl is looked up as a combination of *carbonyl group*, **C = O** and *hydroxyl group*, -OH

General formula

- Alkanoic acids have the general formula **$C_nH_{2n+1}COOH$**, where **n** is number of carbon atoms **minus one.**

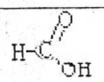
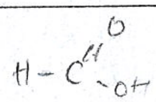
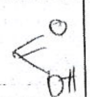
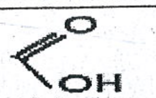
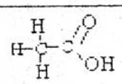
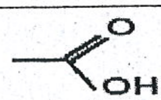
  For instance if n = 4 then **$C_nH_{2n+1}COOH = C_3H_{2(3)+1}COOH = C_3H_7COOH$**

Nomenclature

- A suffix – anoic acid is added to the general prefix indicating the number of carbon atoms in the molecule including the carbon atom in the – COOH group.

No of C atoms	Molecular formula	Prefix	Surfix	Name
1	HCOOH	Meth-	-anoic acid	Methanoic acid
2	CH_3COOH	Eth-	-anoic acid	Ethanoic acid
3	C_2H_5COOH	Prop-	-anoic acid	Propanoic acid
4	C_3H_7COOH	But-	-anoic acid	Butanoic acid
5	C_4H_9COOH	Pent-	-anoic acid	Pentanoic acid
6	$C_5H_{11}COOH$	Hex-	-anoic acid	Hexanoic acid
7	$C_6H_{13}COOH$	Hept-	-anoic acid	Heptanoic acid
8	$C_7H_{15}COOH$	Oct-	-anoic acid	Octanoic acid
9	$C_8H_{17}COOH$	Non-	-anoic acid	Nonanoic acid
10	$C_9H_{19}COOH$	Dec-	-anoic acid	Decanoic acid

Structural, molecular, condensed and skeletal formulae

Name	Structure formula	Condense	Skeletal
Methanoic acid	 	HCOOH 	
Ethanoic acid		CH_3COOH	

C_nH_{2n+1}

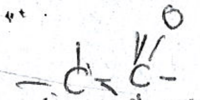
$C_9H_{2(9)+1}$

C_9H_{19}

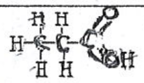
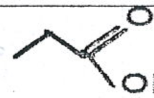
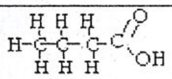
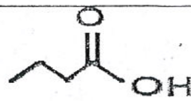
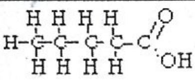
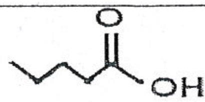
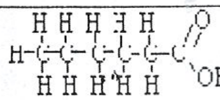
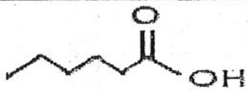
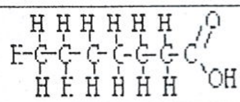
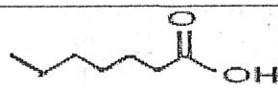
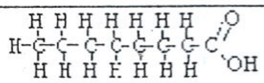
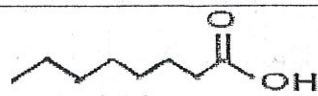
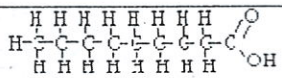
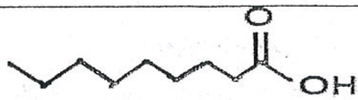
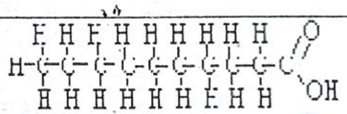
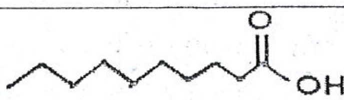
C_2H_{2n+1}..

