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UNIT 1

EXPERIMENTAL TECHNIQUES

LABORATORY SAFETY AND SAFE WASTE DISPOSAL

Chemical waste: is a product or material from a chemical reaction or an expired product which is no longer needed in the next experiment.

CLASSES OF CHEMICAL WASTES

Chemical wastes can be put in two categories as follows:

- hazardous
- non hazardous

Hazardous waste: are chemical wastes which pose a health threat in our lives. These wastes can cause fire accidents such as explosions that can result in serious problems on our body.

Non-hazardous: are wastes which do not pose any health threat on our lives.

CLASSES OF HAZARDOUS CHEMICAL WASTES

Hazardous chemical wastes can be put in two categories as follows:

- Ignitable chemical wastes
- Reactive chemical wastes
- Toxic chemical wastes

Ignitable chemical wastes: these are chemical wastes which are flammable. Flammable means they can easily catch fire. They include organic solvents such as ethanol, methanol, butanol and others. They also include organic fuels such as hexane, acetone and others.

Reactive chemical wastes: these chemicals can easily form new substances when exposed to the environment. When lithium, potassium and calcium are mixed with water, they cause a chemical reaction which is highly explosive. These metals are therefore belonging to reactive chemical wastes.

Toxic chemical wastes: these wastes are poisonous. They can cause suffocation and death to human beings.

SAFE WAYS OF DISPOSING CHEMICAL WASTES

The following are some ways of disposing wastes:

1. Use of normal waste bins.

✎ This way can be used for wastes such as non-recycle plastics, pieces of wood, rubber and dirty papers.

2. Use of special controlled waste containers.

✎ These are used to dispose sharp objects like scalpel, broken glass ware, sample tubes and items contaminated with hazardous chemicals. These must be emptied regularly and not allowed to overflow.

3. Draining using a lot of water.

- Wastes which can be drained include harmless soluble inorganic salts, and contaminated and used acids and alkalis, detergents.

4. Recycling.

- Recyclable materials include unbroken glass, packing waste and paper.

5. Incineration.

- Wastes are put into an incinerator and burned. Wastes that can be incinerated include syringes, needles, all organic solvents, soluble organic wastes, paraffin and mineral oils.

SCIENTIFIC INVESTIGATIONS

- It is a process or method of trying to find an answer or explanation to a problem.
- This process involves a series of stages and it is important to understand each and every stage that is involved.

Planning scientific investigation

- Planning an investigation involves thinking about what you want to do so that you can make predictions more accurately.
- Some of the concepts involved in planning are:
 - Identifying factors affecting the investigation
 - Identifying the materials that will be used for the different activities
 - Describing how to make observations and record suitable data

STAGES OF SCIENTIFIC INVESTIGATION

1. Identifying a problem

- A problem that requires an answer is identified. It is considered to be a problem since its solution is not known. The final statement of the problem must be in a question form.

2. Formulating hypothesis or hypotheses

- A hypothesis is a guessed answer or solution to the problem. The hypotheses are made before the actual experimentation is done.
- Always come up with atleast two hypotheses: atleast one positive and the other negative, so that if one fails the opposite one may be true.
- The experimentation is done so as to confirm or reject the hypothesis.

3. Identifying or selecting the variables

- A variable is a factor that can affect the outcome of an investigation.
- There are two types of variable. The **manipulated variable** (independent variable) causes a change.
- The **responding variable** (dependent variable) changes in response to the manipulated variable.
- For example, if you want to find out the effect of temperature on the rate of reaction between hydrochloric acid and sodium chloride, you would suggest carrying out the

reaction at various temperatures to observe some effects.

- ☞ Temperature is therefore a variable that can be manipulated in this investigation. It is an independent variable, which you will change to observe the change in the rate of reaction.
- ☞ Variables are obtained from the statements of hypotheses or they can be identified from previous stages. At this stage, you will also identify suitable equipment or apparatus that will be used in the tests.

4. Controlling the variables

- ☞ This involves keeping all the variables constant except the one being tested. A statement of how each variable will be controlled is required at this stage.
- ☞ In the example above, to find out the effect of temperature on the rate of reaction of hydrochloric acid with sodium chloride you will need to hold other variable (such as volumes and concentrations) constant.

5. Testing the hypotheses

- ☞ This stage involves testing the hypotheses by doing experimentations. The purpose of testing the hypothesis is to collect data which should help in decision making.

6. Recording data (Data collection)

- ☞ This method involves making observations about what is happening during an investigation. It also involves taking measurements by using apparatus such as graduated instruments such as rulers, thermometers, measuring cylinders and beam balances. Data can be recorded in form of numbers, sentences, patterns and pictures of observation.

7. Data analysis(interpreting experimental data) and presentation

- ☞ Data is any qualitative or quantitative information that is used in determining the results of an experiment.
- ☞ Data analysis involves arranging the data so that in an orderly manner. The aim is to check the relationship between physical quantities. Data can be analysed in form of line graph, a bar graph and pie chart.
- ☞ Once data is analysed, it is presented in a way that enables you to see the relationship you are investigating.

8. Evaluating hypotheses

- ☞ This mainly involves making a decision on whether the hypotheses should be rejected or not rejected. You repeat the experiment (and have others repeat it) to confirm your results.

9. Drawing conclusions

- ☞ Conclusions are derived from the evaluation of the hypotheses. The conclusion must be in relation to the questions that you set out to answer. The conclusion must be supported with relevant reasons.

PURITY OF A SUBSTANCE

- A pure substance is a material that has constant composition and consistent properties throughout the sample.
- In other words, a pure substance does not contain impurities or any contaminants.
- A pure substance is a substance that is made of only one type of atoms or molecules. For example, pure water means that there is nothing in the water except for the water molecules.

Checking the purity of a substance

There are three basic tests carried out to check the purity of a substance. These are;

- ✓ Melting point
- ✓ Boiling point
- ✓ Chromatography

CHECKING PURITY OF A SUBSTANCE BY MELTING POINT

- ✎ A pure substance has a specific melting point. The presence of impurities affects the melting point in two ways:
 - Impurities lower the melting point of the substance. For example, pure ice melts completely at exactly 0°C but the ice we make at home from tap water will start to melt before 0°C.
 - Impurities cause the substance to melt over a range of temperatures.

CHECKING PURITY OF A SUBSTANCE BY BOILING POINT

- ✎ A pure substance has a fixed boiling point. The presence of impurities affects the boiling point in two ways:
 - Impurities raise the boiling point of a substance. The more impurities the substance contains, the higher its boiling point will be.
 - Impurities cause the substance to boil over a range of temperatures. So by looking at the range of boiling point of a substance, we can know whether the given substance is pure or not.

The table below shows some melting and boiling points of some common substances

<i>Substance</i>	<i>Melting point (°C)</i>	<i>Boiling point (°C)</i>
Water	0	100
Ethanol	-114	78
mercury	-15	357
ammonia	-78	-27
Sodium chloride	801	2250

CHECKING PURITY OF A SUBSTANCE BY PAPER CHROMATOGRAPHY

- ✎ Paper chromatography is used to separate mixtures of soluble substances. These are often coloured substances such as food colourings, inks, dyes or plant pigments.
- ✎ In chromatography, substances are picked up and carried by a mobile phase which moves through a stationary phase.

The stationary phase

✎ It is the part of the chromatography which is in solid state e.g. the filter paper.

The mobile phase

✎ It is the part of the chromatography which is in gaseous or liquid state e.g. the solvent.

✎ The different dissolved substances in the mixture are attracted to the two phases in different proportions. This causes them to travel at different rates through the paper.

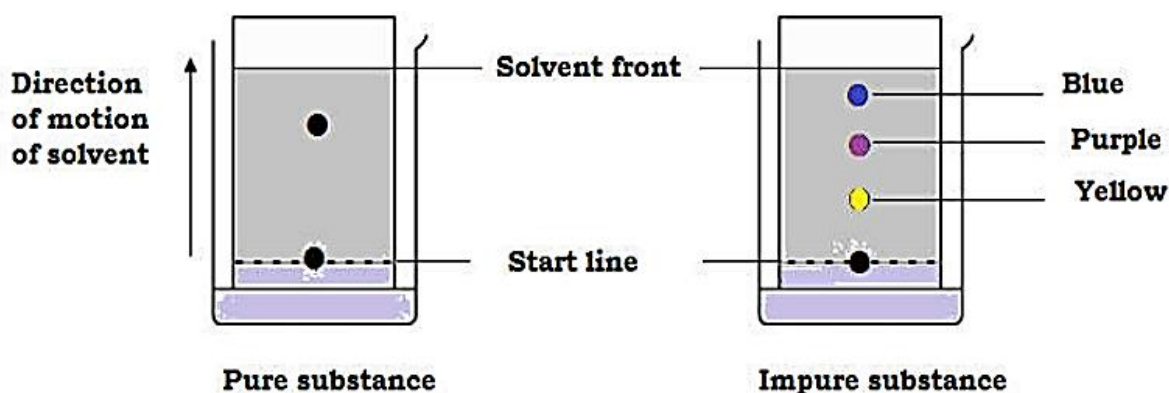
Procedure

- Using a pencil, a 'start line' is drawn near the bottom of the chromatography paper.
- The mixture to be separated is spotted on the start line.
- The bottom of the chromatography paper is dipped into the solvent and the solvent travels up the paper.
- The solvent picks up the substance being tested and carries them up the paper.
- The different components in the substance rise to different heights.
- The "solvent front" is marked. The solvent front is the furthest point reached by the solvent on chromatography paper.

INTERPRETING CHROMATOGRAM

Chromatogram is the pattern formed by substances that have been separated by chromatography. To determine purity of a substance, the chromatogram is interpreted as follows:

- A pure substance produces only one spot on the chromatogram.
- An impure substance produces two or more spots on the chromatogram.

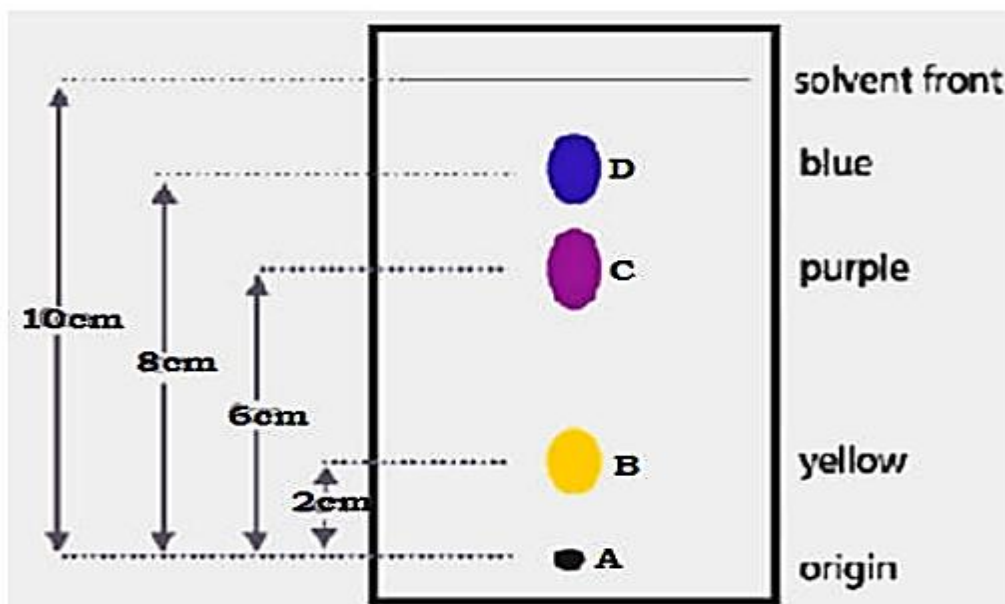


RELATIVE FLOW VALUES (R_f)

Relative flow value is the ratio of the distance travelled by the substance (spot) to the distance travelled by the solvent.

$$\text{Mathematically; } R_f = \frac{\text{distance travelled by the spot}}{\text{distance travelled by the solvent}}$$

Both distances are measured from the center of the sample spot on the base line.



From this sample chromatogram, the relative flow values for components B, C and D are calculated as follows:

1. For component **B**, $R_f = \frac{8}{10} = 0.8$
2. For component **C**, $R_f = \frac{6}{10} = 0.6$
3. For component **D**, $R_f = \frac{2}{10} = 0.2$

The relative flow value is always less than 1 because the distance moved by spot is always less than the distance moved by solvent.

INTERPRETTING RELATIVE FLOW VALUES

The relative flow value can be used to identify the purity of a substance.

- A pure substance has only one relative flow value.
- An impure substance has more than one relative flow value.

TESTS FOR WATER, IONS AND GASES

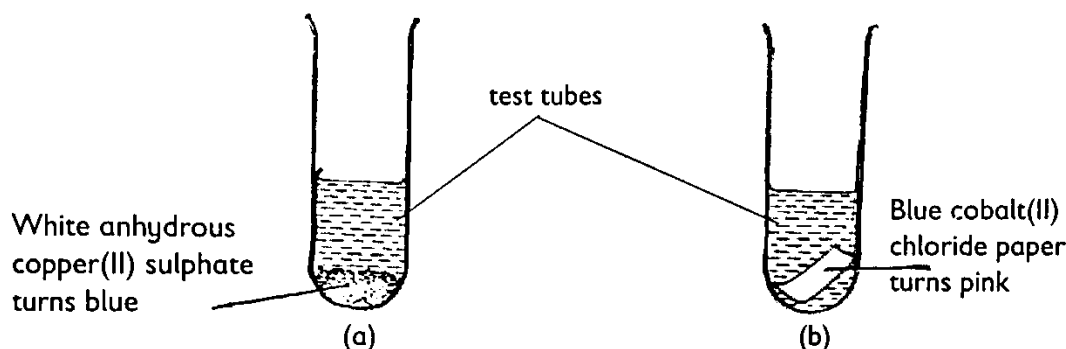
1. Testing for the presence of water

a. *Using anhydrous (II) copper sulphate*

- ☞ Put anhydrous copper (II) sulphate in a test tube. Anhydrous copper (II) sulphate is a white powder. When few drops of the test liquid is added to the anhydrous copper (II) sulphate, the colour of the powder changes to blue if the test liquid is water.

b. *Using anhydrous cobalt chloride*

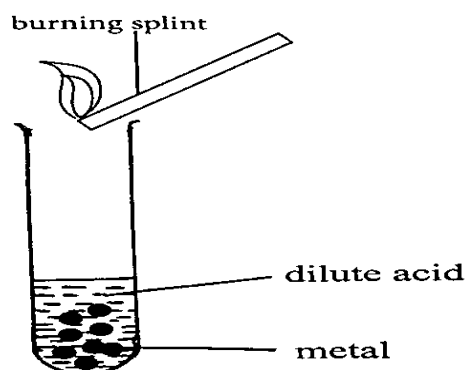
- ☞ Put the anhydrous cobalt chloride in a test tube.
- ☞ When few drops of the test liquid is added to the cobalt chloride, the colour of the powder changes to pink if the test liquid is water.



TESTING FOR THE PRESENCE OF GASES

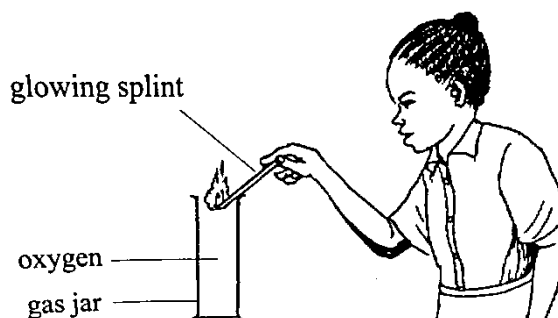
2. Testing for the presence of hydrogen

- ☞ A splint is lit and held near the opening of the tube that contains the sample.
- ☞ The stopper is then removed to expose the splint to the gas. If the gas under the test is hydrogen, it ignites with a “pop” sound.
- ☞ The pop sound is violent and in most cases it completely extinguishes the splint.



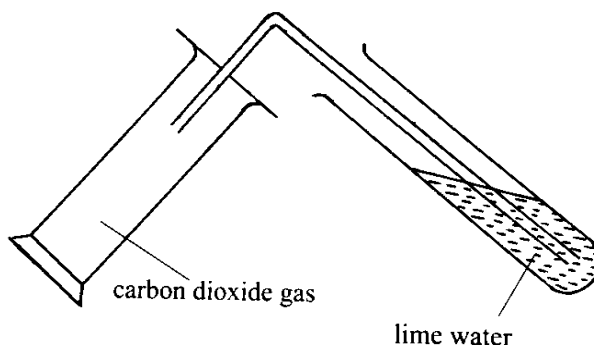
3. Testing for the presence of oxygen

- ☞ A glowing splint is held near the opening of the tube in which the sample is trapped.
- ☞ The stopper is then removed to expose the splint to the gas.
- ☞ If the gas under the test is oxygen, it ignites the glowing splint.



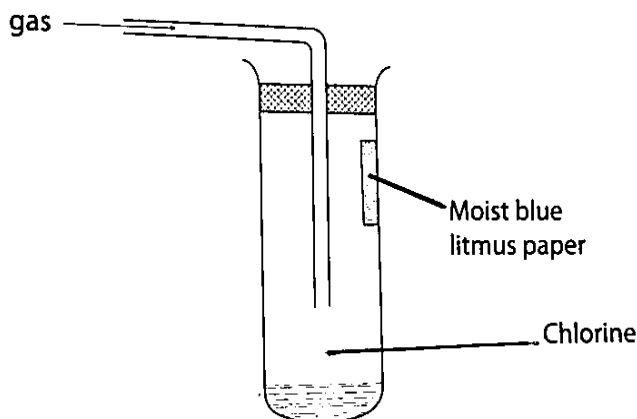
4. Testing for the presence of carbon dioxide

- ☞ Bubble the gas through lime water.
- ☞ If the gas under the test is carbon dioxide, lime water turns milky.
- ☞ This is so because the calcium hydroxide (chemical name for lime water) reacts with carbon dioxide to form calcium carbonate which is insoluble in water and thus forms a milky white precipitate.



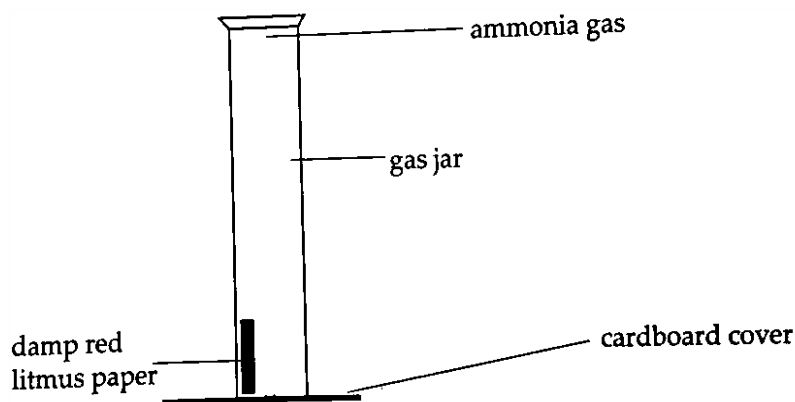
5. Testing the presence of chlorine gas

- ☞ Damp or moist blue litmus paper and expose it to the trapped gas.
- ☞ When the test gas is chlorine, it turns the colour of the blue litmus paper to red and then bleaches it white.
- ☞ $\text{Cl}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{HCl}(\text{aq}) + \text{HClO}$
- ☞ Hydrochloric acid turns the blue litmus paper red while the hypochlorous acid is responsible for bleaching the litmus paper.



Testing for the presence of ammonia gas

Damp or moist blue litmus paper and Expose it to the trapped gas.
When the test gas is ammonia, it will turn moist red litmus paper to blue.



6. Testing for sulphur dioxide gas

Sulphur dioxide has the following properties

- ✓ It is colourless
- ✓ It is a poisonous gas
- ✓ It dissolves in water to form an acidic solution

Activity 1

Aim: testing for sulphur dioxide

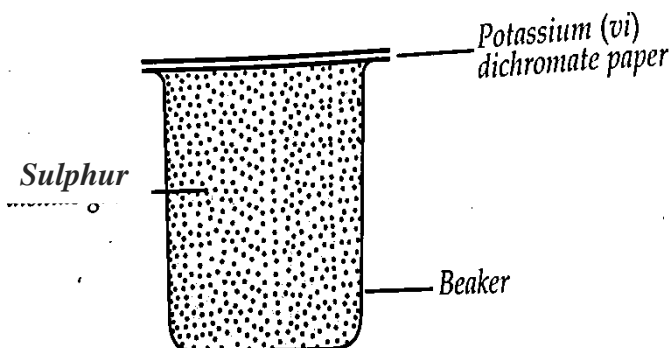
Materials: sulphur dioxide in a beaker and potassium dichromate (vi) paper

Procedure:

Put a freshly made potassium (vi) dichromate paper on top of open beaker containing sulphur dioxide and observe.

Expected results:

- Potassium (vi) dichromate paper which was orange in colour changes to green when exposed to sulphur dioxide.



Conclusion /discussion:

- Potassium (vi) dichromate paper which was orange in colour changes to green when exposed to sulphur dioxide.
- This is so because the dichromate ion ($\text{Cr}_2\text{O}_7^{2-}$) which is orange is reduced to (Cr^{3+}) ion, which is green.

7. Testing for anions

Activity 2

Aim: testing for the presence of nitrates

Materials:

- 2 compounds, x and y (one is a nitrate)
- 2 test tubes
- Test tube rack
- Dropper
- Iron sulphate solution
- Concentrated sulphuric acid

Procedure:

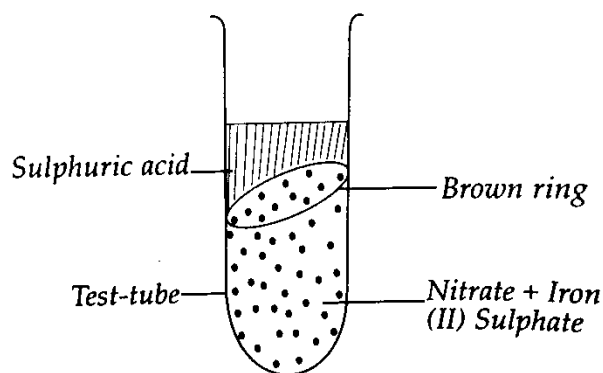
- Put one 5mls of a compound under test in a test tube.
- Add 3cm³ of iron sulphate solution into the test tube.
- Carefully add a few drops of concentrated sulphuric acid.

Expected results:

A brown ring is expected to form in the beaker.

Conclusion/discussion:

A solution of a nitrate compound gives a positive test in the activity. There is a brown ring formed when you mix a nitrate with an iron (II) sulphate solution and sulphuric acid. It is formed where the layer of concentrated sulphuric acid meets the solution of a nitrate and iron (II) sulphate solution.



Test	Anion	Test and observation
Silver nitrate followed by ammonia solution	Chloride(Cl^-)	Add a few drops of dilute nitric acid, followed by a few drops of silver nitrate solution. A white precipitate of silver chloride is formed. The precipitate is soluble in ammonia solution (causes eye damage)
	Bromide (Br^-)	Add a few drops of dilute nitric acid (Irritant) followed by a few drops of silver nitrate solution. A pale yellow precipitate of silver bromide is formed. The precipitate is slightly soluble in ammonia solution. (causes eye damage)
	Iodide (I^-)	Add a few drops of dilute nitric acid (irritant) followed by a few drops of silver nitrate solution. A yellow precipitate of silver iodide is formed. It is insoluble in ammonia solution. (causes eye damage)
Barium chloride	Sulfate (SO_4^{2-})	Add a few drops of barium chloride solution (harmful if swallowed), followed by a few drops of hydrochloric acid. A white precipitate of barium sulphate is formed
Hydrochloric acid	Carbonate (CO_3^{2-})	Add dilute hydrochloric acid to the solution (or add it to the solid). Bubbles of carbon dioxide are given off

8. Testing for Cations

CATION	TEST	OBSERVATION
Copper Cu^{2+} e.g. in CuCl_2 solution or solid. Blue in solution or as a solid.	Add few drops of dilute sodium hydroxide solution	Blue precipitate of copper hydroxide, which does not dissolve in excess NaOH. $\text{CuSO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s}) + \text{Na}_2\text{SO}_4(\text{aq})$ Or $\text{Cu}^{2+} + 2\text{OH}^- \rightarrow \text{Cu}(\text{OH})_2(\text{s})$
Iron (II) Fe^{2+} e.g. in FeCl_2 solution or as a solid. Often pale green in colour.	Add few drops of dilute sodium hydroxide solution	A dark green gelatinous precipitate is formed that does not dissolve in excess NaOH. It turns brown on standing as it is oxidised in air to Fe (III) hydroxide, $\text{Fe}(\text{OH})_3$. $\text{FeCl}_2(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s}) + 2\text{NaCl}(\text{aq})$ Or $\text{Fe}^{2+} + 2\text{OH}^- \rightarrow \text{Fe}(\text{OH})_2(\text{s})$
Iron (III) Fe^{3+} e.g. in FeCl_3 solution or solid; often yellow in solution and red/brown in solid.	Add few drops of dilute sodium hydroxide solution	A dark rusty brown gelatinous precipitate is formed that does not dissolve in excess NaOH $\text{FeCl}_3(\text{aq}) + 3\text{NaOH}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_3(\text{s}) + 3\text{NaCl}(\text{aq})$ Or $\text{Fe}^{3+} + 3\text{OH}^- \rightarrow \text{Fe}(\text{OH})_3(\text{s})$

Aluminium Al^{3+} These are colourless in solids and in solution	Add few drops of dilute sodium hydroxide solution	A white precipitate forms, but it dissolves in excess sodium hydroxide solution, because $\text{NaAl(OH)}_4(\text{aq})$ is soluble. Aluminium hydroxide is said to be amphoteric because it dissolves in an alkali and an acid $\text{AlCl}_3(\text{aq}) + 3\text{NaOH}(\text{aq}) \rightarrow \text{Al(OH)}_3(\text{s}) + 3\text{NaCl}(\text{aq})$ $\text{Al(OH)}_3(\text{s}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaAl(OH)}_4(\text{aq})$
Magnesium Mg^{2+}	Add few drops of dilute sodium hydroxide solution	A white precipitate of magnesium hydroxide, which does not dissolve in excess NaOH.
Calcium Ca^{2+}	Add few drops of dilute sodium hydroxide solution	A white precipitate of calcium hydroxide, which does not dissolve in excess NaOH.
Lead Pb^{2+}	Add few drops of dilute sodium hydroxide solution	A white precipitate forms, but it dissolves in excess sodium hydroxide solution,
Zinc Zn^{2+}	Add few drops of dilute sodium hydroxide solution	A white precipitate forms, but it dissolves in excess sodium hydroxide solution,
Ammonium NH_4^+ e.g. in NH_4SO_4	Add dilute sodium hydroxide solution and warm, test gas with pH paper	A smelly alkaline gas that turns pH paper blue is given off. Ammonia is released from the ammonium salt. $(\text{NH}_4)_2\text{SO}_4 + 2\text{NaOH}(\text{aq}) \rightarrow 2\text{NH}_3(\text{g}) + 2\text{H}_2\text{O} + \text{Na}_2\text{SO}_4(\text{aq})$ or $\text{NH}_4^+(\text{aq}) + \text{OH}^- \rightarrow \text{NH}_3(\text{g}) + \text{H}_2\text{O}$

UNIT 2

NITROGEN, SULPHUR AND PHOSPHORUS

NITROGEN

- ✎ It is a gas made up of molecules, each containing two atoms.
- ✎ It is inert (unreactive) under normal conditions. The inertness comes from the fact that the two atoms hold each other very strongly by a triple covalent bond which is very strong to break.