$C_2H_{2(2)+1}$

C_2H_5



(CH_3COOH)

Propanoic acid		$CH_3CH_2\ COOH$	
Butanoic acid		$CH_3CH_2\ CH_2COOH$	
Pentanoic acid		$CH_3(CH_2)_3COOH$	
Hexanoic acid		$CH_3(CH_2)_4COOH$	
Heptanoic acid		$CH_3(CH_2)_5COOH$	
Octanoic acid		$CH_3(CH_2)_6COOH$	
Nonanoic acid		$CH_3(CH_2)_7COOH$	
Decanoic acid		$CH_3(CH_2)_8COOH$	

Deducing the molecular Formula of a carboxylic acid given its Percentage composition:

- The molecular formula of a carboxylic acid can be worked out using its percentage composition and the general formula of the carboxylic acids

Example:

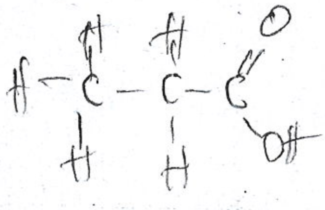
A certain carboxylic acid has the following percentages of elements by mass; carbon = 48.7 %, hydrogen = 8.1% and oxygen = 43.2%. Work out the molecular formula of the carboxylic acid. What is its name?

Solution

	Carbon	**Hydrogen**	**Oxygen**
Mass compositions	48.7g	8.1g	43.2g
Number of moles	$\frac{48.7g \times 1mole}{12g}$	$\frac{8.1g \times 1mole}{1\ g}$	$\frac{43.2g \times 1mole}{16g}$
	= 4.0 moles	= 8.1 moles	= 2.7 moles
Divide by the lowest result;	$\frac{4.0\ moles}{2.7\ moles} = 1.5$	$\frac{8.1moles}{2.7\ moles} = 3$	$\frac{2.7\ moles}{2.7\ moles} = 1$

Multiply each above result by 2 to make 1.5 a whole number;

	1.5 x 2 = 3	3 x 2 = 6	1 x 2 = 2



Therefore the carboxylic acid has 3 carbon atoms, 6 hydrogen and 2 oxygen atoms.

i . From the general formula, $C_nH_{2n+1}COOH$, then the molecular formula is C_2H_5COOH

ii. Its name is Propanoic acid.

Exercise:

Work out the molecular formula and the structure of a carboxylic acid which has the following percentage compositions by mass; 62.1% carbon, 10.3% hydrogen and 27.6% oxygen.

$C_3H_{2(3)+1}$

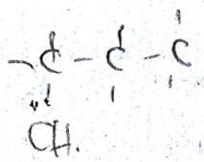
C_3H_5COOH

C_3H_6

Natural sources of alkanoic acids

Name of acid	source
Formic acid	the 'sting' of ants
Ethanoic acid	vinegar
Butanoic acid	butter
Citric acid	in citrus fruits like lemons and oranges
Tartaric acid	Grapes
Malic acid	apples
Lactic acid	in sour milk
Ascorbic acid (vitamin C)	from foods
Hexanoic acid	Goat milk

C_3H_8



Preparation of alkanoic acids

- Alkanoic acids are prepared from oxidation of alkanols and alkanals.
- An alkanol is oxidised to a corresponding alkanal which in turn is oxidised to a corresponding carboxylic acid, eg ethanol is oxidised as follows

i. Reaction with oxygen to produce ethanal

$$CH_3CH_2OH + [O] \rightarrow CH_3CHO + H_2O$$

Ethanol → ethanal

ii. Ethanal Reacts with oxygen (from an oxidising agent) to produce corresponding alkanoic acids

$$2CH_3COH + [O] \rightarrow 2CH_3COOH$$

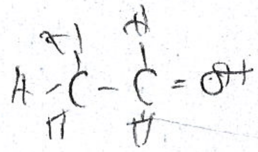
Ethanal + oxygen → ethanoic acid

Physical properties of alkanoic acids

1. **State at room temperature**

- Liquids at room temperature (i.e. methanol to Nonanoic acid)

2. **Solubility**



CH_3CH_2OH

- Small molecule carboxylic acids are soluble in water. Large molecules are less soluble because of the size factor.
- Both carboxylic acids and water contain –OH groups, and therefore acids dissolve in water. "Like dissolves like."
- But as the size of the molecule increases the effect of the – OH group decreases and this makes the bigger carboxylic acids to be insoluble.