SOURCES OF NITROGEN

1. Nitrogen is obtained from the air. Naturally, air contain about 78% nitrogen
2. It is found in animal and plant proteins
3. It is also found in compounds known as nitrates

NITROGEN CYCLE

- ☞ This is a repeated process through which nitrogen gets into the soil from the air, plants and animals and out of the soil.

A. HOW NITROGEN GETS INTO THE SOIL

- Through the work of nitrogen fixing bacteria: Leguminous plants have bacteria in the nodules on their roots.

These bacteria take nitrogen from the atmosphere and convert it into nitrates. Nitrates are soluble compounds and plants absorb and convert them to make proteins.

- Through acid rain Nitrogen dioxide in the air dissolves in the rain making acidic solution which forms nitrates in the soil.

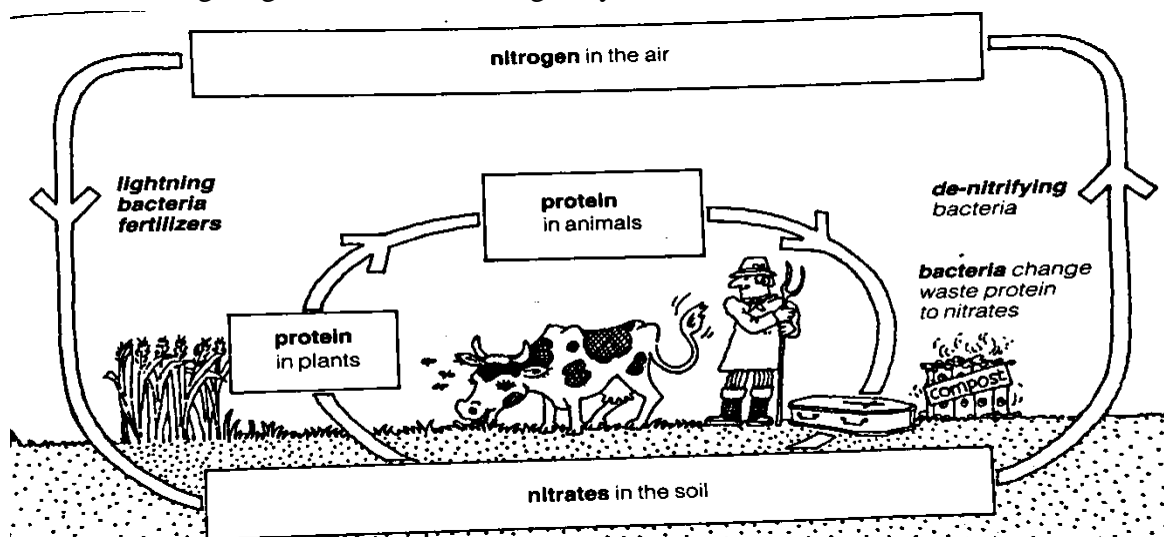
NB: Lightening is a necessary condition for formation of nitrogen dioxide gas.

- Organic matter (Decaying of dead plants and animals): Dead plants and animals produce nitrates as they decay. This is done by putrefying bacteria.
- Artificial fertilisers e.g. potassium nitrate, ammonium nitrate, urea CAN.

B. HOW NITROGEN GETS OUT OF THE SOIL

- By the work of denitrifying bacteria a process called volatilisation.
- Plants use nitrates from the soil and convert them into proteins.
- Through immobilisation microorganisms use nitrogen from the soil to build up their bodies.

The following diagram shows the nitrogen cycle:



PHYSICAL PROPERTIES OF NITROGEN

- ✓ It is a colourless gas.
- ✓ It is odourless gas
- ✓ It is almost insoluble in water
- ✓ It is lighter than air

CHEMICAL PROPERTIES OF NITROGEN

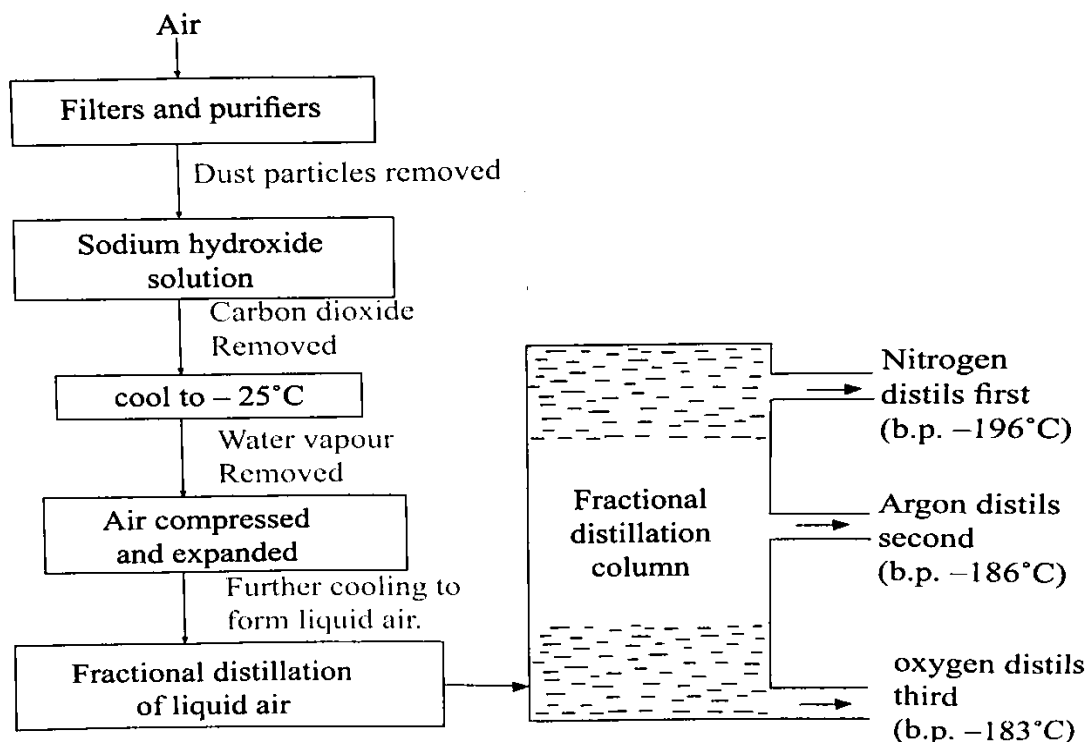
- At room temperature, it is unreactive (we inhale and exhale N_2 without any danger)
- At high temperature, nitrogen reacts with alkali metals to form metal nitrides. For example, when sodium reacts with nitrogen to form sodium nitride.
- Nitrogen reacts with alkaline earth metals. Examples of alkaline earth metals are magnesium and calcium.
- Nitrogen reacts with hydrogen gas to form ammonia gas.

INDUSTRIAL ISOLATION OF NITROGEN FROM THE AIR

Nitrogen is removed from the air by a process known as fractional distillation or liquefaction. This process is comprised of a series of steps as follows:

- ☞ Air is purified by passing it through filters.
- ☞ Carbon dioxide is removed by passing the air through concentrated sodium hydroxide solution.
- ☞ The remaining air is then cooled to about -25 degrees Celsius to remove water which solidifies as ice.
- ☞ The air is compressed to about 200atm. The cooled compressed air is allowed to expand rapidly. This process is repeated until all the liquid air liquefies.
- ☞ The liquid air is then passed into a fractionating column containing compartment. Each compartment is slightly at a higher temperature than the one above it. The components of the liquid air then separate according to their boiling points.
- ☞ Nitrogen distils off first because it has a lower boiling point of -196 degrees Celsius followed by argon at -186 degrees Celsius and finally oxygen at -183 degrees Celsius. These components of air are collected and stored in metal tanks.

The following diagram illustrates the separation of air into its components:



USES OF NITROGEN

1. It is used to form ammonia which is a raw material for the production of fertilizers.
2. Liquid nitrogen is very cold, so it is used for freezing food and other materials that need to be kept frozen.
3. Nitrogen gas, usually mixed with argon is used for filling incandescent light bulbs.

4. Nitrogen is used in food processing. When food is being packaged, nitrogen is used to keep off oxygen and the food stays fresh for a long period of time because there is no oxidation. Most bacteria cannot survive if there is no oxygen
5. It is used to mend leaking pipes. When liquid nitrogen is poured on the pipe, it freezes the liquid inside the pipe while the repairing process is being done.
6. In plant growth, nitrogen
 - Is a major component of proteins
 - Increases leaf area and vegetative growth in crops
 - Is responsible for the deep green colour, chlorophyll.
 - Makes plants and their fruits succulent e.g. melons

AMMONIA

Ammonia is a compound of nitrogen and hydrogen. Most of the world's supply of fertilizer is obtained from ammonia. At room temperature and pressure, it is a gaseous substance.

Physical properties of ammonia

1. It is a gas at room temperature
2. It is lighter than air
3. It readily dissolves in water to form a basic solution
4. It has a pungent smell and it can be felt in urinaries and animal houses.

Chemical properties of ammonia

1. It turns the colour of the red litmus paper to blue. This means ammonia is a base.
2. It reacts with hydrogen chloride to form ammonium chloride.

PREPARATION OF AMMONIA

A. Laboratory preparation of ammonia

In the laboratory, ammonia is prepared by reacting an ammonium salt with a basic compound. For example, Calcium hydroxide can be heated together with ammonium chloride to produce calcium chloride, water and ammonia.



Activity 2

AIM: Preparing ammonia gas in the laboratory

Materials:

- | | |
|------------------------|----------------------|
| • Gas jar | • Calcium oxide |
| • Large test tube | • Calcium hydroxide |
| • Ammonium chloride | • Holed cork stopper |
| • Red wet litmus paper | • Glass wool |
| • Delivery tube | • Mortar and pestle |

Procedure:

1. Grind the mixture of calcium hydroxide and ammonium chloride and put them in the large test tube.

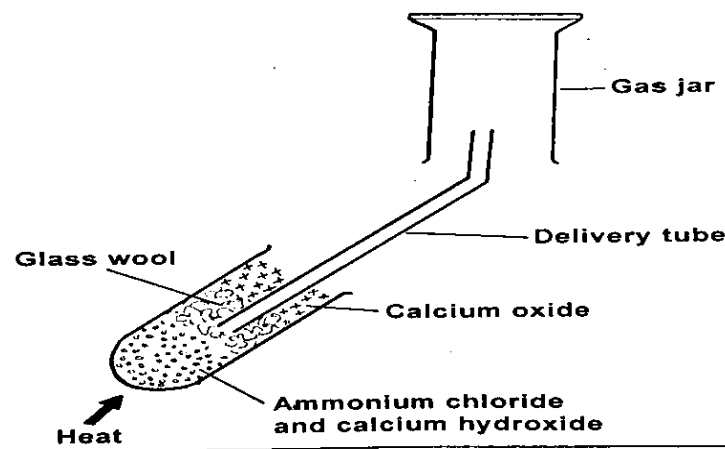


Figure 2.8 Preparing ammonia gas

2. Heat the test tube gently. What do you observe in the delivery tube?
3. Put the moist red litmus paper close to the mouth of the gas jar. What colour does the red litmus paper become? Why is the gas jar positioned upside down above the delivery tube?

Results and discussion:

- ✎ As the mixture is heated, ammonium chloride and calcium hydroxide will react to produce calcium chloride, water and ammonia gas:

$$\text{Ca(OH)}_2(\text{s}) + 2\text{NH}_4\text{Cl}(\text{s}) \rightarrow \text{CaCl}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l}) + 2\text{NH}_3(\text{g})$$
- ✎ Calcium oxide is used as a drying agent because the usual drying agents such as sulphuric acid and calcium chloride would react with ammonia gas.
- ✎ The collecting gas jar is positioned upside down over the delivery tube because ammonia is less dense than air and can only be collected this way. This method of gas collection is called upward delivery.

B. Industrial preparation of ammonia (The Haber process)

Industrially, ammonia is made by reacting hydrogen gas with nitrogen gas.



Optimum conditions necessary for the Haber process

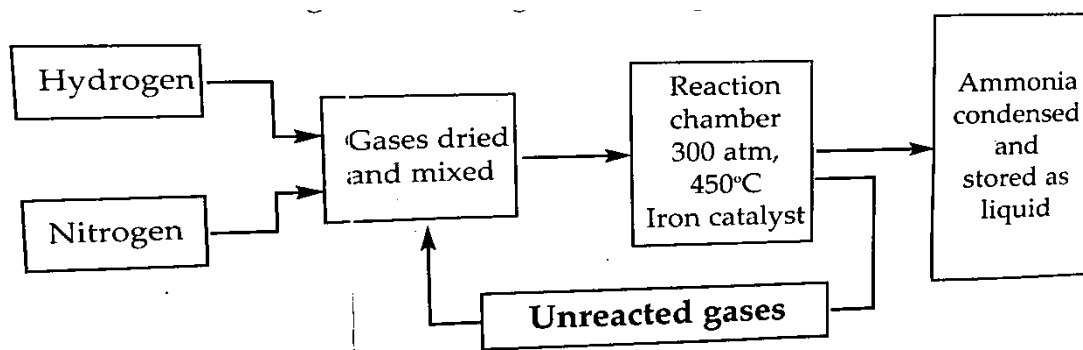
- A mixture of nitrogen and hydrogen in the ratio of 1:3 by volume.
- A pressure of about 200 to 300 atmospheres
- A temperature of about 400 to 500⁰C
- Finely divided iron catalyst mixed with Aluminium oxide as a promoter of the catalyst.

Process:

- ✎ A mixture of nitrogen and hydrogen in the ratio of 1:3 by volume is purified, dried and introduced into the reaction chamber where they are reacted.
- ✎ High pressure and low temperatures give high yield, but the higher the pressure, the greater the cost of equipment to produce and maintain the pressure.
- ✎ At low temperatures, the reaction is slow. Therefore a compromise has to be made between a reasonable rate of reaction and a reasonable yield of ammonia.

- Most industries use a temperature of about 400 to 500°C and a pressure of 200 to 300 atmospheres.
- Finely divided iron catalyst is used to increase the rate of reaction. The catalyst is mixed with Aluminium oxide to increase its efficiency.

Figure 2, 7 is the diagram showing the Haber process.



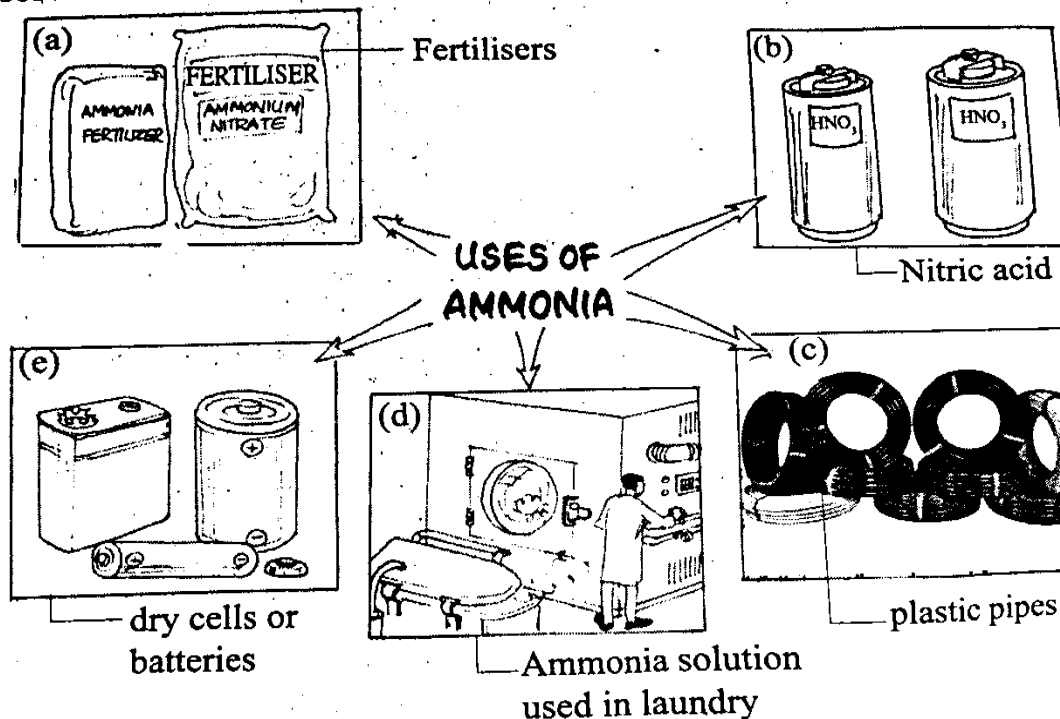
Note:

- The reaction is reversible and conditions must be adjusted to ensure that forward reaction is more favoured than the backward reaction.
- This is done by adjusting the temperature, pressure and the use of catalysts.

USES OF AMMONIA

- It is used in making plastics and fibres such as nylon.
- It is used in the manufacturing of fertilizers (e.g. Ammonium nitrate).
- It is used in the manufacturing of ammonium chloride which is used in dry cells.
- It is used in the making of explosives.
- It is used as a solvent in cleaning.

as nylon.



☒ Nitric acid is a common mineral acid in school laboratories. It is an acid which has been known for many years by scientist. This acid is a powerful oxidizing agent.

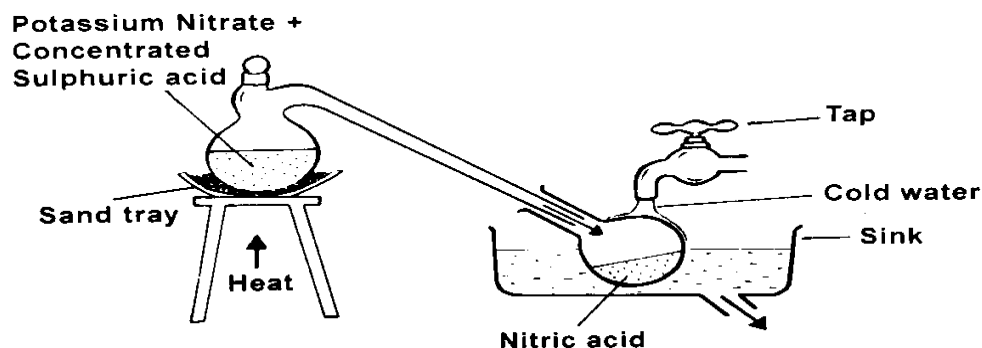
In the laboratory, nitric acid is prepared by heating potassium nitrate with concentrated sulphuric acid

Nitric acid is prepared in large quantities using a method known as **Ostward method**.

- ## The Ostward method

-
- The diagram illustrates the Ostwald process for the manufacture of nitric acid. It begins with two input streams: 'Ammonia from Haber process' and 'Oxygen from liquid air'. These streams are combined in a 'Mixed and compressed' stage. The mixture then enters a 'Reaction chamber' where the temperature is maintained at 900°C and a Platinum and Rhodium catalyst is used. The reaction produces $\text{NO}_2 + \text{H}_2\text{O}$. This mixture is then passed through a series of wash bottles, where 'Water' is added to absorb the gases. The final product is 'Nitric acid'.

- ✖ In the laboratory, nitric acid is prepared by heating potassium nitrate with concentrated sulphuric acid in a retort (long necked distillation apparatus) flask.
- ✖ The reaction happens as follows; $\text{KNO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{KHSO}_4(\text{s}) + \text{HNO}_3(\text{g})$
- ✖ The fumes of nitric acid are cooled by passing them over cold running water.
- ✖ This produces liquid nitric acid as in the following diagram.



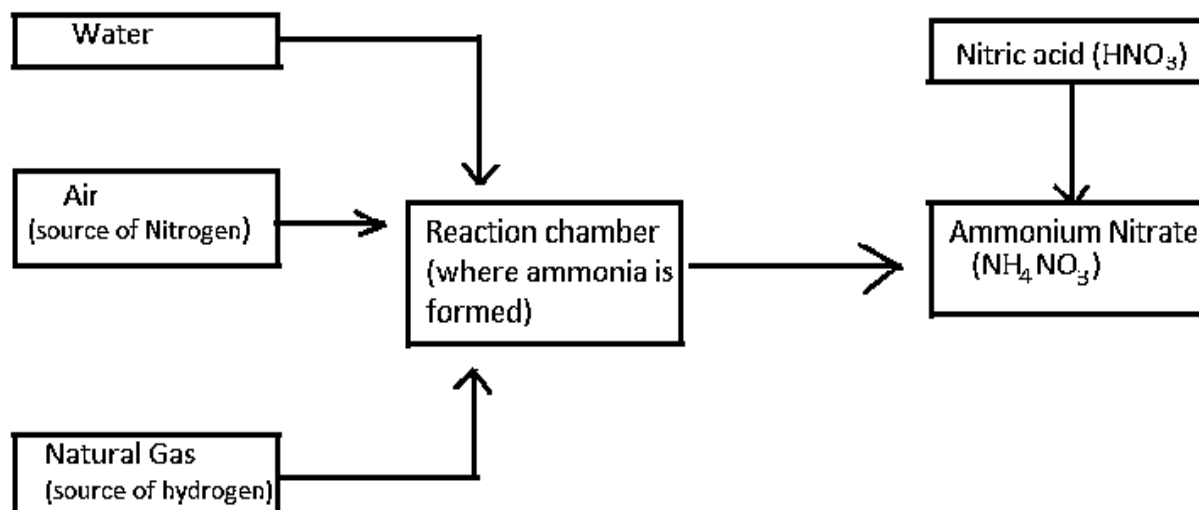
USES OF NITRIC ACID

- ✓ The most important use of nitric acid is in the manufacturing of nitrate fertilisers.
- ✓ It is used as a laboratory chemical when doing experiments in schools.
- ✓ It is used in the manufacturing of explosives like T.N.T.
- ✓ It is used in making nitrogen based compounds like plastics (e.g. nylon).
- ✓ It is used as an oxidizing agent in textiles.
- ✓ It is used in purification of precious minerals like Gold.
- ✓ It is also used in the manufacturing of nitrate salts, dyes and drugs.

AMMONIUM NITRATE FERTILIZER

- ☞ It is a white solid crystal that is highly soluble in water.
- ☞ It is manufactured by reacting nitric acid with ammonia.

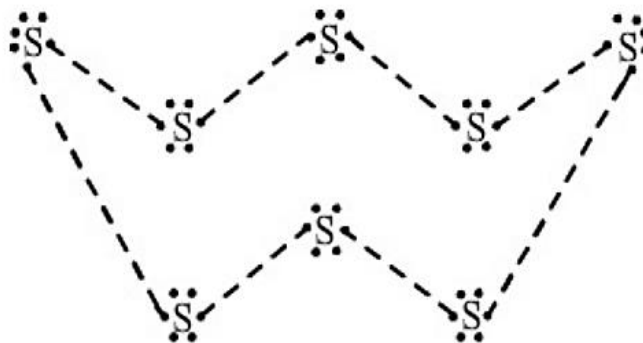
The following flow chart shows the production of fertilizers called ammonium nitrate.



- ☞ The ammonium nitrate is crystallised into granules which can easily be spread on the land. When ammonium nitrate is dissolved in water, it forms two ions; ammonium ions (NH_4^+) and nitrate ions (NO_3^-). Plants can absorb these ions through their roots.
- ☞ Ammonium nitrate is hygroscopic. Because of this, it is packed with a drying agent to prevent the granules from sticking together.

SULPHUR

- ✎ It is a yellowish powder found in group 6 of the periodic table.
- ✎ It is identified by the letter S on the periodic table and its atomic number is 16. It is a multivalent element.
- ✎ Under normal conditions, it forms cyclic octatomic molecules with the formula S_8 .



SOURCES OF SULPHUR

- ✓ It comes from large underground sulphur beds in Poland, Mexico and the USA.
- ✓ Some is extracted from crude oil and natural gas. Crude oil is also called petroleum
- ✓ Sulphur is found combined with metals in metal ores. Examples of such compounds include iron sulphide (FeS) and lead sulphide (PbS).

EXTRACTION OF SULPHUR BY FRASCH PROCESS

- ✎ The sulphur beds in Texas and Louisiana in USA are as much as 30 metres thick and occur at a depth of 160 metres. Between the surface and sulphur deposits, there are layers of sand which makes it difficult to use common methods of mining.
- ✎ The basic principle of the Frasch process is that sulphur is melted underground and pumped up to the surface of the earth using concentric pipes as shown in the following diagram.

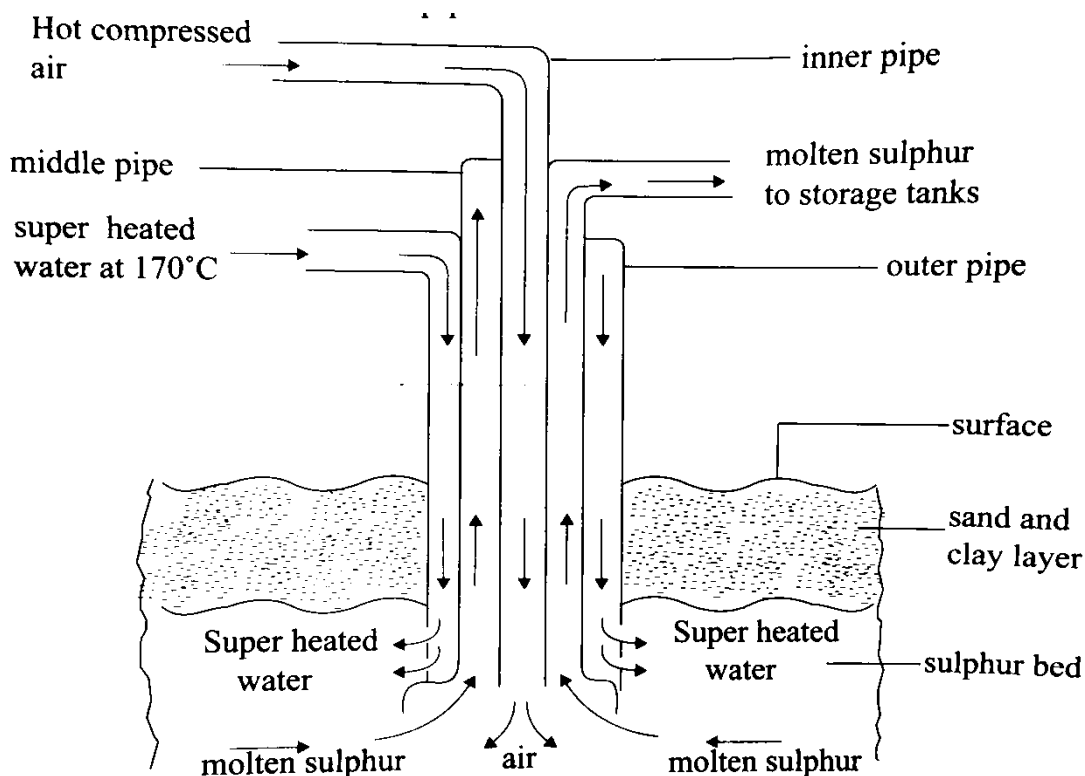


Fig. 2.14: Extraction of sulphur by Frasch process

Functions of the three pipes in the extraction process

- ☞ The outermost pipe which has the widest diameter delivers the superheated water to the sulphur bed
- ☞ The second pipe (held above the sulphur deposits) serves as the outlet of molten sulphur
- ☞ The inner pipe brings hot compressed air into the sulphur deposits to force molten sulphur up the middle pipe to the storage tanks.

The extraction process

- ☞ Super-heated water at 170°C (under 10 a.t.m) is pumped down the outermost pipe.
- ☞ The superheated water melts the sulphur at 113°C. The compressed air at about 15 atmospheres is pumped down the innermost pipe to force up the molten sulphur.
- ☞ This molten sulphur raises through the annular space between the inner most pipe and the second pipe. The molten sulphur is collected and stored in tanks.
- ☞ Sulphur cools and solidifies into solid blocks.

PHYSICAL PROPERTIES

1. It is a brittle solid (i.e. can easily break)
2. It has a yellow colour
3. Like other non-metals, it does not conduct electricity
4. It has a relatively low melting and boiling point.
5. Like most non-metals, it is insoluble in water. It only dissolves in organic solvents such as methylbenzene, carbon tetra chloride and carbon disulphide

CHEMICAL PROPERTIES

1. It reacts with oxygen to form sulphur dioxide i.e. $S(s) + O_2(g) \rightarrow SO_2(g)$
2. Sulphur reacts with hydrogen to form hydrogen sulphide
3. It reacts with metals to form metal sulphides e.g. $Zn(s) + S(s) \rightarrow ZnS(s)$

USES OF SULPHUR

1. Sulphur is used for vulcanization of rubber (i.e. it is added to rubber to make it tough and increase elasticity)
2. Sulphur is added to cement to make it resistant to acid. This is most used in flooring factories places where acid is spilled.
3. Sulphur is used in the production of sulphuric acid
4. Some powdered sulphur is used as an insecticide. It can be spread on plants to kill or drive away insects that feed on plants.
5. Sulphur is used in the production of drugs, matches, gun powder and papers.

SULPHURIC ACID

- ☞ Sulphuric acid is a very useful compound of sulphur. It is manufactured in large quantities than any other product. It is an oily colourless liquid.
- ☞ It forms a very strong acid when it is added to water.

Warning

- *Both teachers and students must take a great care when using sulphuric acid. It can cause severe burns if it touches the skin.*

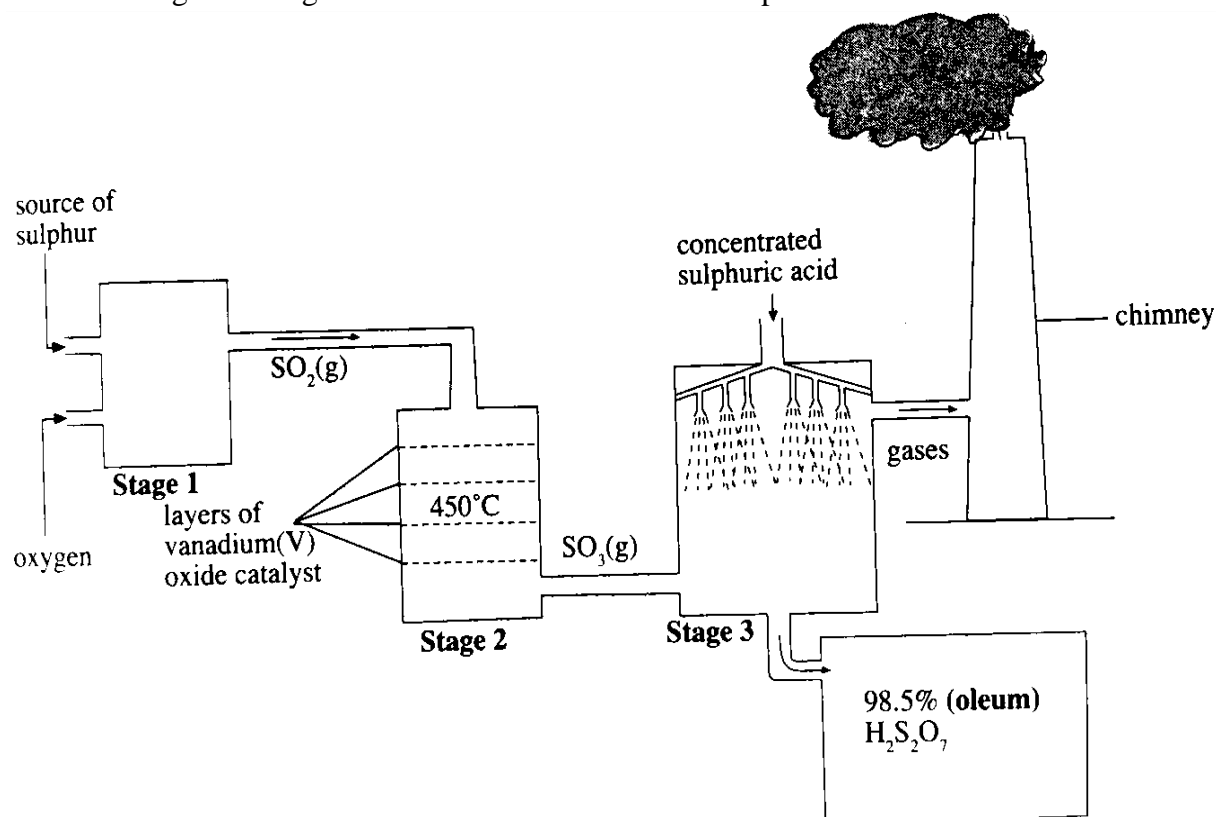
- Always wear protective clothing and gloves when working with it.

MANUFACTURE OF SULPHURIC ACID

Sulphuric acid is made through a process called contact process. This process has four stages as follows:

1. In the first stage, sulphur reacts with oxygen to produce sulphur dioxide.
2. In the second stage sulphur dioxide reacts which is produced in the first stage is made to react with more oxygen to produce sulphur trioxide (SO_3). This chemical reaction is reversible (i.e. it goes in both directions). To increase the rate at which SO_3 is produced, vanadium oxide (V_2O_5) is used as a catalyst.
3. The third stage involves the reaction of sulphur trioxide with concentrated sulphuric acid to form fuming sulphuric acid (oleum) according to the following chemical reaction
4. In the last stage oleum is hydrated to form concentrated sulphuric acid.

The following flow diagram illustrates the formation of sulphuric acid.

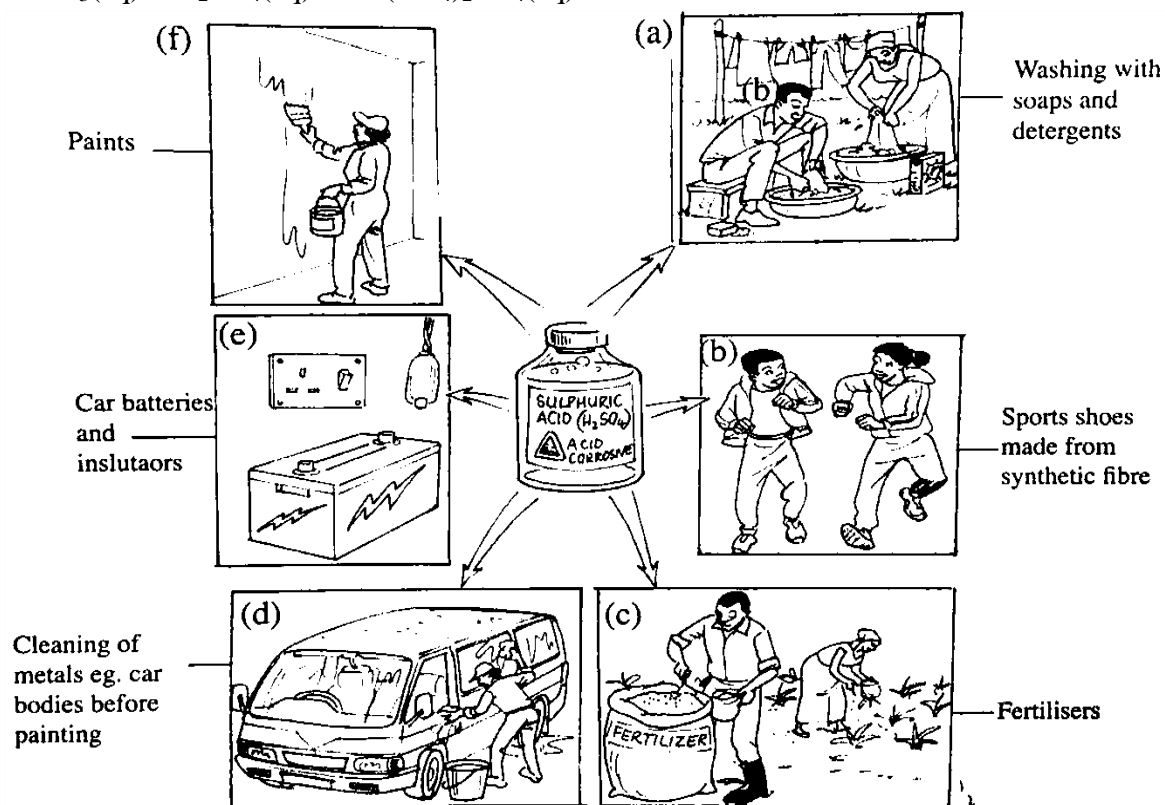
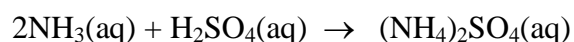


PROPERTIES OF SULPHURIC ACID

- ✓ Sulphuric acid is a diprotic acid (i.e. it can donate two protons (2H^+) to a base.
- ✓ Sulphuric acid reacts with bases to form salts and water. An example of a base is sodium hydroxide (NaOH).
- ✓ It reacts with metals to form salts (known as sulphates) and hydrogen gas.
- ✓ It also reacts with carbonates to form metal sulphates, carbon dioxide and water.

USES OF SULPHURIC ACID

1. It is used for making paints, plastics and paper
2. It is used for making insecticides and drugs
3. It is used for making synthetic fibres and flocculents such as aluminium sulphate ($\text{Al}_2(\text{SO}_4)_3$). A flocculent is used in water treatment processes. It makes small particles to come together and settle down as large particles.
4. It is a raw material for the manufacturing of dyes, soaps and detergents
5. It is used in refining of crude petroleum.
6. Sulphuric acid is used to remove rust on iron and steel to clean them before they are used in making important iron materials.
7. Car batteries use sulphuric acid as a conducting solution called electrolyte.
8. Sulphuric acid is used as a dehydrating agent
9. Sulphuric acid is used to make fertilisers. It is reacted with ammonia to form ammonium sulphate according to the reaction below;



SULPHATES

Other important compounds of sulphur are sulphate such as iron sulphate, copper sulphate, and magnesium sulphate. These are salts produced from sulphuric acid.

FORMATION OF SULPHATES

Sulphates are formed in three main ways;

- ✓ By neutralizing a dilute acid with an alkali
- ✓ By reacting a concentrated or dilute acid with appropriate metals.
- ✓ By double decomposition of soluble salts

USES OF SULPHATES

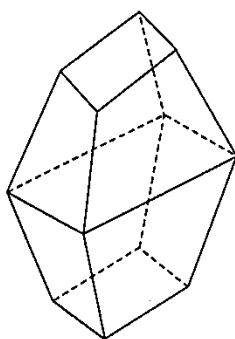
The following are some uses of sulphates

1. For making fertilizers e.g. ammonium sulphate fertilizers.
2. For making plaster of Paris (POP) which is used to support bone fracture. The sulphate used for this role is calcium sulphate.
3. For making medicines e.g. eye ointments, mouth washers
4. For making stomach ant-acid tablets eg Epsom salt ($\text{MgSO}_4 \cdot \text{H}_2\text{O}$)
5. For making fungicides and dyes (hydrated zinc sulphate ($\text{ZnSO}_4 \cdot 5\text{H}_2\text{O}$) is used for this)
6. For making white paints and diagnostic x-ray studies, e.g. Barium sulphate (BaSO_4)

ALLOTROPES OF SULPHUR

Sulphur exists in two main allotropes namely;

- ✓ Rhombic sulphur
- ✓ monoclinic sulphur



(a): Octahedrally shaped rhombic sulphur



(b): Monoclinic sulphur

RHOMBIC SULPHUR

- ❖ This is also called alpha sulphur (α sulphur). It consists of large yellow crystals with an octahedral shape. Crystals of this allotrope are stable below 96°C . Above this temperature, they slowly change to the monoclinic form. The S_8 molecules of rhombic sulphur interlock with each other.

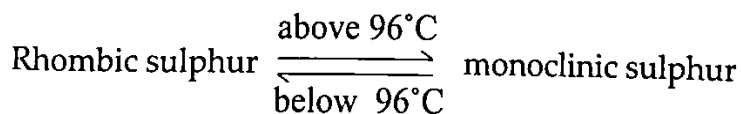
MONOCLINIC SULPHUR

- ✎ Monoclinic sulphur is also called beta sulphur (β sulphur).
- ✎ It consists of needle shaped transparent crystals which are stable above 96°C .
- ✎ Below this temperature, they gradually change to rhombic form. The S_8 molecules of monoclinic sulphur are stacked.

Characteristics of rhombic and monoclinic sulphur

No.	characteristics	Rhombic sulphur	Monoclinic sulphur
1	Shape of the crystal	octahedral	Long prism needle like
2	Colour of the crystal	yellow	yellow
3	Stability	Below 96°C	Above 96°C
4	Density	2.07g/cm^3	1.96g/cm^3
5	Melting point	113°C	119°C

Transition temperature of the two allotropes is 96°C .



PHOSPHORUS

- ✎ It is a non-metal element found in-group 5 and period 3 of the periodic table.
- ✎ Its atomic number is 15 and the electron configuration is 2:8:5
- ✎ It has a valency of 3 but it can also have a valency of 5.

SOURCES OF PHOSPHORUS

Phosphorus has various sources such as;

1. It occurs naturally in rocks such as rock phosphate
2. It is found in banana peels
3. It is found in crab shells
4. It is also found in coal

USES PHOSPHORUS

1. Red phosphorus is used in the manufacturing of safety matches
2. It is also taken as a food additive because it forms part of the enamel.
3. Phosphorus in form of phosphates are used in making tooth paste
4. Phosphorus is used in the making of phosphoric acid which is a raw material for the production of Fertilizers
5. It is used in making detergents e.g. surf
6. White phosphorus is used in making Rat poison
7. It is also used in making bombs

PHYSICAL PROPERTIES OF PHOSPHORUS

1. It is a non-metal
2. It has allotropic property. It exists mainly as red phosphorus and white phosphorus.
3. It has low melting and boiling points
4. It is insoluble in water

ALLOTROPE OF PHOSPHORUS

The main allotropes of phosphorus are:

- ✓ White phosphorus
- ✓ Red phosphorus

WHITE PHOSPHORUS

- ☞ It is a white solid that turns yellow upon exposure to light.
- ☞ Consists of P_4 molecules
- ☞ It has a melting point of 44°C and boiling point of 280°C
- ☞ It is rapidly oxidized in air under normal temperatures.
- ☞ White phosphorus shows phosphorescence(i.e. glowing in the dark)

- ☞ It is highly flammable and ignites on its own when exposed to air at about 30°C.
- ☞ It is highly toxic (gives off very poisonous vapour).
- ☞ It is less stable as compared to red phosphorus
- ☞ It does not react with acids
- ☞ It has a density of 1.2g/cm³

RED PHOSPHORUS

- ☞ It is a deep red powder
- ☞ It is a giant structure
- ☞ It has a density of 2.34g/cm³
- ☞ it sublimates at 400°C
- ☞ it is stored dry because it cannot be oxidized under normal temperatures
- ☞ It does not catch fire on its own unlike white phosphorus
- ☞ It is the stable form at all temperatures.
- ☞ It reacts with acids

CHEMICAL PROPERTIES OF PHOSPHORUS

1. It burns readily with oxygen to give phosphorus (V) oxide which dissolves in water to form phosphoric acid.
 Phosphorus + oxygen → phosphorus pentoxide

$$\text{P}_4(\text{s}) + 5\text{O}_2(\text{g}) \rightarrow 2\text{P}_2\text{O}_5(\text{s})$$
 phosphorus pentoxide + water → phosphoric acid

$$\text{P}_2\text{O}_5(\text{s}) + 3\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{H}_3\text{PO}_4(\text{g})$$
 Or
$$\text{P}_4\text{O}_{10}(\text{s}) + 6\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{H}_3\text{PO}_4(\text{g})$$
2. Phosphorus burns in chlorine gas to form phosphorus trichloride (PCl₃) or phosphorus pentachloride (PCl₅) when excess chlorine is supplied.
$$\text{P}_4(\text{s}) + 10\text{Cl}_2(\text{g}) \rightarrow 4\text{PCl}_5(\text{s})$$
3. White phosphorus reacts with sodium hydroxide to form phosphinate and phosphine i.e
$$3\text{NaOH}(\text{aq}) + \text{P}_4(\text{s}) + 3\text{H}_2\text{O}(\text{l}) \rightarrow 3\text{NaH}_2\text{PO}_2(\text{aq}) + \text{PH}_3(\text{g})$$
4. Phosphorus reacts with metals to form metal phosphides.

UNIT 3

CHEMICAL BONDING AND PROPERTIES OF MATTER

- ☞ **An atom;** is the smallest particle of matter that can take part in a chemical reaction.
- ☞ All matter is made up of atoms.
- ☞ Atoms of most elements are very reactive and do not exist in “Free State.”
- ☞ Chemical reaction refers to rearrangement of atoms to form new substances.
- ☞ When atoms take part in a chemical reaction, they result into the formation of molecules which can either be elements or compounds.

WHY DO CHEMICAL REACTIONS TAKE PLACE?

Chemical reactions happen so that atoms which are involved attain the electron configuration of a noble gas (Octet rule or duplet rule) thereby:

- i. Finding stability
- ii. Reaching a lower energy level
- iii. Experiencing greater forces of attraction.

AN ELEMENT

- ✍ It is a substance that is made up of the same kind of atoms.
- ✍ The molecules of elements contain two or more similar atoms chemically joined together.
- ✍ For example, a molecule of hydrogen contains two hydrogen atoms combined together and is written as H_2 .
- ✍ Other examples of elements are molecular oxygen, Molecular chlorine, Nitrogen and bromine.

HOW DO ATOMS FIND STABILITY?

- By gaining electrons.
- By losing electrons.
- By sharing electrons.

VALENCY: is the number of chemical bonds an atom can make. In Ionic compounds it is given by the number of electrons lost or gained. In molecular substances it is given by the number of electrons offered for sharing.

A COMPOUND

- ✍ It is a substance that is made up of different kinds of atoms.
- ✍ Molecules of compounds contain two or more types of atoms chemically combined together.
- ✍ For example hydrogen chloride is a compound that is made up of one hydrogen atom and one chlorine atom.
- ✍ Other examples of compounds are ethane, ethane and glucose. Compounds are divided into two categories namely ionic compounds and covalent compounds.

TYPES OF CHEMICAL BONDING

1. Covalent bonding
2. Ionic bonding
3. metallic bonding

5. COVALENT BONDING

This is the type of bonding that involves the sharing of electrons between non-metals.

WHAT FORMS THE COVALENT BOND?

The attraction between the shared electrons with the nuclei of the atoms involved.

- As an example of covalent bonding, Consider carbon (C) and hydrogen (H) reacting to form methane. Carbon has 4 electrons in the outermost shell. It requires 4 electrons to attain an electron configuration of a noble gas.

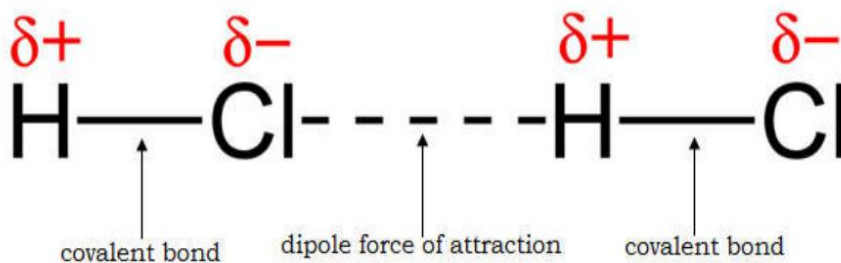
- Hydrogen has 1 electron in the outermost shell. It requires 1 electron to reach an electron configuration of a noble gas.
- When carbon and hydrogen react one carbon atom offers four electrons for sharing to four hydrogen atoms. Each of the four hydrogen atoms offers one electron for sharing with the carbon atom.
- The hydrogen atom with one shell and two electrons in the outermost shell attains a noble gas electron configuration.
- The carbon atom now with eight electrons in the outermost shell also attains an electron configuration of a noble gas.
- The attraction between the shared electrons and the nuclei of carbon and hydrogen forms the covalent bond

PHYSICAL PROPERTIES OF COVALENT (MOLECULAR) COMPOUNDS

1. Most of them are gaseous or volatile liquids at room temperature.
2. Most of them have a scent(a characteristic odour or smell)
3. Covalent compounds tend to be soft and relatively flexible. This is largely because covalent bonds are relatively weak and easy to break.
4. Covalent compounds do not conduct electricity. They dissolve into molecules rather than dissociating into ions so they typically do not conduct electricity.
5. Most covalent compounds have a relatively low melting and boiling points.

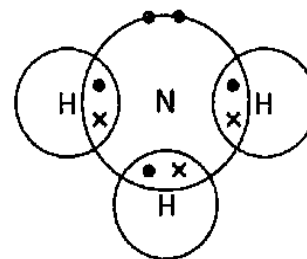
TYPES OF COVALENT BONDS

- (a) **Polar covalent bond:** this is a covalent bond in which electrons are unequally distributed between the atoms which are involved. It involves elements which have varying values of electronegativity. This happens when the atoms have electronegativity differences. Electrons are attracted to the more electronegative atom.
- The more electronegative atom acquires a partial negative charge (symbolized δ^-) and the less electronegative atom acquires a partial positive charge (symbolized δ^+). The bond between two polar molecules is called a dipole.
 - Molecules that have partial charges are called polar covalent molecules. The existence of partial charges in some covalent compounds enables them to conduct electricity in aqueous solutions. A good example of polar covalent compound is hydrogen chloride (HCl).