3. **Visicosity**

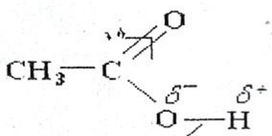
- Carboxylic acids can show high viscosity (stickiness) as liquids.
- Both viscosity and density increase on going along the series due to the increasing size of the IMF as size of molecules increases.

4. **Melting points and boiling points**

- Increases down the homologous series due to the increasing size of the IMF as size of molecules increases.
- They have higher melting and boiling points than those of the corresponding alkanols because

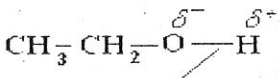
i. Carboxylic acids form stronger hydrogen bonds than the alkanols. This is because the –OH group of the carboxylic acids is more polarised due to the presence of the electron-withdrawing -C=O group as shown below;

ethanoic acid molecule



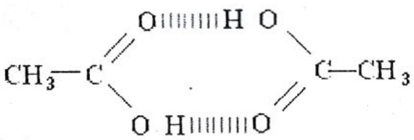
more polarised bond

ethanol molecule



less polarised bond

ii. Carboxylic acids exist as dimmers i.e. two molecules bonded together as follows;



5. **Electric conductivity**

- Alkanoic ionise in water as such they conduct electric current

$CH_3COOH_{(aq)} \rightarrow CH_3COO^-{}_{(aq)} + H^+$

6. **Changing colour of indicator**

Name of indicator	Colour in acid solution
Blue litmus	Red
Methyl orange	Pink
Methyl red	Red
Phenolphthalein	Colourless

Red litmus	Red
Universal indicator	Red

Reactions of alkanoic acids

1. Reaction with metals to produce salts and hydrogen

   $CH_3COOH_{(aq)} + Mg_{(S)} \rightarrow (CH_3COOMg)_{2(aq)} + H_{2(g)}$

   Magnesium ethanoate

   $CH_3COOH_{(aq)} + K_{(S)} \rightarrow (CH_3COOK_{(aq)} + H_{2(g)}$

   Potassium ethanoate

2. They react with bases to form salts, for example;

   $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$.

   ethanoic acid **sodium hydroxide** **sodium ethanoate** **water.**

3. Reaction with carbonates to produce salt, carbon dioxide and water

   $CH_3COOH + NaHCO_{3(aq)} \rightarrow CH_3COONa + CO_2 + H_2O$.

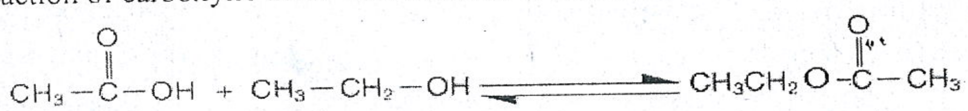
   Ethanoic acid **sodium hydrogen carbonate** **sodium ethanoate**

4. Reaction with oxygen to form carbon dioxide and water

   $CH_3COOH + 2O_{2(g)} \rightarrow 2\ CO_{2(g)} + 2H_2O_{(l)}$.

5. Esterification reaction

   reaction of carboxylic acids with alcohols to form **Esters**



ethanoic acid ethanol ethylethanoate (ethyl acetate)

Test for alkanoic acids

- We use the Acid test:
- It is based on the ability of carboxylic acids to change the colour of an indicator during neutralisation reaction.

Procedure:

- Put 15 drops of sodium hydroxide (NaOH) solution in a test tube
- Add 2 drops of phenolphthalein indicator. This gives a pink colour (phenolphthalein in basic solution).
- To this is added 6 drops of the test liquid.