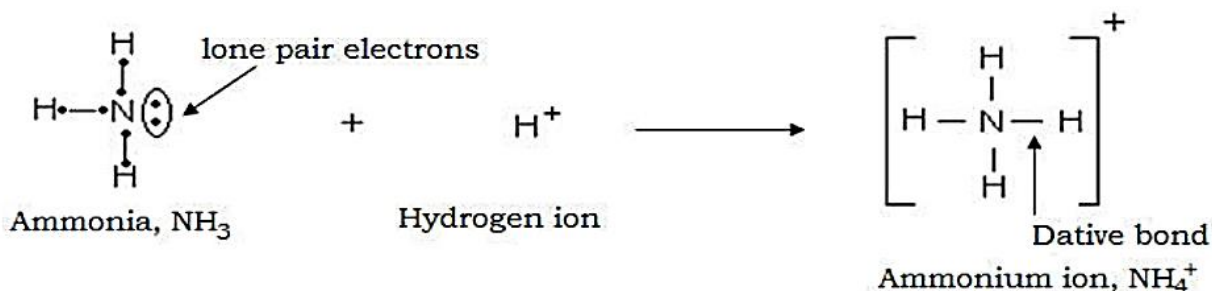
- (b) **Non-polar covalent bond:** this is a covalent bond in which electrons are equally distributed between the atoms which are involved. It involves elements which have nearly equal values of electronegativity e.g. hydrogen and hydrogen in the formation of a molecule of hydrogen. Molecules that have non-polar bonds are called non-polar covalent molecules. They do not have partial charges as a result they do not conduct electricity.

- (c) **Single covalent bond:** this is formed when one pair of electrons is shared between the two non-metal atoms e.g. in chlorine gas and in hydrogen gas.
- The following diagram shows single covalent bonds in a molecule of ammonia



A molecule of ammonia

- (d) **Double covalent Bond:** this is formed when two pairs of electrons are shared between the two nonmetal atoms e.g. e.g. in CO_2 ($\text{O}=\text{C}=\text{O}$).
- (e) **Triple covalent Bond:** This type of bond is formed when three pairs of electrons are shared between the two nonmetal atoms e.g. in nitrogen gas, N_2 ($\text{N}\equiv\text{N}$).
- (f) **Dative/Coordinate Bond:** This is formed in a covalent bond where the shared pair of electrons is provided by one of the bonded atoms. The atom which provides both electrons for the covalent bond is called the DONOR.
- The atom which accepts both electrons is called the ACCEPTOR e.g. in Ammonium ion and hydronium ion.
 - The following diagram shows the formation of a dative bond in ammonium ions



TYPES OF MOLECULAR SUBSTANCES

I. Polar Covalent Compounds

- ☞ These are formed between two atoms which have combined where one is more electronegative than the other resulting into formation of partial charges.
- ☞ Partial charges form because the electron density favours the more electronegative element and electrons spend more time in its vicinity.
- ☞ An example is hydrogen chloride gas (HCl). Chlorine is more electronegative than hydrogen. Electrons spend more time in the vicinity of chlorine. The chlorine end of $\text{HCl}(\text{g})$ is partially negative (delta negative δ^-). The hydrogen end is delta positive (delta positive δ^+).

II. Non Polar Covalent Compounds

- ☞ These are formed between two atoms which have combined and have nearly the same electronegativity. The electron density is shared equally e.g. in O_2

III. Simple Covalent Compounds

☞ These contain few atoms and have weak IMF.

IV. Giant Molecular Substances

☞ These contain hundreds of thousands of atoms. They have strong IMF e.g. Polymers like plastics, diamond and graphite.

6. IONIC BONDING/ELECTROVALENT BONDING

It is the electrical attraction between oppositely charged atoms which are produced when electrons are transferred from one atom to another during a chemical reaction.

WHAT FORMS THE IONIC BOND?

The attraction between oppositely charged ions

- ✎ As an example of ionic bonding, consider Sodium reacting with chlorine to form sodium chloride. Sodium has one valence electron. Chlorine has seven valence electrons. In a chemical reaction sodium donates its valence electron to chlorine.
- ✎ There is electron transfer in which sodium loses an electron and chlorine gains an electron.
- ✎ The sodium atom becomes a sodium cation with an electron configuration of Neon. The chlorine atom becomes a chloride ion with an electron configuration of Argon.

IONIC COMPOUNDS

- ✎ These are made up of electrically charged atoms or molecules as a result of gaining and losing of electrons.
- ✎ Ions of opposite charges form ionic compounds.
- ✎ This usually happens as a result of metals reacting with nonmetals.
- ✎ Examples of ionic compounds are sodium chloride, magnesium chloride and lithium iodide.

NOTE

- ✓ Ionic bonding happens between metals and non-metals
- ✓ Metals lose electrons and become cations.
- ✓ A CATION is a positively charged ion which is attracted to the cathode.
- ✓ Non metals gain electrons and become anions.
- ✓ Ionic Lattice is a giant structure of many ions arranged and packed in a regular manner.

PHYSICAL PROPERTIES OF IONIC COMPOUNDS

1. They are usually in solid form at room temperature
2. They are hard and brittle.
3. Ionic compounds are generally soluble in water but they are insoluble in organic solvents such as propane and benzene.
4. Ionic compounds are electrolytes. They conduct electricity in molten state or if they are dissolved in water.
5. Ionic compound have high melting and boiling points. Ions in these compounds form strong intermolecular forces to neighboring ions and a great amount of heat is needed to separate them.

7. METALLIC BONDING

- ☞ It is the electrical attraction between positive fixed ions and negative delocalised electrons in a metal. The metallic bond forms as gaseous metal cools to form a solid.

WHAT HAPPENS IN THE FORMATION OF METALLIC BONDS?

- ✂ In a gas state the electrons of a metal are localised. Meaning: they are only to be found inside electron shells. Thus they are not free to move.
- ✂ As metal cools to form a solid metal atoms take up fixed positions and the valence electrons are delocalised. Meaning: they are set free from the energy levels. On being delocalised they become more stable, reach a lower energy level and experience stronger forces of attraction.
- ✂ The metallic bond results from electrostatic forces of attraction between the fixed metal ions and the negative delocalised electrons. We say the fixed metal ions are like stationary islands embedded in a sea of electrons.

ELECTRONEGATIVITY

- ✓ It is a tendency of an atom to attract a shared pair of electrons towards itself.
- ✓ Electronegativity is a measure of how strong atoms attract bonding electrons to themselves.

EXERCISE 1

1. Work out the formula of Aluminium oxide
2. ii. Explain the reaction between calcium and chlorine to form calcium chloride.
3. What is the charge of chlorine in the compound calcium chloride?
4. Why are negatively charged ions called anions?
5. What is an anode?
6. Draw the dot and cross diagrams for the following reactions and suggest the molecular formulae of the compounds formed;
 - a. Lithium reacting with fluorine
 - b. Magnesium reacting with chlorine

FACTORS AFFECTING THE ELECTRONEGATIVITY OF ELEMENTS

The electronegativity of elements depends on two factors namely;

- ✓ The number of protons in the nuclei
- ✓ The size of the atom

NUMBER OF PROTONS

- The greater the number of protons in the nucleus of an atom the greater is the attraction for the bonded pairs of electrons. This results into a greater electronegativity of an element.
- In other words, an increase in the number of protons increases the nuclear charge. This has a pulling effect on the outermost shell electrons. This strengthens the attracting ability of an atom to its electrons and so increases electronegativity.

SIZE OF AN ATOM OR ATOMIC RADIUS

- The smaller the atom, the less distance between the nucleus and bonded pair of electrons and the greater is the attraction between the nucleus and the bonded electrons. The electronegativity therefore increases.
- In other words, an increase in the atomic radius means that outermost shell electrons are placed farther away from the nucleus which holds onto the electrons. The effect is that electrons are pulled less strongly to the nucleus. This decreases electronegativity.

INTRERMOLECULAR FORCE

- ☞ These are forces of attraction existing between molecules in a substance.
- ☞ These are forces of attraction that hold molecules together in a substance.
- ☞ The strength of IMF is important because it affects physical properties of substances such as melting and boiling points as well as state at room temperature.

TYPES OF INTERMOLECULAR FORCE

Depending on the type of atoms present between the intermolecular forces, we have the following types of intermolecular forces

- ✓ Dipole-dipole attraction
- ✓ Hydrogen bonding
- ✓ Van der waals forces

DIPOLE-DIPOLE ATTRACTION

- ☞ These are the forces of attraction existing between polar molecules in a substance.
- ☞ The positively charged portion of one molecule is attracted to the negatively charged portion of another molecule. This is the common intermolecular force of attraction.
- ☞ An example of dipole-dipole attraction may be seen in carbonyl monoxide molecules where the carbon atom of one molecule is attracted to the oxygen atom of another molecule.

HYDROGEN BONDING

- ☞ This is the force of attraction that happens when a hydrogen atom of one molecule is attracted to a more electronegative atom of another molecule.
- ☞ This is the force of attraction between the hydrogen atom of one molecule and an unshared electron pair of fluorine, oxygen or nitrogen of neighbouring molecules.
- ☞ It is the strongest intermolecular force of attraction. In this force of attraction, the partially positive hydrogen in one molecule attracts the partially negative atom (i.e. fluorine, oxygen or nitrogen) of another molecule.

VAN DER WAAL'S FORCES

- ☞ This is the force of attraction between neutral molecules of a substance. It occurs in large non-polar molecules. Van der Waal's forces are for example found among the halogens and between different layers of graphite.
- ☞ Their effect increases with the mass of molecules but between two molecules, their effect is negligible.

EFFECTS OF INTRERMOLECULAR FORCES ON THE PHYSICAL PROPERTIES OF SUBSTANCES

- ✓ It affects the melting points of substances. Solid substances which have strong IMF require more heat energy to overcome their attraction so as to change them to liquid form.
- ✓ It affects the boiling points of substances. Substances which have strong IMF usually have high boiling points. More heat energy is required for their particles to completely separate and form a gas. They therefore melt at an increased temperature.
- ✓ It affects the viscosity of fluids. Viscosity is the resistance of a fluid to flow and it increases with the strength of intermolecular force of a fluid.
- ✓ It determines the state of a substance at room temperature. Substances which have very strong IMF are in solid form, those that have weak IMF are in liquid while those that have very weak IMF are in gaseous state at room temperature. For example, the presence of hydrogen bonds in water gives it a liquid state at room temperature. Without these forces of attraction, water would exist as a gaseous substance at room temperature.

ALLOTROPY

- ❖ It is the existence of an element in two or more physical forms at the same physical state.
- ❖ The different forms of the same element are called allotropes. Allotropes of an element have different arrangements of atoms hence they have different physical properties.
- ❖ Oxygen, carbon and sulphur are some of the elements that have allotropic properties.

ALLOTROPE OF OXYGEN

There are two allotropes of oxygen namely;

- ✓ Molecular oxygen or diatomic Oxygen (O_2)
- ✓ Ozone or triatomic Oxygen (O_3)

PROPERTIES OF MOLECULAR OXYGEN (O_2)

- It is a clear, colourless gas
- It has no smell
- It is only slightly soluble in water
- It is very reactive. It reacts with a great many substances to produce oxides and the reactions usually give out a lot of energy.

PROPERTIES OF OZONE (TRIATOMIC OXYGEN, O_3)

- It has a pale blue colour.
- It has a characteristic smell or odour
- It is very unstable and changes spontaneously to oxygen.
- It is a powerful oxidizing agent
- It absorbs ultraviolet (UV) light from the sun, thereby protecting the earth from the sun's UV radiation

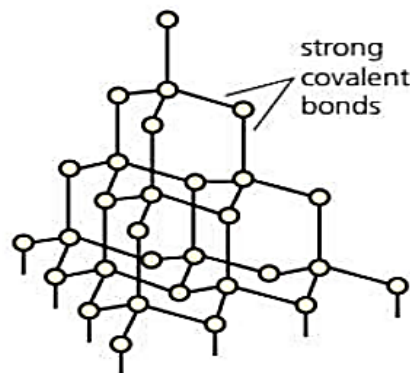
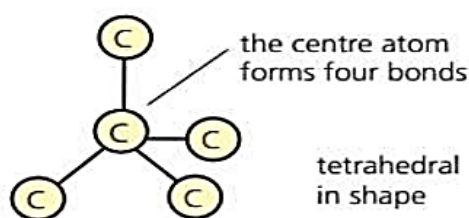
ALLOTROPE OF CARBON

Carbon is an element in group IV of the periodic table. A small amount of carbon occurs as the free element, in the earth's crust. There are two allotropes of carbon namely;

- ✓ Diamond
- ✓ Graphite

DIAMOND

- ✎ This allotrope is a clear, hard substance that sparkles in light.
- ✎ It is a giant structure of carbon atoms in which each carbon atom is covalently bonded to four other atoms.
- ✎ This bonding style results into a giant structure as shown in the diagram on the right.



The bonding arrangement of tetrahedral carbon atoms enables the diamond to be extremely hard with atoms very closely packed. Since all valence electrons are involved in bonding, diamond does not conduct electricity.

PHYSICAL PROPERTIES OF DIAMOND

- ✓ It is clear, but sparkles in the presence of light.
- ✓ It does not conduct electricity
- ✓ It has very high density
- ✓ It is insoluble in water
- ✓ It has very high melting and boiling points
- ✓ It is the *hardest known* and *the least compressible* substance.

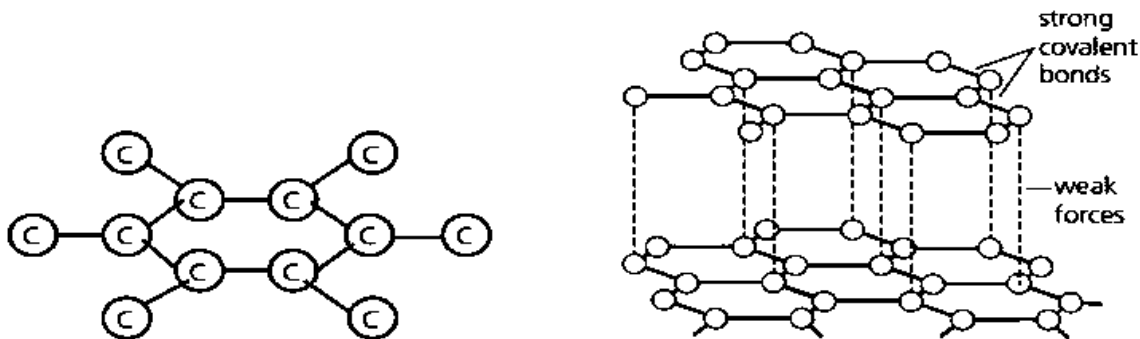
USES OF DIAMOND

- ✓ It is used for making heat sinks for cooling electronic devices. This so because it is a good heat absorber.
- ✓ It is used for making jewellery and ornaments.
- ✓ It is used for making drilling bits in derricks.
- ✓ It is used in making diamond studded saws.
- ✓ It is used in making glass cutters.

GRAPHITE

- ✎ Graphite is made up of plane layers of carbon atoms. Each carbon atom is bonded to three others using covalent bonds.

- ✎ This results into a structure with fused hexagonal rings arranged in an infinite three dimensional network as shown in the diagram on the right.
- ✎ The layers are attracted to each other by weak IMF called van der waals forces.



PHYSICAL PROPERTIES OF GRAPHITE

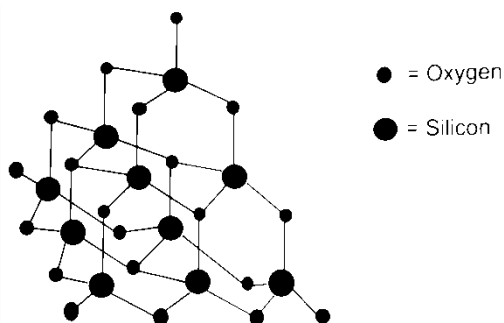
- ✓ It is dark grey in colour and opaque.
- ✓ It conducts electricity. Since the carbon atoms are bonded to three others, which leaves one electron free to move about
- ✓ It has a lower density than that of diamond.
- ✓ It is insoluble in water.
- ✓ It has a lower melting and boiling points than diamond.

USES OF GRAPHITE

- ✓ Graphite is used in making pencils.
- ✓ Graphite is used as a positive terminal in dry cells and as electrodes in industry.
- ✓ It is used for making lubricants. The weak IMF between the layers allows them to easily slide over each other.

COMPARISON BETWEEN DIAMOND AND SILICON (IV) OXIDE

- ✎ Silicon is an element in group 4 of the periodic table. It therefore has a valency of four and the atoms may bond tetrahedrally to other Si atoms just like in diamond.
- ✎ In its elemental form, silicon is a nonconductor of electricity, but when the temperature is increased, it becomes an excellent semi-conductor.
- ✎ In silicon oxide, silicon gets bonded to tetrahedrally to four oxygen atoms, resulting into a giant structure. An example of silicon oxide is common sand or silica.



Activity 3.6

AIM: Finding out physical properties of Silicon oxide

Materials: River

- Beaker
- Spatula
- Pestle
- Gas burner
- Crucible
- 100ml of water

Procedures:

1. Take some river sand. Put approximately 10g on spatula and heat over a burner flame.
2. Dissolve a spatula full of sand in 100ml of water in a beaker. What happens?
3. Crush some sand in a crucible using a pestle. How easily does this happen?

Results/observations:

What happens when you heat river sand? Does it melt? Does the sand really dissolve in water? How easily would you crush the sand?

You will have noticed from Activity 3.6 that sand (silicon oxide)

- ✓ Quite hard; it cannot easily be crushed
- ✓ Cannot dissolve in water
- ✓ Has high melting point (1160°C)

Silicon oxide forms a giant molecular structure, which is rigid and hard. The structure brings about strong covalent bonding between atoms, which cannot easily be broken.

The following table summarises some similarities between silicon oxide and diamond

PROPERTY	DIAMOND	SILICON OXIDE
Bonding arrangement	tetrahedral	tetrahedral
Packing of atoms	Very close	Very close
hardness	Very close	Very close
Density (g/cm^3)	3.51	3.25
Electrical conductivity	Does not conduct as there are no free electrons	Does not conduct as there are no free electrons
Melting point	high	high
Boiling point	high	high

PROPERTIES OF METALS

1. They have high density
2. Their melting and boiling points are high because the metallic bond is very strong.
3. They are good conductors of electricity and heat
4. They have high tensile strength because the metallic bond is flexible.
5. Metals do not dissolve in water.
6. Some metals are sonorous i.e. they produce a ringing sound.
7. Most metals are ductile i.e. they can easily be drawn into wires.
8. Most metals are lustrous i.e. they have shining appearances e. g. silver, gold.
9. Most of them are malleable i.e. they can be hammered into any shape.
10. Metals are in solid form at room temperature, the only exception is mercury which is in liquid form.

SPECIAL METALS

1. Mercury is a liquid at room temperature.
2. Iron is magnetic.
3. Alkali metals are soft and have no sonorous property.

ALLOYS

- An alloy is a mixture of two or more elements whereby atleast one of them is a metal.
- An alloy is mixture of two or more metal and a non-metal.
- Alloys are made so as to achieve the desire combination of properties of elements.
- An alloy is obtained by mixing the base metal with a given element or element so as to combine the desired qualities of strength, malleability weight or hardness among others.
- Generally, it is found that alloying produces a metallic substance which has more useful properties than the original pure metals it was made from.

EXAMPLES OF ALLOYS

The most commonly encountered examples of alloys are;

- (a) Steel
- (b) Stainless steel
- (c) Brass
- (d) Bronze
- (e) Duralumin (Duraluminium)

WHY ARE METALS MXED INTO ALLOYS?

- Because these mixtures tend to have better qualities than original elements
- For example an alloy of gold would have better properties in terms of strength and lustre than unmixed gold.

SOME ALLOYS AND THEIR USES

a. *Steel*

- ✍ This is an alloy of iron and carbon. It is very strong.
- ✍ The added carbon helps to prevent the iron from rusting and make it stronger.
- ✍ It is widely used in making:
 - ✓ Screws
 - ✓ and beams of
 - ✓ Bridges.
 - ✓ nails
 - ✓ building
 - ✓ Cutting tools

b. *Stainless steel*

- ✍ This is an alloy of iron, nickel and chromium in addition to carbon.
- ✍ These added elements keep the metal shiny and increase its resistance to corrosion.
- ✍ It is used for making
 - ✓ Sinks
 - ✓ Cutlery.
 - ✓ kitchen
 - ✓ surgical instruments

c. *Brass*

- ✍ This is an alloy of copper and zinc it does not corrode
- ✍ it is easy to work with and is very sonorous
- ✍ it is used for making:

- ✓ Propellers of ships
- ✓ bolts and nuts
- ✓ Electrical plugs
- ✓ screws
- ✓ door hinges

d. Bronze

✎ It is alloy of copper and tin. It does not corrode and very sonorous.

- ✓ It is used for making coins
- ✓ medals
- ✓ machinery
- ✓ statues
- ✓ Decoration work
- ✓ Springs

e. Duralumin

✎ This is an alloy of Aluminium, copper and small amounts of other elements. On its own, Aluminium is not very strong. For this reason manufactures mix Aluminium with Copper and other elements in order to strengthen it.

✎ It is used in making:

- ✓ Food, cans
- ✓ Foils
- ✓ Bicycle parts
- ✓ Rockets
- ✓ Parts of airplanes
- ✓ Kitchen utensils

UNIT 4

THE STRUCTURE OF AN ATOM

- ✎ An atom has a centre called NUCLEUS.
- ✎ The nucleus contains two types of particles namely PROTONS and NEUTRONS.
- ✎ Protons and neutrons are collectively called NUCLEONS.
- ✎ Around the nucleus electrons rotate at very high speeds.
- ✎ The paths electrons follow are called shells, orbitals or energy levels.
- ✎ They are called energy levels because they are associated with a certain form of energy called kinetic energy.
- ✎ Protons carry a positive charge of +1 and have a mass in atomic mass unit (amu) of 1.
- ✎ Neutrons have no charge but have a mass of 1 amu. Electrons carry a negative charge.
- ✎ The mass of electrons is so negligible (insignificant) that it is regarded as zero.
- ✎ An atom is electrically neutral i.e. it has an equal number of protons and electrons. The positive charges and negative charges balance and cancel each other out.
- ✎ The electrons are at a distance from the nucleus. The radius of the atom is more than 10000 times larger than the nucleus. The atom is mostly empty space.

EXERCISE 2

1. Name the part of the atom which is heaviest?

It is the Nucleus

2. give a reason:

It contains particles (protons and neutrons) which have mass.

3. What is the overall charge of the nucleus?

Positive

4. *Why?*

Because of the presence of the positively charged protons

5. *Why does the nucleus attract the energy levels?*

Because the nucleus and the energy levels are oppositely charged

ATOMIC NUMBER

Atomic number of an element is given by the number of protons in the nucleus of an atom.

SYMBOL for Atomic Number: Z

- ✓ Atomic number gives an atom its identity.
- ✓ It is very unique for a particular element.
- ✓ It determines the electron configuration which decides chemical properties of an element. (Electron configuration is discussed under the periodic table). The element carbon is carbon because it has 6 protons.

ATOMIC MASS/ MASS NUMBER/ NUCLEON NUMBER

☞ It is the total number of protons and neutrons in the nucleus of an atom.

SYMBOL: A

Number of Neutrons = Mass Number (A) – Atomic number (Z)

ISOTOPES

Definition of Isotopes: Isotopes are atoms of the same element that have the same atomic number but different mass numbers because of differences in the number of neutrons in their nuclei.

Why do isotopes have similar chemical properties?

They have the same electron configuration.

Examples of isotopes

- Isotopes of carbon are C-12 (with 6 protons, 6 electrons and 6 neutrons), C-13 (with 6 protons, 6 electrons and 7 neutrons) and C-14 (with 6 protons 6 electrons and 8 neutrons).
- Isotopes of hydrogen are H-1 called a proton (with 1 proton, 1 electron and 0 neutrons), H-2 called deuterium (with 1 proton, 1 electron and 1 neutron) and H-3 called tritium (with 1 proton, 1 electron and 2 neutrons)

WORKING OUT THE AVERAGE MASS OF AN ATOM

- *The average mass of an atom is the average mass of the isotopes of that atom.*

Example

Copper has two main isotopes Cu-63 and Cu-65. In nature they occur in the proportions 69% and 31% respectively. Work out the average mass of the isotopes of copper.

Working:

(Mass of isotope 1 X its percentage) + (mass of isotope 2 X its percentage)
divided by 100

$$\text{Average mass of copper} = \frac{63 \times 69 + 65 \times 31}{100}$$

$$\text{Average mass of copper} = \frac{4347 + 2015}{100}$$

$$\text{Average mass of copper} = \frac{6362}{100}$$

$$= 63.62 \text{ ans}$$

EXERCISE 3

1. The two main isotopes of carbon are C-12 and C-13. These are present in nature in the proportions 98.89% and 1.11% respectively. Work out the average mass of the isotopes of carbon.
2. The two isotopes of Bromine are Br-79 and Br-81. In nature they occur in equal proportions. Work out the average mass of the isotopes of bromine
3. Chlorine has two isotopes, Cl -35 and Cl-37. For every 4 chlorine atoms three are Cl-35 and one is Cl -37. Work out the average mass of the isotopes of chlorine.

UNIT 5

STOICHIOMETRY

- Stoichiometry is the study of the amounts of reactants and products in a chemical reaction. This branch of chemistry is based on the law of conservation of mass.

CHEMICAL REACTION

- Chemical reaction refers to rearrangement of atoms to form a new substance.
- For example, carbon reacts with oxygen to form carbon dioxide.

CHEMICAL EQUATION

- A chemical equation is the short and accurate representation of a chemical reaction.
- It has reactants on the left and products on the right separated by an arrow.

WRITING A BALANCED CHEMICAL EQUATION

- ☞ A good chemical equation must be balanced. A balanced chemical equation should show the same number of atoms of each kind that make up reactants on one side and products on the other side.
- ☞ The kinds of atoms involved in a chemical reaction do not increase or decrease after the chemical reaction, but remain the same.
- ☞ There are two methods of balancing chemical equations: **trial and error** method and the **systematic method**.

BALANCING CHEMICAL EQUATIONS USING THE TRIAL AND ERROR METHOD

- Write a correct symbol or formula for each reactant and products.
- Balance the number of atoms on both sides by putting a coefficient against the molecules. Do not change the subscripts of the atoms in the formula.
- Write the physical state after each reactant and product in brackets. Use the symbol (s) for solid, (l) for liquid, (g) for gas, and (aq) for aqueous solution.

Example:

Balance the equation; $\text{Na(s)} + \text{H}_2\text{O(l)} \rightarrow \text{NaOH(aq)} + \text{H}_2(\text{g})$

- ☞ It is observed from the equation there is a total of 2 hydrogen atoms in reactant 's side and 3 hydrogen atoms in the product 's side.
- ☞ If a coefficient 2 is put in front of NaOH the total number of hydrogen atoms becomes 4 in product 's side. To balance off, a coefficient 2 is put on both Na and H₂O.

☞ The balanced equation becomes; $2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2\text{(g)}$

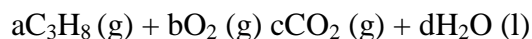
BALANCING CHEMICAL EQUATIONS USING THE SYSTEMATIC METHOD

Example:

Balance the equation $\text{C}_3\text{H}_8\text{(g)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)}$

Solution

Let the balanced equation be



Using C atoms only

$$3a = c \dots\dots\dots[1]$$

Using H atoms only

$$8a = 2d \dots\dots\dots[2]$$

Using O atoms only

$$2b = 2c + d \dots\dots\dots[3]$$

Solving the equations

Let $a = 1$. Then use this assumption to solve for all the equations using substitution method.

$$c = 3a = 3(1) = 3$$

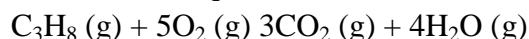
$$2d = 8a = 8(1)$$

$$d = 8/2 = 4$$

$$2b = 2(3) + 4$$

$$b = 10/2 = 5$$

The balanced equation is therefore



EXERCISE 4

Balance the following chemical equations;

1. $\text{Ca(s)} + \text{O}_2\text{(g)} \rightarrow \text{CaO(s)}$
2. $\text{CO(g)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$
3. $\text{Fe(s)} + \text{O}_2\text{(g)} \rightarrow \text{Fe}_2\text{O}_3\text{(s)}$
4. $\text{N}_2\text{(g)} + \text{H}_2\text{(g)} \rightarrow \text{NH}_3\text{(g)}$
5. $\text{N}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow \text{NO(g)}$
6. $\text{PCl}_5\text{(s)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{PO}_4\text{(aq)} + \text{HCl(aq)}$

THE LAW OF CONSERVATION OF MASS

- This law states that matter can change form, but can neither be created nor destroyed. The mass of substances in a chemical reaction remains the same. In any chemical process in a closed system, the mass of the reactants must be equal to the mass of the products.

RELATIVE FORMULA MASS FOR COMPOUNDS

- Relative formula mass for (RFM) is the mass of a molecule relative to the mass of carbon-12 isotope. Hence to work out the RFM of a compound, you should write the correct chemical formula, first you add up the relative atomic mass (RAM) of the elements that make up the compound to get the relative formula.

CALCULATING RELATIVE FORMULA MASS OF COMPOUNDS

To calculate the relative formula mass of a compound, add the relative atomic masses for all the atoms present in that compound.

Example 1

Work out the relative formula mass (R.F.M) of ethanol ($\text{CH}_3\text{CH}_2\text{OH}$).

(RAM of C = 12, H = 1 and O = 16).

Solution

$$\begin{aligned}\text{R.F.M of CH}_3\text{CH}_2\text{OH} &= (12 \times 2) + (1 \times 6) + (16 \times 1) \\ &= 24 + 6 + 16 \\ &= 46 \text{ amu}\end{aligned}$$

Example 2

Work out the relative formula mass (R.F.M) of aluminium sulphate, $\text{Al}_2(\text{SO}_4)_3$.

(RAM of Al = 27, S = 32, and O = 16)

Solution

$$\begin{aligned}\text{RFM of Al}_2(\text{SO}_4)_3 &= (27 \times 2) + (32 \times 3) + (16 \times 4 \times 3) \\ &= 54 + 96 + 192 \\ &= 342 \text{ amu}\end{aligned}$$

Table 5 should help you in working out the relative formula mass.

Table 5: symbols, atomic numbers and relative atomic numbers of some elements.

element	symbol	Atomic number	RAM	element	symbol	Atomic number	RAM
Aluminium	Al	13	27	Nitrogen	N	7	14
Argon	Ar	18	40	Oxygen	O	8	16
Beryllium	Be	5	9	Phosphorus	P	15	31
Boron	B	6	11	Potassium	K	19	39
Calcium	Ca	20	40	Silicon	Si	14	29
Chlorine	Cl	17	35.5	Sodium	Na	11	23
Carbon	C	6	12	Sulphur	S	16	32
Fluorine	F	9	19	Iron	Fe	26	56
Helium	He	2	4	Copper	Cu	29	64
Hydrogen	H	1	1	Zinc	Zn	30	65
Lithium	Li	3	7	Silver	Ag	47	108
Magnesium	Mg	12	24	Manganese	Mn	25	55
Neon	Ne	10	20	Chromium	Cr	24	52

THE MOLE

- A mole refers to the amount of substance which contains 6.023×10^{23} particles.
- The mole is a unit of measurement just like “pair”, “dozen”, “ream” among others.
- The number, 6.023×10^{23} is known as the Avogadro’s number or constant.

RELATIONSHIP BETWEEN MOLE AND MASS OF A SUBSTANCE

- 1 mole of the particles of any substance is equivalent to the relative atomic mass (RAM) of the atom in the molecule of the substance expressed in grams.
- For example, the relative atomic mass of sodium (Na) is 23 amu. Therefore, 1 mole of sodium atoms weighs 23g. Similarly, 23g of sodium atoms contains 6.023×10^{23} sodium atoms.

MOLAR MASS

- Molar mass is the mass of one mole of a chemical substance.
- If we weigh one mole of a substance, the mass obtained is referred to as molar mass.
- The unit of molar mass is the gram per mol. It is denoted as g/mol or gmol^{-1} .
- Molar mass of any element always contain 6.023×10^{23} particles although the molar masses of the elements are different

CALCULATING THE NUMBER OF PARTICLES IN A GIVEN MOLES OF AN ELEMENT OR COMPOUND

- When carrying out calculations involving number of particles, always bear in mind that one mole of a substance contains Avogadro's constant i.e. $L = 6.023 \times 10^{23}$ particles.
- The particles can be atoms, electrons, molecules etc. Since $1 \text{ mole} = 6.023 \times 10^{23} = \text{RFM (RMM, RAM)}$ in grams, there are several formulae that can be derived from this relationship. Use the following formula;

Formula: Number of moles = $\frac{\text{Mass in grams}}{\text{RAM or RFM}}$

EXERCISE 5 “Question three has been worked out as an example!”

- Work out the mass of gold (Au) which contains the same number of atoms as 3g of carbon.
- Work out the number of moles in 10g of sulphur dioxide.
- What mass of CuSO_4 contains 0.2 moles?

Solution: No of moles = $\frac{\text{Mass}}{\text{RFM}}$

$$\text{Mass} = \text{RFM} \times \text{No of moles}$$

$$\text{Mass} = 160 \times 0.2 = 32\text{g}$$

- Work out the percentage of water in hydrated sodium sulphate.

MOLAR VOLUME OF GASES

- It is the volume occupied by one mole of a gas.
 - Its value is 22.4dm^3 at standard temperature and pressure (s.t.p). Standard temperature is taken as 0°C while standard pressure is taken as 1 atm or 760mmHg.
- Its value is 24dm^3 at room temperature and pressure (r.t.p). Room temperature is taken as 25°C while room pressure is taken as 1 atm or 760mmHg.

EXERCISE 6

Consider ethane burning completely in air.

1. Work out a balanced chemical equation
2. Work out the volume of CO_2 evolved at STP.
3. Work out the volume of CO_2 evolved when 70g of ethane are used.

PERCENTAGE OF WATER IN HYDRATED IONIC AND MOLECULAR COMPOUNDS

- Crystals of some compounds, such as sugar and copper sulphate contain water which becomes part of the substance.
- The water that is incorporated in these ionic compounds is called the water of crystallization. Salts with water of crystallization are known as hydrated salts.

Hydrated ionic compounds have different number of water molecules within them. Table 1 shows examples of hydrated compounds.

Name	Formula	Number of water molecules
Magnesium sulphate-7-water	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$	7
Sodium sulphate-10-water	$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$	10
Cobalt chloride-6-water	$\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$	6
Calcium sulphate-2-water	$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$	2
Copper sulphate-5-water	$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	5

- When hydrated salts are heated, they lose their water of crystallization. It is possible to determine the mass and percentage of water of crystallisation in hydrated ionic compounds.

NOTE:

- When we completely heat copper sulphate it turns into a white powder. This powder is called anhydrous copper sulphate. The change in colour comes about because of the loss of water from hydrated copper sulphate.
- The molecules of sugar are made up of carbon, hydrogen and oxygen atoms as in the formula $\text{C}_{12}\text{H}_{22}\text{O}_{11}$.
- When sugar is heated, it disintegrates; carbon and hydrogen atoms combine to form water. Carbon remains in the test tube as the black residues
- The equation for the disintegration of the sugar after heating is as follows;
$$\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s}) \rightarrow 12\text{C}(\text{s}) + 11\text{H}_2\text{O}(\text{g})$$

EMPIRICAL FORMULA

- It is the simplest formula that shows the atoms that make up the compound in their lowest ratio.
- It is a chemical formula showing the simplest ratio of elements in a compound. For example, CH_2O is the empirical formula for glucose.

Working out empirical formula given grams of elements making up a compound

- ✓ Work out the number of moles.
- ✓ Work out the lowest mole ratio.

Example 1

A compound contains 52.2% carbon (C), 13.0% hydrogen (H), and 34.8% oxygen (O). Work out its empirical formula. (RAM of C = 12, H = 1 and O = 16)

Solution

Since percentage is usually out 100, then assume you have 100 g of the substance. Then convert the percentage of each element into mass in grams. Thus

Carbon = 52.2 g, H = 13.0 g and 34.8 g.

Element	Mass (g)	Number of moles	Simplest mole ratio
Carbon (C)	52.2	$\frac{52.2}{12} = 4.35$	$\frac{4.35}{2.175} = 2$
Hydrogen (H)	13.0	$\frac{13.0}{1} = 13$	$\frac{13}{2.175} = 6$
Oxygen (O)	34.8	$\frac{34.8}{16} = 2.175$	$\frac{2.175}{2.175} = 1$

The empirical formula of the compound is: **C₂H₆O**

Example 2

A compound was found to contain 3.2 g of copper, 0.6 g of carbon and 2.4 g of oxygen. Find its empirical formula. (RAM of Cu = 64, C = 12 and O = 16)

Solution

Element	Mass (g)	Number of moles	Simplest mole ratio
Copper (Cu)	3.2	$\frac{3.2}{64} = 0.05$	$\frac{0.05}{0.05} = 1$
Carbon (C)	0.6	$\frac{0.6}{12} = 0.05$	$\frac{0.05}{0.05} = 1$
Oxygen (O)	2.4	$\frac{2.4}{16} = 0.15$	$\frac{0.15}{0.05} = 3$

The empirical formula of the compound is: **CuCO₃**

EXERCISE 7

- In an experiment 2.6g of chromium were in excess of Chlorine gas 7.93g of Chromium chloride were formed. Work out the correct formula of Chromium chloride and work out a balanced chemical equation for the reaction.*
- Calculate the empirical formula of an organic compound containing 48.0g of carbon, 12.0g of Hydrogen and 32.0g of Oxygen.*

MOLECULAR FORMULA

- It is a chemical formula that gives the total number of atoms of each element in each molecule of a substance.
- It is a chemical formula that shows the actual number of atoms in the various elements present in one molecule of molecule of that compound. For example C₆H₁₂O₆ is the molecular formula for glucose

CALCULATING MOLECULAR FORMULA

Use the following steps to come up with the molecular formula of a substance

- Identify or work out the empirical formula
- Identify the molar mass of the compound
- Calculate the empirical formula mass
- Find the empirical formula units using the formula:

$$\text{Empirical formula units} = \frac{\text{molar mass}}{\text{empirical formula mass}}$$

- Multiply the subscripts in the empirical formula. The result is the required molecular formula.

Example 3

Calculate the molecular formula of a compound if its empirical formula is CH_2O and has a molar mass of 180 g/mol. (RAM of C=12, H = 1 and O = 16).

Solution

Empirical formula	=	CH_2O
Empirical formula mass	=	$(12 \times 1) + (1 \times 2) + (16 \times 1)$
	=	30 g/mol
Molar mass of compound	=	180 g/mol
Empirical formula units	=	$\frac{180 \text{ g/mol}}{30 \text{ g/mol}}$
	=	6
Molecular formula	=	$6(\text{CH}_2\text{O})$
	=	$\text{C}_6\text{H}_{12}\text{O}_6$

CONCENTRATION OF A SOLUTION

- ❖ Concentration of a solution is the amount of solutes dissolved in a solvent.
- ❖ It is a measure of the number of solute particles in a certain volume of a solvent.

WAYS OF EXPRESSING THE CONCENTRATION OF A SOLUTION

The most common ways of expressing the concentration of a solution are|:

1. In moles per unit volume

$$\text{Concentration} = \frac{\text{Number of moles of solute}}{\text{Volume of solution}}$$

The units of concentration in this case are mol/dm³, mol/cm³ and mol/litre

2. In mass per unit volume

$$\text{Concentration} = \frac{\text{Mass of solute}}{\text{Volume of solution}}$$

The units of concentration in this case are g/dm³, g/cm³ and g/litre

3. As a percentage

$$\text{Concentration} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100\%$$

CALCULATING CONCENTRATION OF SOLUTIONS

Example 1

Calculate the concentration of a solution containing 4 moles of sodium hydroxide (NaOH) in 2 dm³ of water.

Solution

$$\begin{aligned}\text{Concentration} &= \frac{\text{Number of moles}}{\text{Volume of solution}} \\ &= \frac{4 \text{ moles}}{2 \text{ dm}^3} \\ &= 2 \text{ mol/dm}^3\end{aligned}$$

Example 2

Calculate the concentration of sodium carbonate (NaCO₃) solution containing 53 g of the salt dissolved in 2 dm³.

Solution

$$\begin{aligned}\text{Number of moles} &= \frac{\text{Mass}}{\text{Molar mass}} \\ &= \frac{53 \text{ g}}{106 \text{ g/mol}} \\ &= 0.5 \text{ mole}\end{aligned}$$

$$\begin{aligned}\text{Concentration} &= \frac{0.5 \text{ mole}}{2 \text{ dm}^3} \\ &= 0.25 \text{ mol/dm}^3\end{aligned}$$

Example 3

12 g of sodium chloride (NaCl) is dissolved in 100g of water. Calculate the concentration of the solution.

Solution

$$\begin{aligned}\text{Mass of solute} &= 12 \text{ g} \\ \text{Mass of solution} &= 12 \text{ g} + 100 \text{ g} \\ &= 112 \text{ g} \\ \text{Concentration} &= \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \\ &= \frac{12 \text{ g}}{112 \text{ g}} \times 100\% \\ &= 0.1\% \text{ NaCl}\end{aligned}$$

MOLARITY

Molarity is defined as the number of moles of a solute per unit volume of the solution. The unit of molarity is the mol per cubic decimeter. It is also called the molar, abbreviated (M).

FORMULAE

A. $\text{MOLARITY} = \frac{\text{No of moles}}{\text{volume}}$

B. $\text{No of moles} = \frac{\text{mass}}{\text{RFM}}$

C. By substitution $\text{Molarity} = \text{No of moles} = \frac{\text{mass}}{\text{RFM} \times \text{Volume}}$

D. $\text{Mass} = \text{Molarity} \times \text{RFM} \times \text{volume}$

This is rewritten as $\frac{\text{mass}}{\text{Volume}} = \text{RFM} \times \text{Volume}$

But $\frac{\text{mass}}{\text{Volume}}$ gives us concentration.

So Concentration = Molarity \times RFM

Concentration is mass of solute particles dissolved in a given volume of a solution.

1 litre = 1dm³ = 1000cm³ = 1000ml. 1 litre of water has a mass of 1kg.

B. $C_1V_1 = C_2V_2$. where C is concentration and V is volume.

PROBLEM SOLVING ON MOLARITY

Example 1

Calculate the molarity of a solution of sodium hydroxide (NaOH), which was made by dissolving 10 g of solid sodium hydroxide in 250 cm³ of water. (RAM of Na = 23, O = 16 and H = 1)

Solution

$$\text{Number of moles} = \frac{\text{mass of NaOH}}{\text{molar mass of NaOH} \times \text{volume}}$$

$$\text{Volume} = \frac{250}{1000} \text{ dm}^3 = 0.25 \text{ dm}^3$$

$$\text{Molar mass of NaOH} = (23 \times 1) + (16 \times 1) + (1 \times 1) \text{ g/mol}$$

$$= 40 \text{ g/mol}$$

$$\text{Molarity} = \frac{10 \text{ g}}{40 \text{ g/mol} \times 0.25 \text{ dm}^3}$$

$$= 1 \text{ M}$$

Example 2

Calculate the mass of potassium hydroxide (KOH), which needs to be used to prepare 500 cm³ of a 2M solution in water. (RAM of K = 39, O = 16 and H = 1).

Solution

$$\text{Number of moles} = \text{concentration of KOH} \times \text{volume}$$

$$= 2 \text{ mol/dm}^3 \times \frac{500}{1000} \text{ dm}^3$$

$$= 2 \text{ mol/dm}^3 \times 0.5 \text{ dm}^3$$

$$= 1 \text{ mole}$$

$$\begin{aligned}
 \text{Mass of KOH} &= \text{Number of moles} \times \text{molar mass} \\
 &= 1 \text{ mole} \times [(39 \times 1) + (16 \times 1) + (1 \times 1)] \text{ g/mol} \\
 &= 1 \text{ mole} \times 56 \text{ g/mol} \\
 &= 56 \text{ g}
 \end{aligned}$$

EXERCISE 8

- What is the mass of solute particles in 250cm^3 , 1 molar hydrochloric acid solution?
- Given 20g of sodium hydroxide (NaOH) pellets, in what volume of water can you dissolve it to have 2 molar solution?
- What is the molarity of a solution of sodium chloride labeled 50% (HINT: 50% means 50g dissolved in 100ml of water).
- Work out the molarity of NaCl solution prepared by dissolving 5.8g in 200cm^3
- What is the molarity of 21% HNO_3 ?
- 60cm^3 of a solution whose concentration is 15g/cm^3 was diluted with distilled water by raising its volume to 80cm^3 . Calculate the concentration of the new solution.
- Calculate the mass of solute particles when 100ml salt solution is removed from a stock of 200g/l
- Explain how you can make a 20g/l concentration of 500cm^3 sugar solution from a stock solution whose concentration is 50g/l.

STANDARD SOLUTIONS

DEFINITION: A standard solution is a solution whose concentration is known.

PREPARATION OF STANDARD SOLUTIONS

There are two methods of preparing a standard solution. These are;

- Dissolution method
- Stock dilution (dilution of another standard solution)

DISSOLUTION METHOD

- ✎ In this method we use measure out an exact mass of a solute and dissolve it in the required volume of a solvent.
- ✎ To prepare a standard solution using the dissolution method, do the following steps;
 - Weigh the solute that you want to dilute.
 - Add water to the solute in the beaker and stir the solution with the rod until all the solute is dissolved.
 - Transfer the solution into the appropriate volumetric flask.
 - Rinse the beaker and funnel with plenty of distilled water
 - Then add more distilled water up to the graduated mark

Example

Describe how you would prepare 250 ml of 2M potassium carbonate (K_2CO_3) solution using solid potassium carbonate. (RAM of K = 39, C = 12 and O = 16).

Solution

- Firstly, the mass of potassium carbonate to be dissolved is calculated using the formula:

$$\begin{aligned}\text{Mass of } K_2CO_3 &= \text{molarity} \times \text{molar mass} \times \text{volume} \\ &= 2M \times [(39 \times 2) + (12 \times 1) + (16 \times 3)] \text{g/mol} \times 0.25 \text{ dm}^3 \\ &= 2M \times 138 \text{ g/mol} \times 0.25 \text{ dm}^3 \\ &= 69 \text{ g}\end{aligned}$$

- 69g of solid potassium carbonate is weighed using a balance and placed in a beaker.
- Using distilled water, the solid potassium carbonate is dissolved quantitatively.
- Then, the potassium carbonate solution is transferred into a 250 ml volumetric flask. The beaker is rinsed with distilled water several times, adding the solution into the volumetric flask.
- Distilled water is added into the flask up to the mark.
- The flask is covered on top and shaken to mix thoroughly. The concentration of the solution will be 2M potassium carbonate.

STOCK DILUTION (DILUTION OF ANOTHER STANDARD SOLUTION)

☞ Dilution refers to the process of adding additional solvent to a solution to decrease its concentration. This process keeps the number of solutes constant but increases the total volume of a solution.

☞ The volume of a solvent needed to prepare the desired concentration of a new, diluted solution can be calculated mathematically from the following relationship; $M_1 V_1 = M_2 V_2$ where,

M_1 = concentration of the original (stock) solution

V_1 = volume of the original (stock) solution

M_2 = concentration of the diluted solution

V_2 = volume of the diluted solution

How to prepare a standard solution by stock dilution?

- ✓ Firstly, you should know the volume and concentration of the stock solution
- ✓ Secondly, you should determine the concentration and total volume of the new solution you want to make.
- ✓ Thirdly, you add the correct volume of the stock solution to the correct volume of water so that the total volume is what you want.

Example 1

60 cm³ of the solution whose concentration is 15 g/cm³ was diluted with distilled water by raising its volume to 80 cm³. Calculate the concentration of the new solution.

SOLUTION

By dilution formula:

$$\begin{aligned}
 C_1 V_1 &= C_2 V_2 \\
 15 \text{ g/cm}^3 \times 60 \text{ cm}^3 &= C_2 \times 80 \text{ cm}^3 \\
 C_2 &= \frac{15 \text{ g/cm}^3 \times 60 \text{ cm}^3}{80 \text{ cm}^3} \\
 &= 11.25 \text{ g/cm}^3
 \end{aligned}$$

Example 2

Calculate the amount of water that must be added to 5 cm³ of 2M hydrochloric acid to dilute it to 0.1M.

Solution

By dilution formula

$$\begin{aligned}
 C_1 V_1 &= C_2 V_2 \\
 2\text{M} \times 5 \text{ cm}^3 &= 0.1 \text{ M} \times V_2 \\
 V_2 &= \frac{2\text{M} \times 5 \text{ cm}^3}{0.1\text{M}} \\
 &= 100 \text{ cm}^3 \\
 \text{Water added} &= V_2 - V_1 \\
 &= 100 \text{ cm}^3 - 5 \text{ cm}^3 \\
 &= 95 \text{ cm}^3
 \end{aligned}$$

DETERMINING THE CONCENTRATION OF A SOLUTION BY TITRATION

- Titration is the gradual addition of one solution to another until the end point is reached.
- Titration is the gradual addition of one solution to another until the reaction between them is complete.
- The best known example of titration is neutralization of an acid by a base and redox reaction.

FACTS ABOUT ACID – BASE TITRATION

During the titration, one solution (titrant) is titrated against or added to another solution (analyte) until the reaction between the components in the solutions is complete.

Titrant: This is the solution of a known concentration which is added to another solution to determine the concentration of the unknown solution.

Analyte: It is the solution of unknown concentration.

End point: It is the balanced point whereby an acid is completely neutralized by a base. The end point is shown by an indicator.

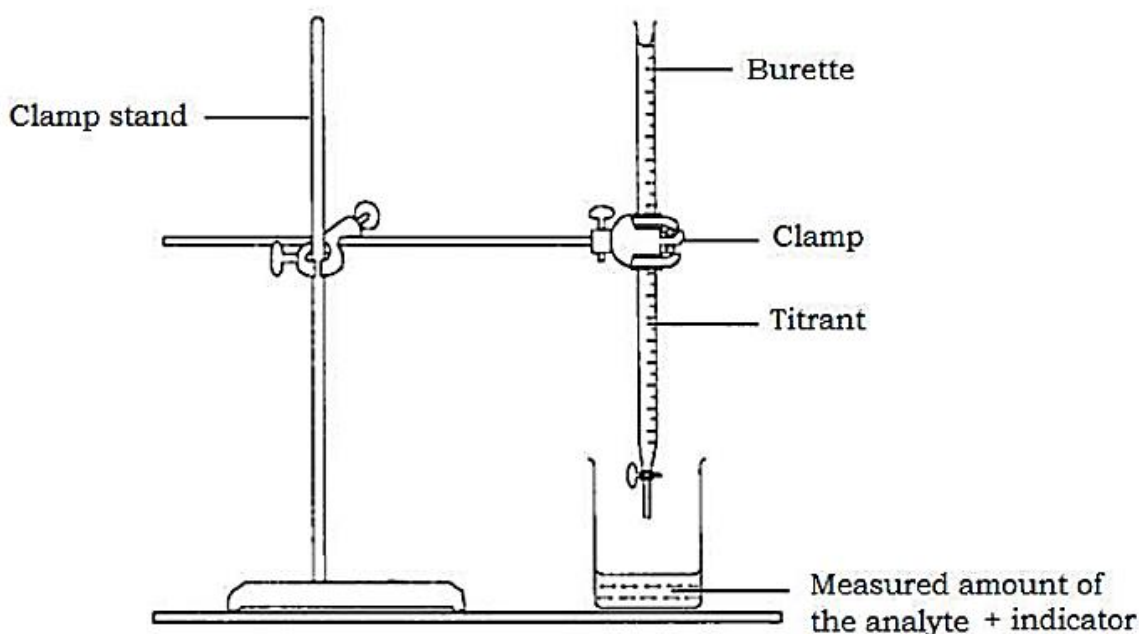
Indicator: An indicator is a dye that has a different colour in acidic or basic solutions. Examples of indicators are used during a titration are phenolphthalein and Methyl orange. When the end point is reached, there is a change in colour of the indicator.