Results:

- No change in colour means that the added liquid is not an acid. A change to colourless shows the test liquid contains the carboxylate group, -COOH and therefore an acid.

Uses of alkanoic acids

1. Solvents for most organic compounds
2. Food preservatives, the acidity prevents bacteria from being active and inhibits their growth
3. Production of house hold vinegar (ethanoic acid)
4. Production of esters
5. Removing rust from base metal (citric and oxalic acid)
6. Reagent in the production of polyvinylacetate, cellulose acetate, synthetic fibre and fabrics.

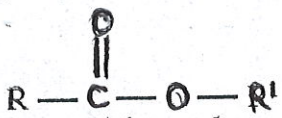
1:30
2:40

Topic 9: Alkanoates (also called Esters)

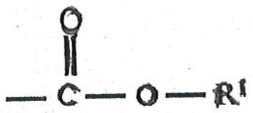
- These are sweet smelling compounds formed from the reaction between an acid and an alkanol.

  Alkanol + Alkanoic acid → Alkanoate (ester) + water

- This type of reaction is called **Esterification.**
- The alkanoates have the general formula and functional group below where R is an alkyl group



general formula



functional group

Nomenclature of alkanoates

- Alkanoates have two parts; the acid part and the alkanol part joined together by an ester link.

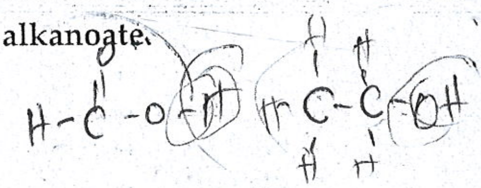
acid part alkanol part

- The names of esters also have two parts, the stem of an alkanol with surfix **–yl** which comes first followed by a stem of the name of the acid with the sulfix **–oate.**
- From **ethanol** and **ethanoic acid** the name would be **ethylethanoate.**

Rules for naming alkanoates

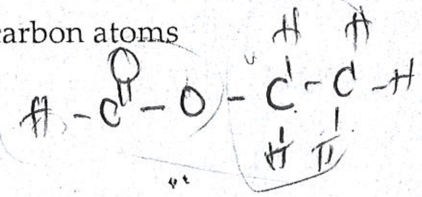
1. Divide the formula into two parts using a line
   R_1COOR_2 → R_1COO / R_2
2. Name the last part R_2 first an alkyl group from an alkanol
3. Name the first part R_1COO as an alkanoic acid but substitute **ic** with **oate.**
4. Combine 2 and 3 to get a full name of the ester **Alkyl alkanoate.**

   Example



What is the name of $CH_3COOC_2H_5$.

- From rule 1 we have CH_3COO / C_2H_5.
- The last part C_2H_5 is called ethyl since it has 2 carbon atoms
- The first part is ethanoate 2C as well



- Full name **ethylethanoate.**

Exercise

Name the following alkanoates

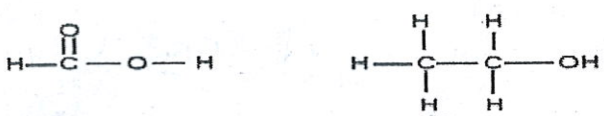
1. $COOC_2H_5$
2. $C_2H_5COOC_2H_5$
3. $C_2H_5COOCH_3$
4. $CH_3COOC_5H_{11}$
5. $C_2H_5COOC_3H_7$
6. $CH_3COOC_4H_9$
7. $C_2H_5COOC_4H_9$
8. $C_4H_9COOC_2H_5$
9. $C_7H_{15}COOC_6H_{13}$
10. $HCOOCH_3$

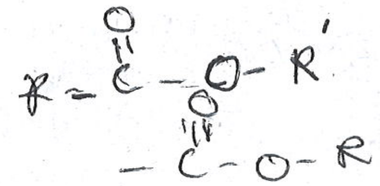
Drawing structures of alkanoates from a given alkanoic acid and alkanol

Example

Draw the structure of an alkanoate formed from ethanol and methanoic acid

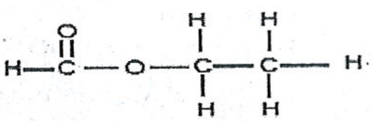
- Draw the structure of ethanol and methanoic acid





- Take out H from the acid and –OH from the alkanol

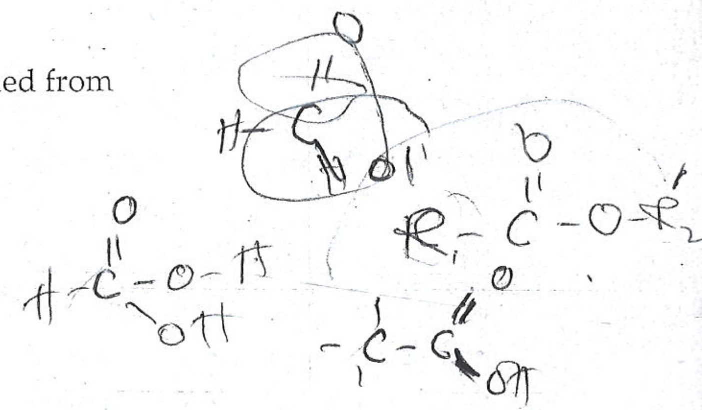
- Join the alkanoate from methanoic acid to the ethyl group from ethanol



Ethylmethanoate

Exercise

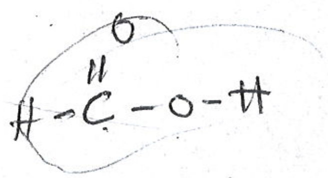
Draw the structure and name an alkanoate formed from



1. Methanol and ethanoic acid
2. Pentanol and methanoic acid
3. heptanol and propanoic acid
4. propanol and butanoic acid
5. hexanol and ethanoic acid
6. ethanol and nonanoic acid

Deducing the reactants of esterification

- We can deduce the structure and names of an alkanol and an alkanoic acid that were used to make a particular ester



Example

Deduce the structure of reactants that produce $CH_3CH_2COOCH_3$.

Solution

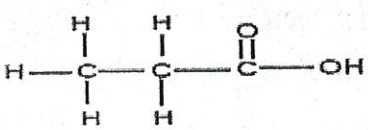
- Divide the alkanoate into the acid part and alkanol part

  CH_3CH_2COO $\qquad$ CH_3

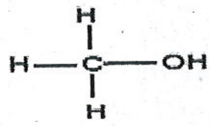
- Add H to the acid part and –OH to the alkanol part

  CH_3CH_2COOH $\qquad$ CH_3OH

- Draw the structures and name the acid and the alkanol



Propanoic acid



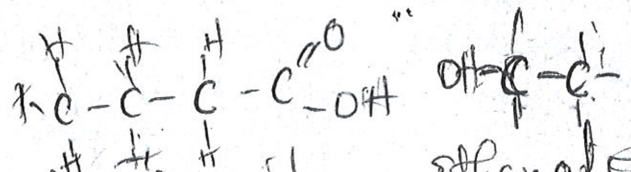
Methanol

Exercise

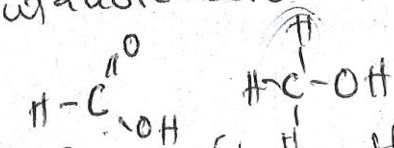
Draw the structures and name the acid and the alkanol that produce

1. $CH_3CH_2\ CH_2COO\ CH_2CH_3$.
2. $CH_3CH_2COO\ (CH_2)_3CH_3$
3. $HCOOCH_3$
4. $C_6CH_{13}COO\ CH_2CH_3$
5. $CH_3(CH_2)_8CH_2COO\ C_3CH_7$

$CH_3CH_2\ CH_2COOH\ \ CH_2CH_2OH$



butanoic acid $\qquad$ Ethanol



methanoic acid $\qquad$ methanol

propyl

Sources of alkanoates

1. **Natural sources**
   - Fruits and flowers
- Propylethanoate from pear
- Butylethanoate from banana
- Octylethanoate from orange
- Methylbutanoate from apple
- Ethylbutanoate from pineapple
   - Plants where they occur as fats and oils
   - Animals such as whales in which they occur in form of fats as well.