APPARATUS USED IN A TITRATION

- | | | |
|--------------------------|----------------------|---------------------------|
| ✓ Clamp and clamp stands | ✓ Dropper bottle | ✓ Conical flask or beaker |
| ✓ Burette or syringe | ✓ Measuring cylinder | ✓ White tile |
| | ✓ Funnel | |

STEPS TO BE FOLLOWED WHEN CARRYING OUT A TITRATION

1. The apparatus is set up as shown below.



2. The titrant is added into the burette.
3. A specific volume of the solution to be titrated is poured into a conical flask.
4. The indicator is added into the conical flask.
5. The titrant is slowly added to the solution being titrated until there is a colour change.
6. The volume of the titrant added from the burette is recorded.

CALCULATING CONCENTRATION OF UNKNOWN SOLUTION

If the titrant and analyte have a mole ratio of 1:1 in the balanced equation, the concentration of the unknown solution is calculated using the formula:

$$C_a \times V_a = C_b \times V_b$$

where

C_a	=	concentration of the acid used
V_a	=	volume of the acid used (cm ³)
C_b	=	concentration of the alkali used
V_b	=	volume of the alkali used (cm ³)

On the other hand if the titrant and analyte have a mole ratio of greater than 1:1 in the balanced equation (e.g. 2:1), the concentration of the unknown solution is calculated using the formula:

$$\frac{C_a \times V_a}{N_a} = \frac{C_b \times V_b}{N_b}$$

where

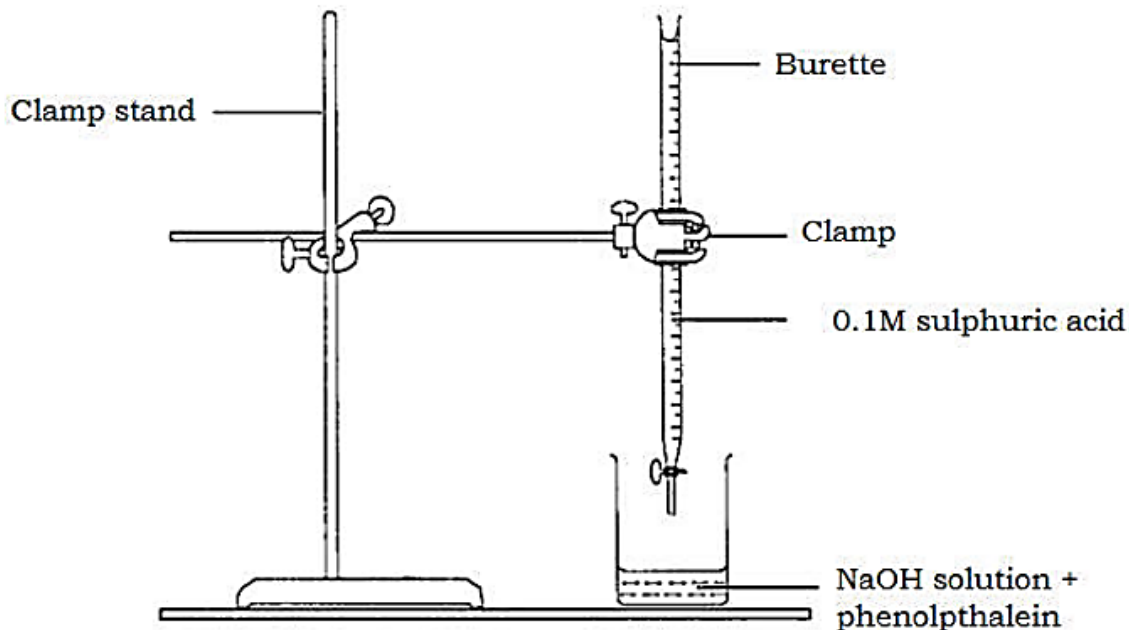
N_a	=	number of moles of the acid shown in the chemical equation
N_b	=	number of moles of the alkali shown in the chemical equation

PROBLEM SOLVING ON TITRATION

Example 1

You are provided with a burette, a funnel, a measuring cylinder, a beaker, clamp and clamp stand, 0.1M sulphuric acid (0.1M H_2SO_4), sodium hydroxide solution of unknown concentration, phenolphthalein indicator and a white tile or paper.

a. Set up the apparatus as shown below.



b. Fill the burette with sulphuric acid (H_2SO_4) up to the zero mark.

c. Pour 15 cm³ of sodium hydroxide into a beaker.

d. Add three drops of phenolphthalein indicator to the sodium hydroxide.

e. Slowly, add the sulphuric acid from the burette to sodium hydroxide, shaking the beaker all the time, until the pink colour disappears.

f. Note and record the volume of sulphuric acid used

Initial volume of sulphuric acid (cm ³)	Final volume of sulphuric acid (cm ³)	Volume of sulphuric acid used (cm ³)

g. Empty the beaker and rinse it with distilled water.

h. Repeat steps c to f.

i. Calculate the average volume of sulphuric acid used.

j. Write a balanced equation for the reaction.

k. Calculate the concentration of hydrochloric acid.

Example 2

In a titration, 11.6 cm³ of 3.0M of sulphuric acid (H₂SO₄) was required to neutralize 25 cm³ of sodium hydroxide (NaOH) of unknown concentration.

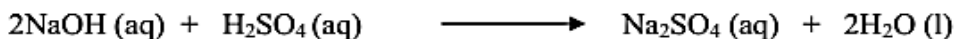
- a. Identify each of the following in the titration; titrant, analyte

Solution

The titrant is sulphuric acid. The analyte is sodium hydroxide.

- b. Write down a balanced equation for the reaction.

Solution



- c. Calculate the concentration of sodium hydroxide.

Solution

The titrant and analyte have a mole ratio of 2:1. Therefore concentration of sulphuric acid is calculated using the formula:

$$\begin{aligned}\frac{C_a \times V_a}{N_a} &= \frac{C_b \times V_b}{N_b} \\ C_b &= \frac{C_a \times V_a}{N_a} \times \frac{N_a}{V_b} \\ &= \frac{3.0\text{M} \times 11.6\text{cm}^3}{1} \times \frac{2}{25\text{cm}^3} \\ &= 2.784\text{M}\end{aligned}$$

The concentration of sodium hydroxide is **2.784M**.

THE LIMITING AND EXCESS REAGENTS

- Reactants must react in a given mole ratios in order for the reactants to be all used up.
- Sometimes in a chemical reaction, the amount of reactants available may not be in their recommended mole ratios according to stoichiometric equation.
- In this case, some amount of reactants will be left over once the reaction is completed. Also, the amount of products produced by such a reaction is expected to be equally restricted by the amount of reactant that has been completely used up in the reaction.

LIMITING REAGENT

- ❖ It is the reactant that restricts the amount of products that can be formed in a reaction.
- ❖ It is the reactant that is totally consumed when the chemical reaction is complete.
- ❖ The reaction cannot continue without it. Once it is finished, the reaction process stops.

EXCESS REAGENT

- ❖ It is the reactant in a chemical reaction that remains when a reaction is completed.
- ❖ The excess reagent remains because there is nothing which can react with it.

IDENTIFYING THE LIMITING AND EXCESS REAGENTS

There are **two** ways of finding the limiting reagent.

A. The first method is to calculate the mass of products that can be formed from the given quantities of reactants. The reactant that produces the smallest amount is the limiting reagent. The other one is therefore the excess reagent.

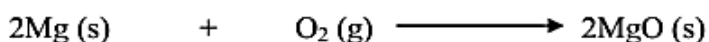
☞ To determine the limiting reagent and the excess reagent using this method, there are four important guidelines you should follow:

1. Ensure the chemical equation is balanced.
2. Convert given masses into moles.
3. Calculate the number of moles of product produced by each reactant.
4. The reactant that produces the smaller number of moles of product is the limiting reagent.

B. The second method is to divide the moles of each reactant with its stoichiometric coefficient. The one with the lowest ratio is the limiting reagent. The other one is therefore the excess reagent.

Example 1

Magnesium reacts with oxygen to form magnesium oxide according to the equation:



If 6g of magnesium reacts with 2 g of oxygen determine the:

- a. limiting reagent and excess reagent
- b. mass of magnesium oxide formed. (RAM of Mg = 24 and O = 16)

Solution

$$\begin{aligned} \text{a. Number of moles of magnesium} &= \frac{\text{mass of magnesium}}{\text{molar mass}} \\ &= \frac{6\text{ g}}{24\text{ g/mol}} \\ &= 0.25\text{ mole} \end{aligned}$$

$$\begin{aligned} \text{Number of moles of oxygen} &= \frac{\text{mass of oxygen}}{\text{molar mass}} \\ &= \frac{2\text{ g}}{32\text{ g/mol}} \\ &= 0.0625\text{ mole} \end{aligned}$$

To find the number of moles of product formed by each reactant, we use the balanced equation.

$$\begin{aligned} \text{i. From the balanced equation} \\ 2\text{ moles Mg} &= 2\text{ moles MgO} \end{aligned}$$

$$0.25 \text{ mole Mg} = 0.25 \text{ mole MgO}$$

ii. From the balanced equation

$$1 \text{ mole O}_2 = 2 \text{ moles MgO}$$

$$0.0625 \text{ mole O}_2 = y$$

$$y = 0.0625 \times 2$$

$$= 0.125 \text{ mole MgO}$$

Comparing the number of moles of product each reactant produces, it can be seen that O_2 gives out the smaller number.

Hence, O_2 is the limiting reagent while Mg is the excess reagent.

b. The quantity of the product formed is controlled by the limiting reagent.

Thus

$$0.0625 \text{ mole of O}_2 = 0.125 \text{ mole of MgO}$$

$$\text{Mass of MgO} = \text{Number of moles} \times \text{molar mass}$$

$$= 0.125 \text{ mole} \times 40 \text{ g/mol}$$

$$= 5 \text{ g}$$

YIELD AND PERCENTAGE YIELD IN A CHEMICAL REACTION

- ❖ The amount of product that is obtained by experiment is called actual yield.
- ❖ The amount of product obtained from a balanced chemical equation without doing the experiment is called theoretical yield.
- ❖ Actual yield and theoretical yield are often different.

The actual yield is often less than the theoretical yield due to the following factors;

1. Poor collection of products
 2. Presence of impurities in the reactants
 3. Poor experimental design
- ❖ To calculate the percentage yield, use the following equation;

$$\text{Percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Example 1:

2.34 g of aluminium reacts with excess copper (II) sulphate solution to produce 3.89 g of copper according to the equation:



Determine the theoretical, actual and percentage yield of copper.

(RAM of Cu = 64 and Al = 27)

Solution

- a. To find the theoretical yield

From the balanced equation

$$2 \text{ moles Al} = 3 \text{ moles Cu}$$

$$2 \times 27\text{g Al} = 3 \times 64\text{g Cu}$$

$$54\text{g Al} = 192\text{g Cu}$$

$$2.34\text{g Al} = y$$

$$y = \frac{2.34 \times 192}{54}$$

$$= 8.32\text{g}$$

The theoretical yield of copper is 8.32g;

- b. The actual yield of copper is 3.89g

$$\begin{aligned} \text{c. Percentage yield} &= \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 \\ &= \frac{3.89\text{g}}{8.32\text{g}} \times 100 \\ &= 46.75\% \end{aligned}$$

Exercise 9

- a. Use the following equation to answer questions 1-3: $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$

- How many moles of NH_3 can be produced from the reaction of 28 g of N_2 ?
- How many moles of NH_3 can be produced from the reaction of 25 g of H_2 ?
- If 28 g of N_2 and 25 g of H_2 are reacted together, which one would be the limiting reactant?

- b. At high temperatures, sulfur combines with iron to form the brown-black iron (II) sulfide: $\text{Fe (s)} + \text{S (l)} \rightarrow \text{FeS (s)}$

- In one experiment, 7.62 g of Fe are allowed to react with 8.67 g of S. a. What is the limiting reagent, and what is the reactant in excess?
- b. Calculate the mass of FeS formed.

- c. Work out the mass of CO_2 formed when 72g of methane burn completely in air.
- d. 10g of calcium metal reacts with dilute sulphuric acid which is in excess. CaSO_4 and hydrogen gas are the products.
 - i. Write down a balanced equation?
 - ii. Work out the number of moles of calcium metal used up?
 - iii. Work out the moles of sulphuric acid needed?
 - iv. Work out the volume of hydrogen gas evolved at STP?

UNIT 6

HEATS OF REACTIONS

- Heat is a form of energy measured in joules.
- Reactions are classified by how they behave in terms of heat changes.
- There are two types of chemical reactions from this view, exothermic and endothermic reactions.

ENTHALPY OF A SUBSTANCE

- This is the amount of heat energy contained in a substance.
- The symbol for enthalpy is H. All matter contains energy in form of kinetic energy and potential energy.
- ☞ Enthalpy of a substance can change due to various factors. Formation and breaking of bonds change of states of matter and dissolution of substance in water cause changes in enthalpy of the substance.
- ☞ Enthalpy change is denoted by the symbol ΔH . The change itself is denoted by Δ .
- ☞ The heat change is calculated as the difference between the heat content of the products of the reaction and that of the reactants.

$$\text{Enthalpy change} = \text{Enthalpy of products} - \text{Enthalpy of reactants}$$

$$\Delta H = H_2 - H_1$$

- Chemical reactions that cause a change in temperature of the surrounding are called **thermal-chemical** reactions. An example of such reactions is combustion.

EXOTHERMIC REACTION

- This is the type of reaction that gives out heat energy to the surroundings.
- It causes the container to heat up when the heat is given to the surroundings.
- In this type of reaction the energy content of the reactants is more than the energy content of the products.
- When the reaction gives heat to the surrounding it means that there is an energy loss and because of this, the products formed have lower energy content than the reactants.

- When the energy absorbed during bond breaking is less than the energy given out during bond making, the excess energy is released in form of heat.
- The overall reaction is thus exothermic and the heat change ΔH is indicated by a negative sign.

Heat energy diagram profile for exothermic reaction

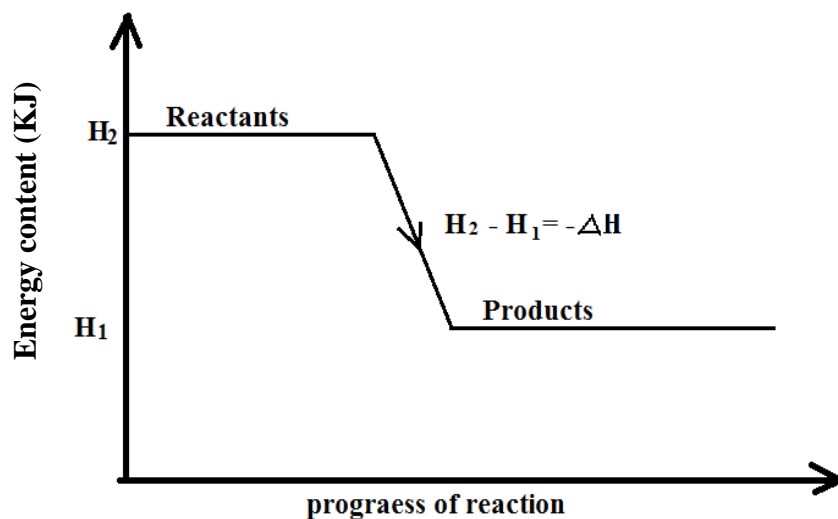
When heat energy of reactants is subtracted from heat energy of products, the answer is negative. When the temperature of the reactants is subtracted from the temperature of the products, the answer is positive i.e. $\Delta T = T_p - T_r$ gives a positive answer:

ΔT = change in temperature

T_p = temperature of the products

T_r = temperature of the reactants

- For an exothermic reaction, the reactants are at higher energy level than the products as shown below;



Examples of exothermic reactions

The following are some of the examples of exothermic reactions;

- ✓ Mixing of concentrated sulphuric acid with water.
- ✓ A reaction of a strong acid and a base e.g. $\text{NaOH}_{(aq)} + \text{HCl}_{(aq)} \rightarrow \text{NaCl}_{(aq)} + \text{H}_2\text{O}_{(l)}$
- ✓ Most combustion processes.

ENDOTHERMIC REACTIONS

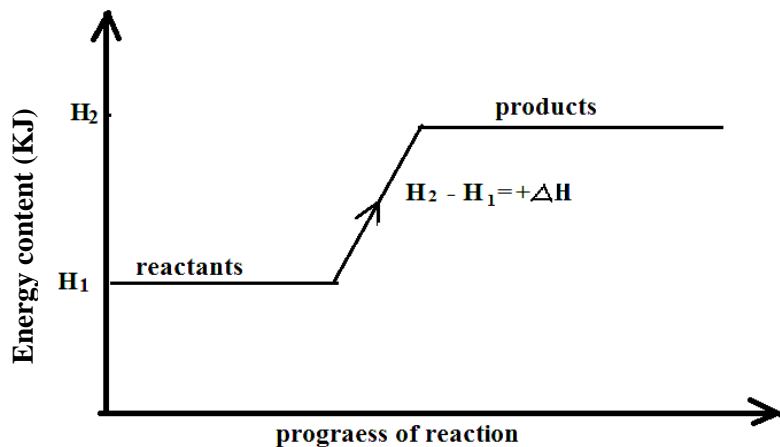
- This is a reaction that takes heat in from the surrounding.
- In Endothermic reactions, the products have more energy than reactants.
- In Endothermic reactions, energy taken in to break the bonds is greater than the energy given out when new bonds are formed.
- The container where the reaction is happening feels cold.
- When heat energy from the reactants is subtracted from the heat energy of the products, the answer is positive i.e. $\Delta H = H_2 - H_1 = +\Delta H$
- In term of temperature, when the temperature of the reactants is subtracted from the temperature of the products, the answer is negative. $\Delta T = T_p - T_r = \text{negative}$.

Note: When $\Delta H = \text{Positive}$, $\Delta T = \text{negative}$

When ΔH = negative, ΔT = Positive

HEAT ENERGY DIAGRAM FOR ENDOTHERMIC REACTION

- For an endothermic reaction, the reactants are at a lower energy level than the products as shown below.



EXAMPLES OF ENDOTHERMIC REACTIONS

- ✓ The dissolving of ammonium nitrate in water
- ✓ Reaction of oxygen with nitrogen: $N_2(g) + O_2(g) \rightarrow 2NO_2(g)$
- ✓ The process of photosynthesis in green plants.

THERMAL-CHEMICAL REACTIONS

- These are chemical reactions that release or absorb heat to or from their surroundings thereby affecting the temperature. These reactions can be identified using their equations and graphs.

IDENTIFYING ENDOTHERMIC AND EXOTHERMIC REACTINS FROM THEIR THERMICAL EQUATIONS

Chemical equations of thermo-reactions can tell us whether the reaction is endothermic or exothermic. They are presented with heat energy either as one of the reactants or one the products.

THERMO-CHEMICAL EQUATION OF ENDOTHERMIC REACTIONS

- ✎ An endothermic reaction is easily identified because its chemical equation includes heat as one of the reactants. Since the reaction is absorbing heat from the surroundings, its equation will be; $\text{Reactants} + \text{Heat} \rightarrow \text{products}$
- ✎ The equation of an endothermic reaction between ammonium nitrate and water will therefore be, $NH_4NO_{3(s)} + H_2O_{(l)} + \text{Heat} \rightarrow NH_4^+_{(aq)} + NO_{3(aq)}$

THERMO-CHEMICAL EQUATIONS OF EXOTHERMIC REACTIONS

- An exothermic process releases heat to the surrounding. Its equation will therefore include heat as one of the products (ie $\text{Reactants} \rightarrow \text{Products} + \text{Heat}$)

- The equation of neutralization reaction between hydrochloric acid and sodium hydroxide solution will therefore be: $\text{NaOH}_{(\text{aq})} + \text{HCl}_{(\text{aq})} \rightarrow \text{NaCl}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} + \text{Heat}$

BOND ENERGY

- It is the amount of energy in kilojoules associated with breaking or forming of 1 mole of chemical bonds in an element or compound.
- Some bonds are stronger than others. Strong bonds require more energy to break.
- The amount of energy required to break a bond is the same as the amount of energy given off when the same bond is formed.

BOND BREAKING –ENDOTHERMIC PROCESS

- When a solid melts to form a liquid, it absorbs heat energy from its surroundings.
- When a liquid changes to a gas at a boiling point it absorbs more heat energy. Likewise, when a bond breaks to form individual atoms, it absorbs heat energy from the surroundings

BOND FORMATION-EXOTHERMIC PROCESS

- When water vapour is cooled, particles join together to form a liquid.
- When liquid water is cooled further it freezes to form ice. Both Condensation and freezing processes are exothermic. When atoms join together to form a bond, energy is also given off. So bond formation is an exothermic process.
- Just like condensation and freezing, it gives off heat to the surroundings.
- When atoms join together to form a bond, energy is also given off. So bond formation is an exothermic process, just like condensation and freezing, it gives off heat to the surroundings.

DETERMINING THE KIND OF REACTION USING BOND ENERGY

- We can determine whether a reaction is exothermic or endothermic from the value of the overall energy change. The overall energy change is worked out by subtracting energy given when bonds form from the energy absorbed when bonds break.

CALCULATING THE OVERALL ENERGY CHANGE USING BOND ENERGY

- Since exothermic change gives out energy while endothermic change takes in energy, the overall energy can easily be worked out using the data given below:

Bond	Energy(Kj/mol)	Bond	Energy(Kj/mol)
H-H	436	N-N	160
C-H	413	N=O	631
N-H	393	N≡N	941
O=O	498	N-O	201
C-C	347	C=O	805
Cl-Cl	242	O-H	464
C-Cl	397	H-Cl	433
C-O	358	O-Cl	269
C=C	607	O-O	204

Example

Methane burns in excess oxygen to produce carbon dioxide and water according to the equation:



- Calculate the overall enthalpy change for the reaction.
- State whether the reaction is endothermic or exothermic based on enthalpy change.

[Hint: use the bond energies data given in the table above]

Solution

- Breaking 4 C–H bonds in methane requires

$$4 \times 413 \text{ kJ} = 1652 \text{ kJ}$$

Breaking 2 O=O bonds in oxygen requires

$$2 \times 498 \text{ kJ} = 996 \text{ kJ}$$

$$\begin{aligned} \text{Total energy required to break bonds} &= (1652 + 996) \text{ kJ} \\ &= 2648 \text{ kJ} \end{aligned}$$

Making 2 C=O bonds in carbon dioxide gives out

$$2 \times 805 \text{ kJ} = 1610 \text{ kJ}$$

Making 4 O–H bonds in water gives out

$$4 \times 464 \text{ kJ} = 1856 \text{ kJ}$$

$$\begin{aligned} \text{Total energy given out} &= (1610 + 1856) \text{ kJ} \\ &= 3466 \text{ kJ} \end{aligned}$$

$$\text{Overall energy change} = 2648 - 3466$$

$$\Delta H = -818 \text{ kJ}$$

- Basing on enthalpy change the reaction is exothermic.

UNIT 7

ALKANOLS

- ☞ Alkanols are organic compounds containing Carbon, Hydrogen and Oxygen.
- ☞ They are therefore classified as oxycarbons. They are also known as **alcohols**.
- ☞ Alkanols form a homologous series with **–OH** as the functional group. They may be regarded as being derived from an alkane by replacing one of the hydrogen atoms with an –OH group. They are saturated organic compounds.

GENERAL FORMULA AND NOMENCLATURE

- ☞ General formula of alkanols is $C_NH_{2N+1}OH$ or **ROH**, where R is the alkyl group.
- ☞ An alkyl group is the name given to a group of atoms remaining when one of the hydrogen atoms of an alkane is removed.
- ☞ For example, when we remove one hydrogen atom from:
 - Methane (CH_3), we get methyl (CH_3) group.
 - Ethane (CH_3CH_3), we get ethyl (CH_3CH_2) group and
 - Propane ($CH_3CH_2CH_3$), we get propyl ($CH_3CH_2CH_2$) group.
- ☞ In this methyl, ethyl and propyl are examples of alkyl groups and can be represented by letter **R** as described above.
- ☞ To get the number of carbon and hydrogen atoms, of an alkanol molecule, the general formula used is: $C_NH_{2N+1}OH$, where N is 1, 2, 3, 4...and so on.

NAMING ALKANOLS

- ☞ The names of alkanols are obtained by replacing the suffix *–e* of the corresponding alkane by *–ol*. For example, when $N = 1$ in alkanes, the name is **methane**. In this example, if we replace the suffix *–e* in methane with *–ol*, the name becomes **methanol**. It is the first member of the family of alkanols.
- ☞ The IUPAC (International Union of Pure and Applied Chemistry) rules demand that the position of the *–OH* group be included in the name.
- ☞ This is why the number 1 is included in **propan-1-ol** to show that the functional group is attached to first carbon of the straight carbon chain.

MOLECULAR AND CONDENSED FORMULAE OF ALKANOLS

- ✂ A molecular formula is a formula shows the actual number of atoms of each element in a molecule. For alkanols, the functional group is also shown in the molecular formula.
- ✂ A Condensed formula is a formula which shows how atoms are attached to each other in the structure of the molecule.
- ✂ The table below shows the molecular and condensed formulae of the first ten alkanols.