2. **Synthetic alkanoates**
- Man-made alkanoate by a process called esterification eg
- ✓ Polyester is a synthetic fibre used as a substitute for cotton and wool
- ✓ Ethylethanoate commonly used in flavouring for sweets and drinks

Formation of an ester (ethyl ethanoate)

Materials: 10ml measuring cylinder, test tube, ethanoic acid, sulphuric acid, ethanol, beaker 150ml, gas burner, distilled water

Procedure

a. Put 2ml of ethanol into a test tube
b. Add 3 drops of sulphuric acid to ethanol
c. Add 1ml of ethanoic acid to the mixture
d. Heat the mixture
e. Pour the heated mixture into a beaker containing 50ml of distilled water.
f. Smell the mixture in the beaker

Results and discussion

- The mixture produces a sweet fruity smell
- The product is an alkanoate, ethylethanoate according to the equation below;

$$CH_3COOH + C_2H_5OH \rightarrow CH_3COOC_2H_5 + H_2O$$

- Sulphuric acid is just a catalyst

Physical Properties Of Alkanoates

- Colourless volatile liquids
- Insoluble in water
- Have a good sweet smell and flavor
- Have low melting points and boiling points; no OH hence cannot hydrogen bond.

Chemical Properties Of Alkanoates

1. **Combustion**

- Burn in air with a bright flame to produce carbon dioxide and water

$$CH_3COOC_2H_5 + 5O_2 \rightarrow 4CO_2 + 4H_2O$$

2. **Hydrolysis**

- Reaction with water to produce an alkanol and a carboxylic acid

$$\underset{\text{Ethylethanoate}}{CH_3COOC_2H_5} + \underset{\text{Water}}{H_2O} \rightarrow \underset{\text{Ethanoic acid}}{CH_3COOH} + \underset{\text{Ethanol}}{C_2H_5OH}$$

- Reaction with an alkaline solution to produce an alkaline salt and alkanol

$$\underset{\text{Ethylethanoate}}{CH_3COOC_2H_5} + \underset{\text{sodium hydroxide}}{NaOH} \rightarrow \underset{\text{sodium ethanoate}}{CH_3COONa} + \underset{\text{ethanol}}{C_2H_5OH}$$

Uses of esters

1. Sorvents for adhesion eg ethylethanoate is a solvent for polyester
2. Flavouring sweets and drinks such as yoghurt and juices
3. Fragrance in soap, petroleum jellies and perfumes

4. Manufacture of clotheseg terelyne
5. Manufacture of cooking oil
6. Soap making: esters are reacted with sodium hydroxide to make soap.

Fats and oils

- Natural esters used as energy storage compounds in both plants and animals.
- They are grouped as glycerides. Fats are solid at room temperature, and of animal origin while oils are liquids and of plant origin.
- Made from alkanols and a large molecule of carboxylic acid by a condensation reaction
- For example glycerol (propane-1,2,3-triol) reacts with stearic acid (fatty acid) an alkanoate called glycerol tristearate as follows

$$\begin{matrix} CH_2OH \\ | \\ CHOH \\ | \\ CH_2OH \end{matrix} + \left(\begin{matrix} HOOC(CH_2)_{16}CH_3 \\ HOOC(CH_2)_{16}CH_3 \\ HOOC(CH_2)_{16}CH_3 \end{matrix}\right) \longrightarrow \begin{matrix} CH_2OOC(CH_2)_{16}CH_3 \\ | \\ CHOOC(CH_2)_{16}CH_3 \\ | \\ CH_2OOC(CH_2)_{16}CH_3 \end{matrix} + 3H_2O$$