Name of alkanol	Molecular formula	Condensed formula
Methanol	CH ₃ OH	CH ₃ OH
Ethanol	C ₂ H ₅ OH	CH ₃ CH ₂ OH
Propanol	C ₃ H ₇ OH	CH ₃ (CH ₂) ₂ OH
Butanol	C ₄ H ₉ OH	CH ₃ (CH ₂) ₃ OH
Pentanol	C ₅ H ₁₁ OH	CH ₃ (CH ₂) ₄ OH
Hexanol	C ₆ H ₁₃ OH	CH ₃ (CH ₂) ₅ OH
Heptanol	C ₇ H ₁₅ OH	CH ₃ (CH ₂) ₆ OH
Octanol	C ₈ H ₁₇ OH	CH ₃ (CH ₂) ₇ OH
Nonanol	C ₉ H ₁₉ OH	CH ₃ (CH ₂) ₈ OH
Decanol	C ₁₀ H ₂₁ OH	CH ₃ (CH ₂) ₉ OH

STRUCTURAL FORMULAE OF ALKANOLS

A structural formula is a formula which shows how atoms are bonded to each other in a molecule. The structural formulae of the first ten alkanols are shown below.

Molecular formula	Structural formula
CH ₃ OH	$ \begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{OH} \\ \\ \text{H} \end{array} $
C ₂ H ₅ OH	$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H} - \text{C} - \text{C} - \text{OH} \\ \quad \\ \text{H} \quad \text{H} \end{array} $
C ₃ H ₇ OH	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array} $
C ₄ H ₉ OH	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $
C ₅ H ₁₁ OH	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $
C ₆ H ₁₃ OH	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $
C ₇ H ₁₅ OH	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $
C ₈ H ₁₇ OH	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $
C ₉ H ₁₉ OH	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $
C ₁₀ H ₂₁ OH	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array} $

SKELETAL FORMULAE OF ALKANOLS

- A skeletal formula is a diagrammatic representation of a molecule in which lines represent bonds between atoms.

- The lines are drawn in zigzag fashion. Carbon atoms are implied when two bonds meet and no atom listed except for functional groups. The skeletal formulae of the first ten alkanols are shown below.

Molecular formula	Skeletal formula
CH ₃ OH	
C ₂ H ₅ OH	
C ₃ H ₇ OH	
C ₄ H ₉ OH	
C ₅ H ₁₁ OH	
C ₆ H ₁₃ OH	
C ₇ H ₁₅ OH	
C ₈ H ₁₇ OH	
C ₉ H ₁₉ OH	
C ₁₀ H ₂₁ OH	

TYPES OF ALKANOLS

Alkanols fall into different classes depending on how the hydroxyl (OH) is positioned on the carbon atoms. They include;

1. *Primary alkanols*

These are alkanols in which the carbon carrying the hydroxyl (OH) group is attached to only one alkyl group. This carbon atom is also attached to two hydrogen atoms. The only exception is methanol although it is also classified as a primary alkanol.

2. *Secondary alkanols*

In these alkanols the carbon atom bearing the hydroxyl group is attached to two alkyl groups. It is also attached to only one hydrogen atom. Some examples of secondary alkanols are butan-2-ol and pentan-2-ol.

3. *Tertiary alkanols*

These are alkanols in which the carbon atoms containing the hydroxyl group is attached to three alkyl groups. This carbon is not attached to any hydrogen atoms. Some examples of tertiary alkanols are 2-methylbutan-2-ol and 2-methylpentan-2-ol.

PREPARATION OF ETHANOL

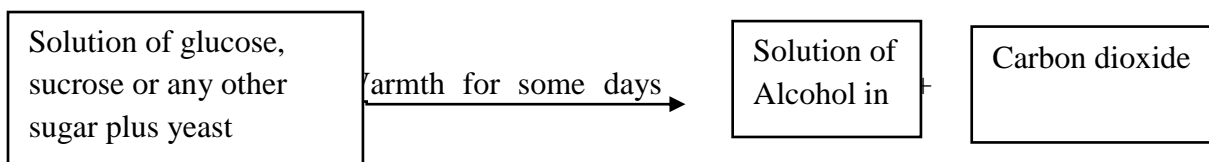
Ethanol is prepared by two main methods as follows

- Fermentation of sugars.
- Hydration of alkenes

PREPARATION OF ETHANOL BY FERMENTATION OF SUGARS

✎ **Fermentation** is the reaction whereby sugars are converted to alcohols and carbon dioxide by the action of enzymes from yeast.

✎ The fermentation process can be summarized as follows.



✎ Fermentation becomes possible when the yeast respire anaerobically, so a container where fermentation is taking place should be air tight.

✎ The chemical equation for the fermentation process is as follows:



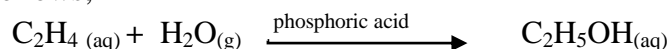
✎ The ethanol produced during fermentation is toxic to the enzymes.

✎ It kills the yeast at concentrations greater than 15% by volume.

✎ To increase the concentration of ethanol, distillation process is carried out

PREPARATION OF ETHANOL BY HYDRATION OF ALKENES

✎ Hydration is the addition of water molecules into other molecules. During this process, a water molecule adds across the double bond of alkenes. The equation for this reaction is as follows;



✎ Phosphoric acid is used as a catalyst. This chemical reaction occurs at high pressure and temperature (about 300⁰C and 60 atm). The produced ethanol is condensed to form a liquid and unreacted ethane is recycled.

PHYSICAL PROPERTIES OF ALKANOLS

1) *conductivity*

Generally, alkanols do not conduct electricity.

2) *melting and boiling points*

The melting and boiling points of alkanols increase with increase in relative molecular mass. As the molecular mass increases, intermolecular forces of attraction between the molecules increase. As these forces increase, more energy is needed to separate the molecules. This increases the melting and boiling points.

3) *density*

The densities increase with increase in relative molecular mass. This is because of the increase in intermolecular forces of attraction between molecules making them very close to

each other. This reduces the volume occupied by the molecule and from the formula $D = \frac{M}{V}$, the density then increases.

4) solubility

The first three alkanols are soluble (miscible) in water. The rest are immiscible.

Why is it that alkanols of larger molecular sizes do not dissolve in water?

- ✎ What makes alkanols to dissolve in water is the –OH group. The –OH group makes alkanols to be like water as water molecule is H-O-H. So when the molecular size of an alkanol increases the –OH group remains the same. This means that for larger molecules of alkanols the relative size of the –OH group is small to bring about their solubility in water. Hence larger alkanols are insoluble in water.

The trends in physical properties of alkanols are summarized in the following table;

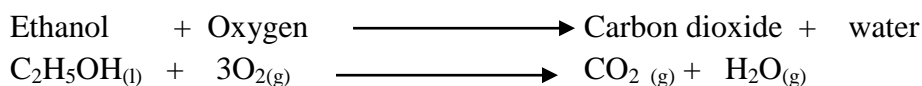
Properties Name	B.Pt (°C)	M. pt (°C)	Density (g/cm³)	Solubility in water
Methanol	-98	65	0.791	Soluble
Ethanol	-117	78.5	0.789	Soluble
Propan-1-ol	-127	97	0.803	Soluble
Butan-1-ol	-89.5	117	0.810	Fairly soluble
Pentan-1-ol	-79	138	0.814	Slightly soluble
Hexan-1-ol	-55.6	157	0.814	Slightly soluble
Heptan-1-ol	-34	176	0.822	Slightly soluble
Octan-1-ol	-15	195	0.825	Slightly soluble
Nonan-1-ol	-7	212	0.827	Slightly soluble
Decan-1-ol	6	228	0.829	Slightly soluble

CHEMICAL PROPERTIES OF ALKANOLS:

Alkanols form a homologous series with similar chemical properties as follows:

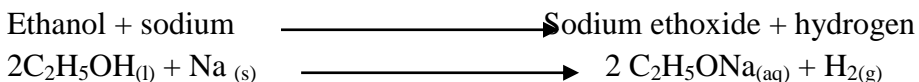
1. Combustion of ethanol

- ✎ When ethanol is ignited in plentiful air, it burns with a blue flame to produce carbon dioxide and water.



2. Reaction with sodium

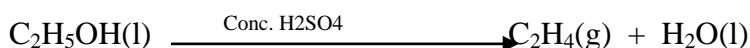
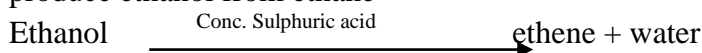
- ✎ When sodium reacts with ethanol, it forms a compound called *sodium ethoxide*.
 ✎ Effervescence occurs and a colorless gas which burns with a pop sound is produced. The gas is hydrogen. The reaction is similar to the reaction of sodium with water but is less vigorous.



Ethanol also reacts with other metals such as potassium to produce the corresponding ethoxide and hydrogen gas.

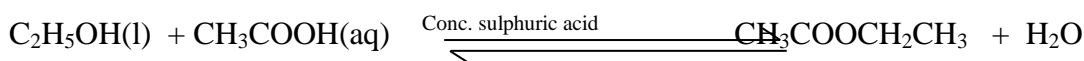
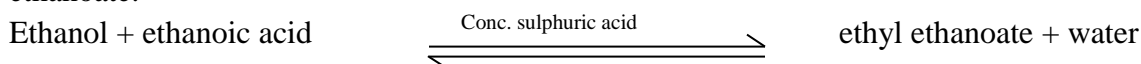
3. Dehydration

- ✎ In the presence of concentrated sulphuric acid, alkanols undergo dehydration process to form ethene and water. This reaction happens at 180°C. This is reverse of the reaction used to produce ethanol from ethane



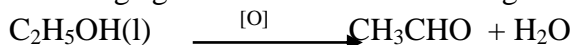
4. Ester formation (esterification)

- ✎ Alkanols react with alkanolic acids to form sweet smelling compounds called **esters**.
- ✎ Ethanol for example, reacts with ethanoic acid in the presence of a few drops of concentrated acid (a catalyst) to form a substance with a characteristic sweet and fruity smell called ethyl ethanoate.

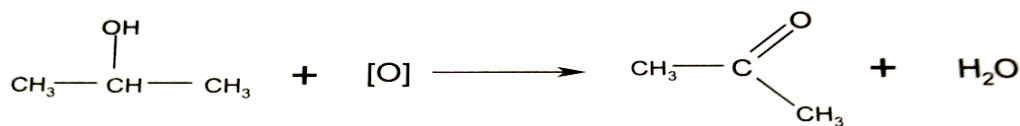


5. Oxidation of ethanol

- ✎ Oxidation of alkanols involves adding an oxygen atom to the alkanol molecule. This is done when alkanols are mixed with oxidizing agents such as acidified potassium dichromate (K₂Cr₂O₇).
- Primary alkanols are oxidized to alkanals (aldehydes) and water in the presence of an oxidizing agent. Consider the following examples;



- Secondary alkanols are oxidized to alkanones (ketones) and water in the presence of an oxidizing agent.



Propan-2-ol

Propanone

- Tertiary alkanols cannot be oxidised because the carbon to which the OH is attached to is attached to three other carbon atoms and no hydrogen can be released from it.

USES OF ETHANOL

- 1) It is used as an ingredient in the manufacturing of medicines eg Philips gripe water for the new born babies.
- 2) It is used as a solvent. It dissolves substances such as paint, varnish and drugs.
- 3) It is used in the manufacturing of alcoholic drinks or beer.
- 4) It is used as fuel.
- 5) It is used as an antiseptic and some antibacterial soaps and wipes.
- 6) It is used as a thermometric liquid.
- 7) It is used in making food flavorings.

- 8) It is used as a raw material for the production of methylated spirit. Methylated spirit is ethanol with a small amount of methanol (a poisonous substance added to stop people from drinking).

DANGERS OF EXCESSIVE CONSUMPTION OF ALCOHOLIC DRINKS AND MEDICINES ON HUMAN HEALTH

- 1) It leads to accidents such as car crash, falls burns and drowning.
- 2) It causes increased injuries when doing the job and loss of productivity
- 3) It causes people to commit suicide, sexual assault and domestic violences
- 4) It damages some body organs such as lungs and kidneys
- 5) It causes high blood pressure, stroke and other heart related problems
- 6) In pregnant women, it may lead to the impairment of the foetus. This is why pregnant women are not encouraged to take alcohol

TEST FOR ALKANOLS

I. SOLUBILITY TEST

Add 15 drops of water to a test tube.

Add at least 10 drops of the test liquid.

Positive result: One layer forms

Interpretation: The sample liquid could be an alkanol.

II. SODIUM TEST

Add 15 drops of test liquid in a test tube.

Place a match head amount of freshly cut sodium metal

Positive result: Sodium metal disappears. Bubbles of a gas form. The gas produces a pop sound when ignited. The liquid formed turns phenolphthalein solution to pink colour.

Interpretation: The test liquid is either an alkanol.

Exercise 10

1. Write the formula of alkanols where $n = 2$ and 3 .
2. Name the alkanols in (1) above.
3. Using a chemical equation, explain how ethanol is prepared from fermentation.
4. Which gas is produced in the process of fermentation of carbohydrates?
5. Explain the trends in melting and boiling points of the first ten alkanols.
6. Compare the melting and boiling points of alkanols to those of alkanes with the same number of carbon atoms. Give reason for your answer.
7. Explain why larger alkanols are insoluble in water yet they have an $-OH$ group.
8. Write a word equation for the reaction between ethanol and ethanoic acid in the presence of concentrated sulphuric acid. What is the reaction called?
9. Explain with an equation what happens when ethanol is burnt in air
10. State the physical properties of ethanol.
11. Write a chemical equation for the fermentation of glucose ($C_6H_{12}O_6$).
12. Explain why alkanols do not undergo oxidation reaction
13. Describe the formation of ethanol by fermentation.
14. How would you increase the alcohol content of ethanol after fermentation?
15. Explain any five negative effects of consuming alcohol.

UNIT 8

ALKANOIC ACIDS

- ☞ Alkanoic acids are organic compounds belonging to the group of oxycarbons.
- ☞ They are also called carboxylic acids because they have a carboxyl group (**-COOH**) in their formula. The carboxyl group serves as their functional group. They may be regarded as being derived from alkanes by replacing one of the hydrogen atoms with a carboxyl group; -COOH group.

GENERAL FORMULA AND IUPAC NOMENCLATURE OF ALKANOIC ACIDS

- ✎ The general formula of Alkanoic acids is **$C_NH_{2N+1}COOH$** (where N = 0, 1, 2, 3, 4...).
- ✎ The systematic name of a carboxylic acid comes from an alkane with corresponding number of carbon atoms. The **-e** in the name of alkane is replaced with **-oic acid**.
- ✎ The first member of the series is therefore 'methanoic acid'. And the molecular formula can be worked out as follows:
 - General formula: **$C_NH_{2N+1}COOH$** ,
 - Molecular formula when N = 0, $C_0H_{(2 \times 0)+1}COOH = HCOOH$.
- ✎ For methanoic acid we substitute zero in the general formula because there is already one carbon atom which is given by the functional group; -COOH

MOLECULAR AND CONDENSED FORMULAE OF ALKANOIC ACIDS

The table below shows the molecular and condensed formulae of the first ten alkanoic acids.

Name of alkanoic acid	Molecular formula	Condensed formula
Methanoic acid	HCOOH	HCOOH
Ethanoic acid	CH ₃ COOH	CH ₃ COOH
Propanoic acid	C ₂ H ₅ COOH	CH ₃ CH ₂ COOH
Butanoic acid	C ₃ H ₇ COOH	CH ₃ (CH ₂) ₂ COOH
Pentanoic acid	C ₄ H ₉ COOH	CH ₃ (CH ₂) ₃ COOH
Hexanoic acid	C ₅ H ₁₁ COOH	CH ₃ (CH ₂) ₄ COOH
Heptanoic acid	C ₆ H ₁₃ COOH	CH ₃ (CH ₂) ₅ COOH
Octanoic acid	C ₇ H ₁₅ COOH	CH ₃ (CH ₂) ₆ COOH
Nonaic acid	C ₈ H ₁₇ COOH	CH ₃ (CH ₂) ₇ COOH
Decanoic acid	C ₉ H ₁₉ COOH	CH ₃ (CH ₂) ₈ COOH

STRUCTURAL FORMULAE OF ALKANOIC ACIDS

The table shows the structures of the first ten alkanoic acids.

Molecular formula	Structural formula
HCOOH	$\begin{array}{c} \text{O} \\ \\ \text{H} - \text{C} - \text{OH} \end{array}$
CH ₃ COOH	$\begin{array}{c} \text{H} \quad \text{O} \\ \quad \\ \text{H} - \text{C} - \text{C} - \text{OH} \\ \\ \text{H} \end{array}$
C ₂ H ₅ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{O} \\ \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \\ \text{H} \quad \text{H} \end{array}$
C ₃ H ₇ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \\ \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array}$
C ₄ H ₉ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \\ \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$
C ₅ H ₁₁ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \\ \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$
C ₆ H ₁₃ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \\ \quad \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$
C ₇ H ₁₅ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \\ \quad \quad \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$
C ₈ H ₁₇ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \\ \quad \quad \quad \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$
C ₉ H ₁₉ COOH	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{O} \\ \quad \quad \quad \quad \quad \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} - \text{OH} \\ \quad \quad \quad \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$

SKELETAL FORMULA FOR THE ALKANOIC ACIDS

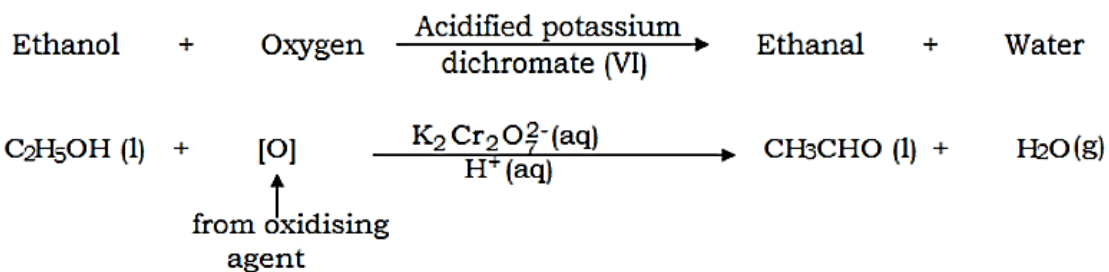
Name of alkanoid acid	Condensed formula	Skeletal formula
Methanoic acid	HCOOH	HCOOH
Ethanoic acid	CH ₃ COOH	
Propanoic acid	CH ₃ CH ₂ COOH	
Butanoic acid	CH ₃ (CH ₂) ₂ COOH	
Pentanoic acid	CH ₃ (CH ₂) ₃ COOH	
Hexanoic acid	CH ₃ (CH ₂) ₄ COOH	
Heptanoic acid	CH ₃ (CH ₂) ₅ COOH	
Octanoic acid	CH ₃ (CH ₂) ₆ COOH	
Nonanoic acid	CH ₃ (CH ₂) ₇ COOH	
Decanoic acid	CH ₃ (CH ₂) ₈ COOH	

PREPARATION OF ALKANOIC ACIDS

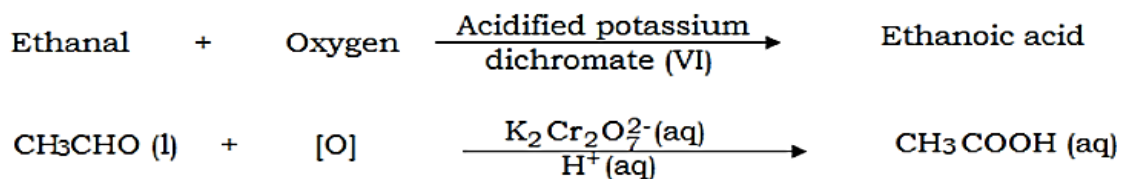
Alkanoic acids can be prepared by the oxidation of corresponding alkanols. The oxidising agents are acidified potassium permanganate (VII) or acidified potassium dichromate (VI).

The oxidation is done in two steps:

a. First, the alkanol is oxidized to an alkanal. For example



b. Then, the alkanal is further oxidized to an alkanoid acid.



PROPERTIES OF ALKANOIC ACIDS

Physical properties

- 1) They are soluble in water. However, their solubility decreases with increase in molecular sizes. This is because with an increase in molecular size, the relative size of $-\text{COOH}$ which brings about their solubility decreases.
- 2) The boiling and melting points increases with an increase in molecular sizes.
- 3) The densities also increase with an increase in size of the molecules.
- 4) They have a strong odour/choking smell
- 5) They have Acid physical properties: sour taste and pH less than 7.

Chemical Properties

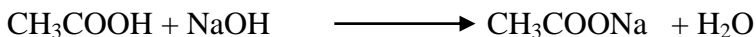
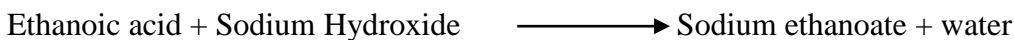
1. They change the color of indicators. For instance, the blue litmus paper is changed to red by a carboxylic acid. **Indicators** are substances that change colors when mixed with either acids or bases. Some common indicators used in the laboratory are shown below together with their colors in acid and basic solution.

Name of Indicator	Acid-color	Basic color
Universal indicator	red	Purple/blue
phenolphthalein	colourless	pink
Litmus	red	blue

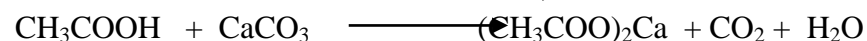
2. Carboxylic acids react with a bases to produce salt and water as shown by the following word equation:



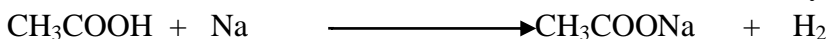
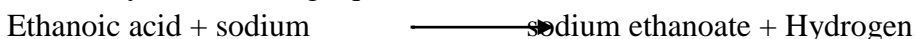
For example;



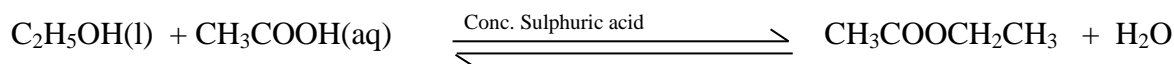
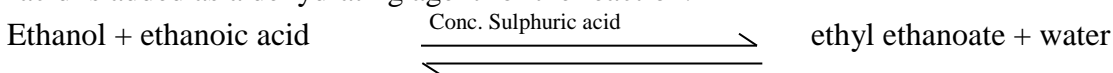
3. Alkanoic acids react with carbonates to form salt, water and carbon dioxide as shown by the word and chemical equations below;



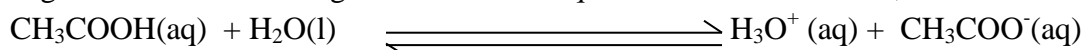
4. Alkanoic acids react with metals to form salt and hydrogen gas. For example, ethanol reacts with alkali metals such as Na, Li, and K to produce hydrogen(H_2) and an ethanoate salt as shown by the following equations:



5. Alkanoic acids react with alkanols to form esters and water. This reaction process is known as esterification. For example, Ethanoic acid can react with ethanol in the presence of a few drops concentrated sulphuric acid to form ethyl ethanoate and water. Concentrated sulphuric acid is added as a dehydrating agent for the reaction.



6. Alkanoic acids are weak acids and they ionize slightly/partially in water, giving few hydrogen ions. The following is an *ionisation equation* for ethanoic acid;



The hydronium ions (H_3O^+) and ethanoate ions (CH_3COO^-) produced when ethanoic acid is added to water allow ethanoic acid to partially conduct electricity.

NATURAL SOURCES OF CARBOXYLIC ACIDS

Some alkanolic acids occur naturally in the following;

- Vinegar (has ethanoic acid).
- In ants and bee stings (methanoic acid)
- Cheese (butanoic acid)
- Sour milk (lactic acid)

UNIT 9

ALKANALS AND ALKANONES

- ☞ Alkanals and alkanones are two separate homologous series of organic oxycarbonyls which are closely related.
- ☞ Alkanals are commonly called aldehydes while alkanones are commonly called ketones. Just like alkanols, alkanals and alkanones are compounds of organic oxycarbonyls since they are made up of oxygen in their formula.
- ☞ However, alkanals and alkanone have different functional groups from alkanols.

FUNCTIONAL GROUPS OF ALKANALS AND ALKANONES

- ☞ Alkanals and alkanones are simple compounds which have a **carbonyl group** – ‘a carbon-oxygen double bond’ as their functional group.
- ☞ They are simple in the sense that they don't have other reactive groups like -OH or -Cl attached directly to the carbon atom in the carbonyl group - as you might find, for example, in carboxylic acids containing -COOH.

GENERAL MOLECULAR FORMULA

Aldehydes: $\text{C}_n\text{H}_{2n+1}\text{CHO}$

Ketones: $\text{C}_n\text{H}_{2n+1}\text{COR}$ (R is the alkyl group)

DIFFERENCES IN THE STRUCTURES OF ALKANALS AND ALKANONES

- ✗ In alkanals, the carbonyl group has a hydrogen atom attached to it together with either a second hydrogen atom or, an alkyl group.
- ✗ In alkanones, the carbonyl group has two hydrocarbon groups (alkyl groups) attached.
- ✗ Alkanones contain a carbon-oxygen double bond just like (alkanals) aldehydes, but it is in the middle of a carbon chain. There isn't a hydrogen atom attached to the group as there is in aldehydes

NOMENCLATURE OF ALKANALS AND ALKANONES

- ✎ The names of alkanals (aldehydes) end in **al**. The name accounts for the total number of carbon atoms in the longest chain including the one in the carbonyl group.
- ✎ If you have side groups attached to the chain, you always count from the carbon atom in the carbonyl group as being number 1.
- ✎ The following table gives the names of the first 5 alkanals.

Molecular formula	No. of carbon atoms	Prefix	Name
HCHO	1	Meth	Methanal
CH ₃ CHO	2	Eth	Ethanal
C ₂ H ₅ CHO	3	Prop	Propanal
C ₃ H ₇ CHO	4	But	Butanal
C ₄ H ₉ CHO	5	Pent	Pentanal

- ✎ Alkanones have '**anone**' as their suffix. For example, an alkanone with four carbon atoms is called butanone.

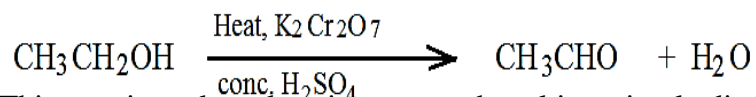
The following table gives the names of the first five alkanones.