Glycerol (alkanol) — stearic acid (fatty acid) — Tristearate (fat /alkanoate/ ester) — water

Uses of fats and oils

- Manufacture of soap by saponification.
- For making glycerol.
- For the production of margarine.
- Used in the manufacture of paints, candles

Soaps and detergents

- Detergents are substances used as cleansing agents eg soap. They are used to
  - ✓ Make water wetter
  - ✓ Help in the removal of dirt and grease

Manufacture of soap

- Raw materials are animal fats and vegetable oils for a process called saponification.
- In the process fats or oils are boiled with an alkaline solution usually sodium hydroxide to produce glycerol and soap (sodium salt of the acid)

$$\begin{matrix} CH_2OOC(CH_2)_{16}CH_3 \\ | \\ CHOOC(CH_2)_{16}CH_3 \\ | \\ CH_2OOC(CH_2)_{16}CH_3 \end{matrix} + NaOH \longrightarrow 3CH_3(CH_2)_{16}COONa + \begin{matrix} CH_2OH \\ | \\ CHOH \\ | \\ CH_2OH \end{matrix}$$

Tristearate (fat /alkanoate/ ester) — Sodium hydroxide — sodium stearate (soap) — Glycerol (alkanol)

- Sodium chloride is then added to the mixture which solidifies the soap which floats on top.

8

14

- Soap is then skimmed and processed where colours, perfumes and fillers are added.

Project: Soap Making

Materials: sodium hydroxide solution, cooking oil, sodium chloride, measuring cylinder, 200ml beaker, large spoon, funnel, filter paper, conical flask, stirring rod, gas burner.

Procedure

- Measure 2ml of cooking oil into a beaker
- Add 10ml of concentrated sodium hydroxide
- Heat the mixture gently while stirring
- Add 2 spoonful of sodium chloride to the mixture
- Add 10ml of distilled water and boil the mixture again
- Cool the mixture, then filter it to remove soap.

$\rightarrow 3CH_3(CH_2)_{16}COONa + CH_2OH$

$CHOH$

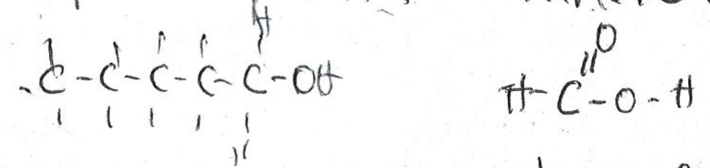
~~$NaOH + CH_2OOC$~~ CH_2OH

$NaOH + CH_2OOC(CH_2)_{16}CH_3$

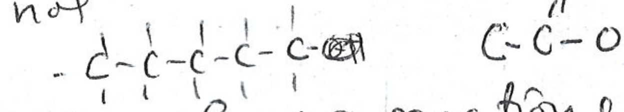
$CHOOC(CH_2)_{16}CH_3$

$CH_2OOC(CH_2)_{16}CH_3$

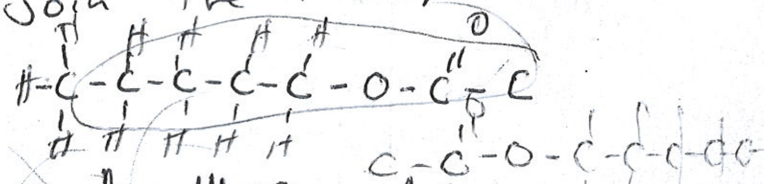
① pentanol and methanoic acid



pentakes out H from acid OH from alkanol

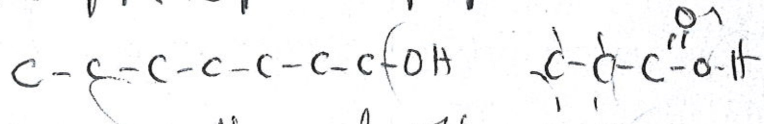


Join the two equations

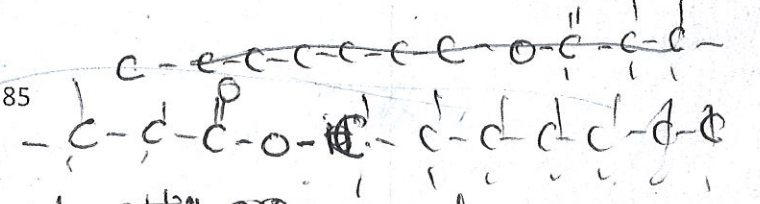


Pentyl-Methanoate

② heptanol and propanoic acid



remove H and OH



heptyl propanoate

Topic 10: Identification of unknown organic compounds

Deducing the family of an unknown organic compound

1. Functional Groups

- Recall that each organic family has its own functional group except for the alkanes.

Homologous series	**Functional group**	**Name of functional group**
Alkanes		Alkanes
Alkenes	C = C	Carbon – carbon double bond
Alkanols	–OH	Hydroxyl
Alkanals and Alkanones	–C = O	Carbonyl
Alkanoic acids	–COOH	Carboxyl
Alkanoates	–COO	Carboxylate

2. General Formula

- Each family has its own general formula which represents all the members of that family.

Homologous series	**General Formula**
Alkanes	C_nH_{2n+2}
Alkenes	C_nH_{2n}
Alkanols	$C_nH_{2n+1}OH$
Alkanals	$C_nH_{2n+1}COH$
Alkanoic acids	$C_nH_{2n+1}COOH$
Alkanones	$C_nH_{2n}O$ or RCOR
Alkanoates	RCOOR

NB: n is for number of carbon atoms

3. **Products of chemical reactions**

- Products of chemical reactions can also be used to identify organic compounds involved in the reaction.

a. Alkanes

- Generally unreactive because they have no functional group.

b. Alkenes

- React with halogens to produce haloalkanes.

~~CnHanta~~ C_nH_{2n} ~~RCOO~~ $C_nH_{2n}O$ RCOR

c. Alkanols

- React with sodium to produce a salt and hydrogen gas.

d. Alkanoic acids

- React with metals to produce a salt and hydrogen gas.
- Reacts with carbonates to produce salt, carbon dioxide and water.
- React with alkanols to produce esters and water.

4. Physical and Chemical properties

- Some unique physical and chemical properties can also be used to identify organic compounds.
- Table below shows example

Organic Family	Physical Properties	Chemical Properties
Alkanes	Insoluble in water	Substitution reaction
Alkenes	Insoluble in water	Decolourise bromine solution
Alkanols	Soluble in water	Dehydration to alkenes
Alkanals	Soluble in water	Reduce silver ions into silver metal
Alkanones	Soluble in water	Reduced to secondary alkanols
Alkanoic acids	Soluble in water and Electrolyte	Neutralise a base
Alkanoates	Small ones have sweet aroma Big ones are coloured	Hydrolysed to alkanoic acid and ~~and~~ alkanols

5. Tests

Family	Test	Results showing presence
Alkenes	Bromine test	Colour changes from brown to colourless
Alkanols	Sodium test	Vigorous reaction and H_2 is given out
Alkanals	Fehling's or Tollen	Green suspension and a red precipitate for Fehling's Silver plating in a container for Tollen
Alkanones	Braday's test	Bright orange or yellow precipitate. Same results for alkanals.
Alkanoic acids	Acid test	Blue litmus turns red
Alkanoates	Hydrolysis with NaOH followed by acid test	Vigorous reaction and the product turns blue litmus red.

6. Flow Diagrams

- Shows test and results for specific families in order to identify them. In addition to the tests shown above we also have the solubility test which in most instances is the first to be carried out.

  Example