Molecular formula	No. of carbon atoms	Prefix	Name
CH ₃ COCH ₃	3	Prop-	Propanone
C ₂ H ₅ COCH ₃	4	But-	Butanone
C ₃ H ₇ COCH ₃	5	Pent-	Pentanone
C ₄ H ₉ COCH ₃	6	Hex-	Hexanone
C ₅ H ₁₁ COCH ₃	7	Hept-	Heptanone

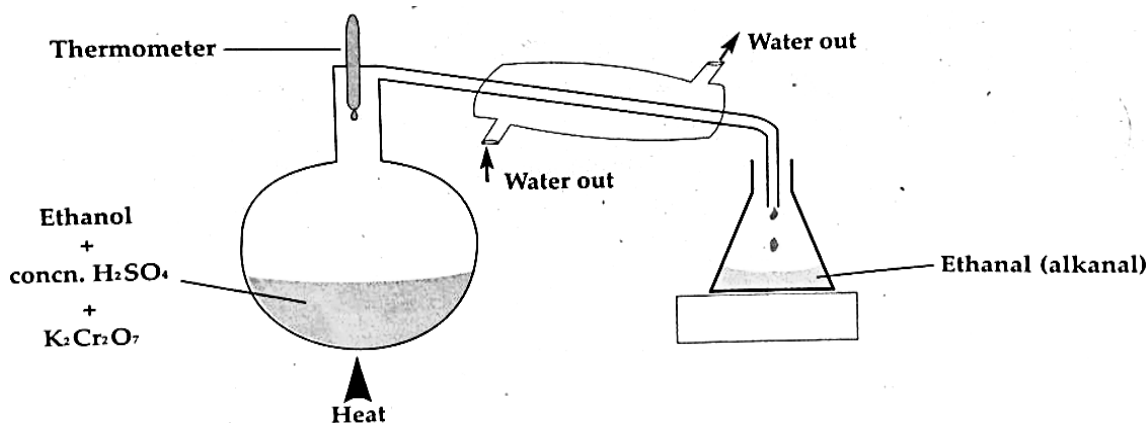
- ✎ However, the position of the carbonyl group in ketones is not fixed as in aldehyde.
- ✎ The IUPAC system of naming ketones includes the position of the functional group. This is why we have compounds such as pentan-2-one and pentan-3-one

PREPARATION OF ALKANALS AND ALKANONES

- 1) Alkanals are produced from the oxidation of primary alkanols. For example, ethanal is produced by heating ethanol with a mixture of potassium (IV) dichromate (K₂Cr₂O₇) and concentrated sulphuric acid (H₂SO₄).
 - ✎ This acidified potassium (IV) dichromate acts as an oxidizing agent. Ethanol loses two hydrogen atoms to form ethanal. This equation below illustrates this process.



- ✎ This reaction takes place in excess ethanol in a simple distillation apparatus. Alkanals have a lower melting and boiling point temperatures.
- ✎ When heating the alkanol with the oxidizing agent in the boiling tube/ flask, the alkanal which is more volatile vaporizes and rises up the distillation column where it is condensed and collected as distillate in the collecting tube.



- 2) Ketones are made from the oxidation of secondary alkanols. For example, propanone is prepared from the oxidation of secondary propanol (propan-2-ol).
- ✎ The process of making propanone is by heating secondary propanol in the mixture of acidified potassium (IV) dichromate using simple distillation apparatus.
- ✎ Propanol loses two hydrogen atoms to produce propanone which vaporizes first from the boiling flask then collected as a distillate. Water is also produced from this process.

PHYSICAL PROPERTIES OF ALKANALS AND ALKANONE

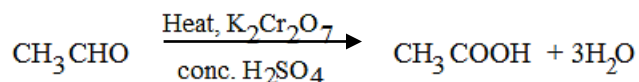
- 1) Small aldehydes and ketones are soluble in water, but the solubility decreases with increasing length of carbon chains. The longer the chains, the lower the solubility. The solubility of alkanones and alkanals is possible regardless of the absence of OH group, since they form hydrogen bonds with water molecules.
- 2) The first member of alkanals (alkanal) is a gas at room temperature while the rest of alkanals are liquids. All alkanones are in liquid form at room temperature.
- 3) Both aldehydes and ketones are less dense than water. They therefore float on top of water. However the density of these homologous series increases with the size of the molecule.
- 4) Alkanals and alkanones have high melting and boiling points (mpt and Bpts) as compared to hydrocarbon of similar relative molecular mass (RMM), since they have strong IMF. Passing along the homologous series, the molecular mass increase hence the IMF increases, as a result mpt and bpt also increases. The following table gives mpts and bpts of the first five alkanals and alkanones.

Alkanal	Mpt (°C)	Bpt (°C)	Alkanones	Mpt (°C)	Bpt (°C)
Methanal	-92	-21	propanone	-95	56
Ethanal	-124	21	butanone	-87	80
Propanal	-81	49	pentanone	78	102
Butanal	-99	46	hexanone	-55	128
pentanal	-60	103	heptanone	-36	151

CHEMICAL PROPERTIES OF ALKANALS AND ALKANONES

- ✎ Aldehydes and ketones undergo similar chemical properties because they contain the same functional group. The following are some of the chemical reactions they undergo;

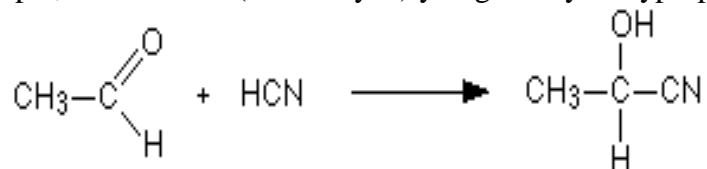
- Both aldehydes and ketones burn in air to produce carbon dioxide and water. For example;
 - Combustion of propanal $\text{CH}_3\text{CH}_2\text{CHO}_{(l)} + 4\text{O}_{2(g)} \rightarrow 3\text{CO}_{2(g)} + 3\text{H}_2\text{O}_{(l)}$
 - Combustion of propanone $\text{CH}_3\text{COCH}_{3(l)} + 4\text{O}_{2(g)} \rightarrow 3\text{CO}_{2(g)} + 3\text{H}_2\text{O}_{(l)}$
- Aldehydes are easily oxidized to form an organic acid. The reaction takes place in the presence of oxidizing agents such as acidified potassium (IV) dichromate solution. For example, the oxidation of ethanal occurs as;



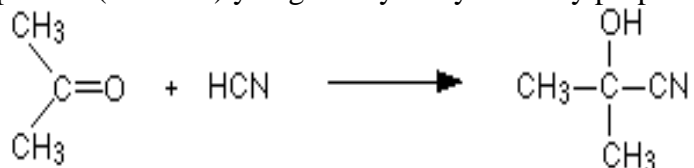
- Alkanals and alkanones undergo a kind of addition reaction called nucleophilic addition. In this reaction, compounds like hydrogen cyanide (HCN), sodium bisulphate (NaHSO_4), and water are added to ketones and aldehydes.
- ✎ This reaction is possible due to the presence of the carbon-oxygen double bond which breaks when nucleophiles are added. A nucleophile is a species which has a tendency to donate electrons or a lone pair of electrons to other reactants.

a. Addition of hydrogen cyanide to aldehydes and ketones

- ✎ Hydrogen cyanide adds across the carbon-oxygen double bond in aldehydes and ketones to produce compounds known as hydroxynitriles.
- ✎ For example, with ethanal (an aldehyde) you get 2-hydroxypropanenitrile:

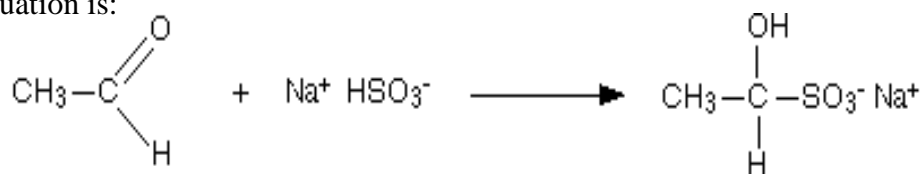


- ✎ With propanone (a ketone) you get 2-hydroxy-2-methylpropanenitrile:

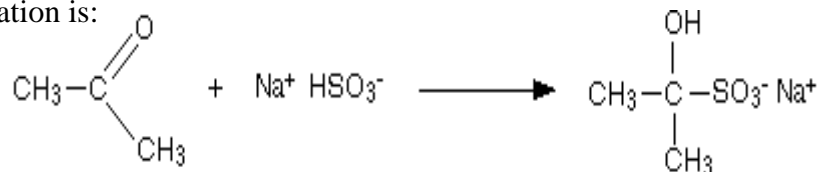


b. Addition of sodium hydrogen sulphite to aldehydes and ketones

- ✎ This reaction only works well for aldehydes. In the case of ketones, one of the hydrocarbon groups attached to the carbonyl group needs to be a methyl group.
- ✎ Large groups attached to the carbonyl group get in the way of the reaction happening. In the case of ethanal, the equation is:



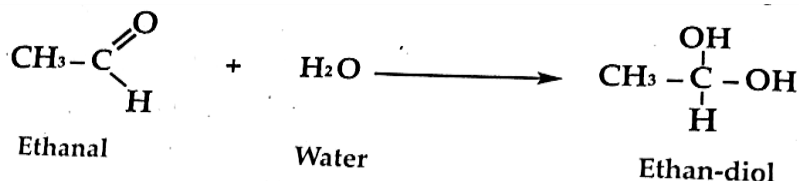
with propanone, the equation is:



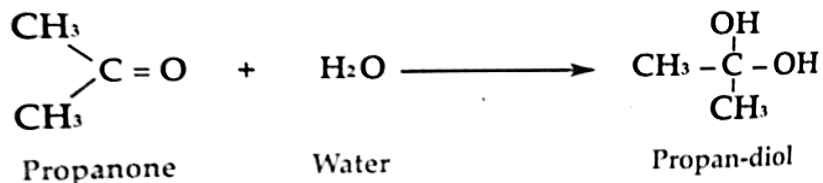
These compounds are rarely named systematically, and are usually known as "hydrogensulphite (or bisulphite) addition compounds".

c. Addition of water to ketones and aldehydes

When water is added to the carbon-oxygen bond, a hydrated compound with two hydroxyl groups (diol) is produced. For example when water is added to ethanal, the following is produced;



In the case of propanone (alkanone), the following is produced;



TEST FOR ALKANALS AND ALKANONES

There are many tests that can be carried on alkanals and alkanones. Some tests are used to detect the presence of a carbonyl group while others are used to differentiate alkanals from alkanones.

1) BRADY'S TEST

This test is used to detect the presence of a carbonyl group ($-\text{C}=\text{O}$) in compounds.

Method

The following steps are followed when carrying out a brads test;

1. Put a match head amount of 2, 4-dinitrophenylhydrazine (2,4-DNPH) in a test tube.
2. Add two drops of ethanol to moisten the 2,4-DNPH powder.
3. Add two drops of concentrated sulphuric acid and a clear yellow liquid is formed.
4. Add one drop of the liquid under the test (i.e alkanal) to the test tube.
5. Observe what happens.

Positive result: A red precipitate forms

Interpretation: The test liquid has a carbonyl functional group.

It can either be an alkanal or an alkanones because both give a positive result.

FEHLING'S TEST

This test is used to differentiate alkanals from alkanones.

Method

The following steps are followed when carrying out a Fehling's test;

1. Put 15 drops of Fehling's solution into a test tube.
2. Add 15 drops of the liquid under the test to the Fehling's solution in the test tube.
3. Heat the contents of the test tube gently for a minute.

Positive result: A red precipitate forms.

Interpretation: The test liquid is an alkanal.

TOLLEN'S TEST

This is also called the silver mirror test. Just like the Fehling's test, it is used to differentiate alkanals from alkanones. It uses Tollen's reagent which oxidizes the aldehyde into a corresponding carboxylic acid. Tollen's reagent is a colourless basic aqueous solution containing silver ions coordinated to ammonia $[\text{Ag}(\text{NH}_3)_2]^+$

Method

The following steps are followed when carrying out a Tollen's test;

1. Put 1ml of Tollen's reagent into a clean test tube.
2. Add 2 drops of the liquid under the test to the Tollen's reagent in the test tube.
3. Heat the contents of the test tube gently for a minute.

Results and interpretation

If the liquid under the test is an alkanal then the silver mirror or a black precipitate is formed. When the liquid under the test is not an alkanal then no silver mirror nor is black precipitate formed.

USES OF ALKANALS AND ALKANONES

- 1) Alkanals such as formal aldehyd~~n~~ is used to preserve dead animals.
- 2) They are used in the manufacturing of organic acid since they can be oxidized further to produce carboxylic acids.
- 3) They are used in the production of perfumes and flavourings
- 4) They are used in the production of explosives
- 5) They are used in the formation of industrial polymers such as melamine and nylon.
- 6) They are used in the manufacturing of synthetic resins.

ALKANONES

- 1) They are used as solvents and fingernail polish remover.
- 2) They are used in the manufacturing of insecticides
- 3) Ketones are also used in making perfumes, flavourings, dyes in addition to making of plastics..

UNIT 10

ALKANOATES (ESTERS)

Introduction

Esters are a group of organic compounds best known for their interesting odors (scent). Esters are one of the most widely used organic compounds, making up fats, oils, fragrances, flavorings, explosives, plastics, and clothing. They are commonly created by combining a carboxylic acid with an alcohol in the presence of a strong acid. This process is referred to as esterification. Esterification is an example of condensation reaction (which is defined as the process in which two or more molecules combine to form a larger molecule with the release of water).

MORE ON ESTERS

- ✎ Esters are products of *condensation reactions* between alcohols and carboxylic acids.
- ✎ They can also be described as derivatives of carboxylic acids in which the OH of the carboxyl group has been replaced by an –O-R from an alcohol.
- ✎ Esterification is the process of forming an ester from the reaction of alkanolic acids and alkanols. In this process, water is removed from the two reacting materials. The general esterification process can be represented as follows;
Alcohol + carboxylic acid \rightarrow ester + water.
- ✎ Esters have the general formula, **R-COOR'**, where R and R' are alkyl groups.

NAMING ALKANOATES

- The name of an ester contains two parts; the alkanolic acid part and the alkanol part.
- The side of the ester that was originally the alcohol part is named first like a branch.
- The side of the ester that was originally the acid part is named last and has the ending '**oate**' to indicate the ester group.
- The two parts of the name are two separate words. Therefore, an ester derived from methanol and propanoic acid is called **methyl propanoate**.

The following are some of the *common and popular* alkanooates and their names;

Formula	Name and where it is found
$\text{CH}_3 - \overset{\text{O}}{\parallel} \text{C} - \text{O} - \text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$	Pentyl acetate or ethanoate – banana oil
$\text{CH}_3 - \overset{\text{O}}{\parallel} \text{C} - \text{O} - \text{C}_8\text{H}_{17}$	Octyl acetate or ethanoate – orange oil
$\text{CH}_3\text{CH}_2\text{CH}_2 - \overset{\text{O}}{\parallel} \text{C} - \text{O} - \text{CH}_3$	Methyl butanoate – apple oil
$\text{CH}_3 - \overset{\text{O}}{\parallel} \text{C} - \text{O} - \text{CH}_2\text{CH}_3$	Ethyl ethanoate – industrial solvent
$\text{CH}_3\text{CH}_2\text{CH}_2 - \overset{\text{O}}{\parallel} \text{C} - \text{O} - \text{CH}_2\text{CH}_3$	Ethyl butanoate – pineapple oil

SOURCES OF ALKANOATES

Natural alkanooates

- Alkanooates are found naturally in many plants and animals. For example, the flavor and fragrance (sweet smell) in fruits like banana and oranges are due to the alkanooates that they contain. They are used as energy storage compounds.
- In some animals such as whales, esters are in the form of fats. Just like In plants, animals use esters as energy storage compounds.

Synthetic alkanooates

- ✎ **Polyesters** are synthetic fibres used as substitutes for cotton and wool with greater strength.
- ✎ **Ethylethanoate** is another synthetic alkanooate. It is used as fruit flavouring for sweets and drinks.
- ✎ Other synthetic esters are used in making perfumes, youghurt and orange flavoured juice.

PHYSICAL PROPERTIES OF ESTERS

- Are colourless volatile liquids
- Insoluble in water (if you add it to water, the mixture forms two layers)
- Relatively low melting and boiling points since they have weak attractions (no hydrogen bonds)
- Have very pleasant fragrances of fruits and flowers.

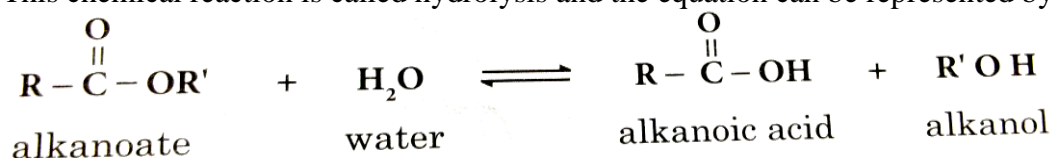
CHEMICAL PROPERTIES OF ESTERS

- Combustion:** alkanooates burn in air to produce a bright flame with water and carbon dioxide. The following equation illustrates this reaction:

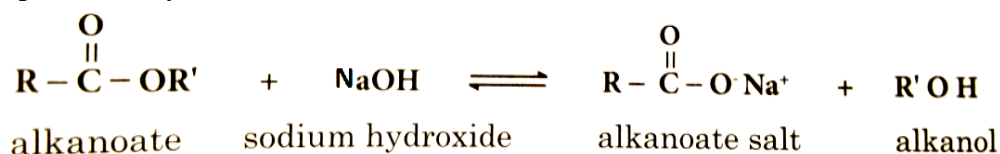


- Hydrolysis with water and sodium hydroxide or potasium hydroxide:**

Esters can be broken apart under acidic conditions by water to form an alkanooic aci and alkanol. This chemical reaction is called hydrolysis and the equation can be represented by;



Esters can also be broken apart under basic conditions by sodium hydroxide or potassium hydroxide to form alkanooate salts and alkanols. This chemical reaction can generally be represented by;



USES OF ALKANOATES

- ✓ **Solvents;** some alkanooates are used as solvents of adhensives. Ehyl ethanoate fpor example is a solvent of polysterene.
- ✓ **Flavourings;** alkanooates are used to give sweets and drinks a fruit flavor. These fruit flavoured sweets and drinks, such as youghurt and juices are very common in shops.
- ✓ **Fragrances;** some alkanooates are used soaps, petroleum jellies and perfumes to give them a sweet fragrance.
- ✓ **Soap making;** some esters are used in making soap and detergents.

FATS AND OILS

- ✎ Fats and oils are natural esters which are used as energy storage compounds in both plants and animals. They belong to a larger class known as lipids.
- ✎ Because they are esters, they are made through the same process of esterification involving an alkanol (known as glycerol) and large molecules of carboxylic acids (known as fatty acids).

- ✎ A glycerol has three hydroxyl (–OH) groups. Hence it requires three molecules of fatty acids to combine with so as to form a molecule of a fat or oil (alkanoates) and three molecules of water are produced as by-product.
- ✎ Oils contain a large proportion of unsaturated acids and have low melting points. They are therefore in liquid form at room temperature. However fats contain saturated acids and have higher melting points. They are therefore in solid form at room temperature’
- ✎ Oils can be converted to fats by reacting them with hydrogen gas

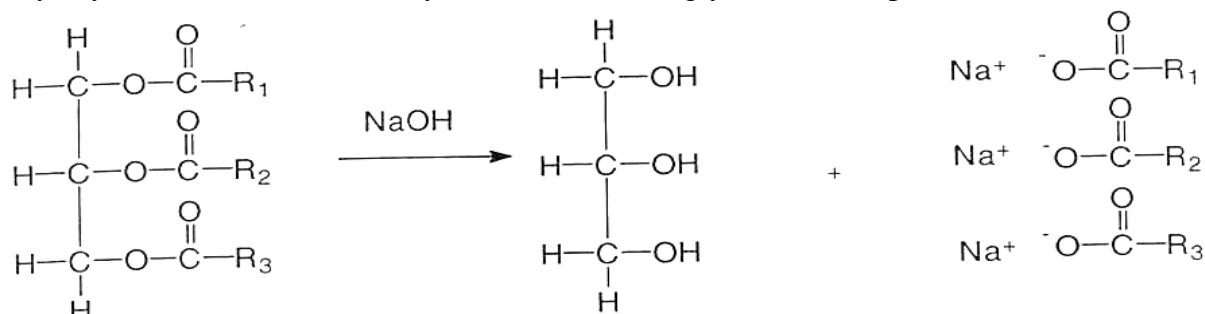
SOAPS AND DETERGENTS

- ✎ Soaps and detergents are substances which are used for cleaning.

SOAP MAKING

- ✎ The main raw materials for the soap making process are animal fats or vegetable oils and any strong base (e.g. sodium hydroxide or potassium hydroxide).
- ✎ The process involves hydrolyzing a fat under alkaline (basic) conditions. Heat is needed for the reaction to take place between the reactants.
- ✎ The reaction is called saponification, and produces one molecule of glycerol and three molecules of soap, for each molecule of fat taken.
- ✎ The following chemical equation illustrates the formation of soap from the reaction of a Glyceryl tristearate (Triglyceride) with sodium hydroxide solution:

Glyceryl tristearate + Sodium hydroxide \longrightarrow glycerol + 3 soap molecules



- ✎ Once the soap molecules are produced, special fragrance oils are added. These improve the smell of the soap. Colorants are also added to give the soap a particular desired colour.

IDENTIFICATION OF UNKNOWN ORGANIC COMPOUNDS

- ✎ In the laboratory, unknown organic compounds can be identified by doing different tests. These tests are based on the physical and chemical properties of the test substances. It is therefore important to understand the physical and chemical properties of all homologous series for the concept of identification.

☞ The following table summarizes the properties of different organic families.

Organic compound	Physical properties	Chemical properties
Alkanes	<ul style="list-style-type: none"> Insoluble in water Small molecules of alkanes are gases while large molecules are liquids or solids 	<ul style="list-style-type: none"> Generally unreactive Burn in air
Alkenes	<ul style="list-style-type: none"> Insoluble in water Small molecules of alkenes are gases while large molecules are liquids or solids 	<ul style="list-style-type: none"> Burn in air changes bromine from brown to colourless
Alkanols	<ul style="list-style-type: none"> Liquid at room temperature Soluble in water. The solubility decreases with increasing size of molecules 	<ul style="list-style-type: none"> Burn in air Reacts with sodium to release hydrogen gas
Carboxylic acids	<ul style="list-style-type: none"> Soluble in water but The solubility decreases with increasing size of molecules 	<ul style="list-style-type: none"> Turn phenolphthalein in sodium hydroxide from pink to colourless React with metals to form salt and hydrogen gas Reacts with carbonates to form salt, water and carbon dioxide. Reacts with alkanols to produce esters
Alkanals (aldehydes)	<ul style="list-style-type: none"> Soluble in water. Large molecules of alkanals are insoluble. 	<ul style="list-style-type: none"> Is readily oxidized to red precipitate in Fehling's solution.
Alkanones (ketones)	<ul style="list-style-type: none"> Soluble in water. Large molecules of alkanones are insoluble. 	<ul style="list-style-type: none"> Is readily oxidized to red precipitate in 2,4 DNPH

IDENTIFYING ORGANIC COMPOUNDS USING TESTS

Identification of carboxylic acids

☞ Carboxylic acids, like other acids affect the colours of indicators, such as litmus and phenolphthalein. This property is used to identify carboxylic acids and is called the acid test.

Method

- 1) Put 15 drops of sodium hydroxide solution in a test tube.
- 2) Add 2 drops of phenolphthalein indicator.
- 3) Then add about 6 drops of the test liquid.

Expected results

When the pink colour in a test tube disappears, the liquid is a carboxylic acid.

Identification of alkenes

☞ Alkenes are very reactive hydrocarbons because of their carbon double bond carbon (C=C) functional group. They react readily with halogens. For example, Hexane reacts with

bromine to form dibromohexane. This reaction of alkenes with bromine is used to identify alkenes. It is called bromine test.

Method

- 1) Put 15 drops of bromine solution in a test tube.
- 2) Add 2 drops of the test liquid.

Expected results

If the test liquid is an alkane, the bromine solution changes its colour from brown to colourless

Identification of aldehydes (alkanals)

☞ Aldehydes contain the carbonyl group in their formula. They react with Fehling's solution (blue in colour) and produce a red precipitate. The felling's test, therefore is used to identify the aldehydes.

Method

- 1) Put 15 drops of felling's solution in a test tube.
- 2) Add about 15 drops of a test liquid.
- 3) Heat for a minute.

Expected results

When the test liquid is an aldehyde, a red precipitate will be produced in a test tube during heating.

Identification of alkanones (ketones)

☞ Ketones, like aldehydes have a carbonyl group ($C=O$) in their structure. But this carbonyl is not at the end of a carbon skeleton as is the case with aldehydes. Ketones do not react with felling's solution. So it cannot be used to identify them.

☞ They can be identified by using a Brady's test. In this test, an orange powder of 2,4-DNPH is used. Both aldehydes and ketones give a positive result. A red precipitate is produced when either an aldehyde or a ketone is mixed with this powder.

Method

- 1) Place a match-head amount of 2,4-DNPH powder in a test tube.
- 2) Add 2 to 3 drops of ethanol to moisten the powder in a test tube.
- 3) Add 3 drops of concentrated sulphuric acid. A yellow liquid is formed.
- 4) Then add 1 drop of a test liquid.

Expected results

When the test liquid is a ketone, a red precipitate will be formed.

Identification of soluble organic compounds

❖ The common physical property of organic compounds is their solubility in water. When an organic compound is added to water in a test tube, if there is only one layer in it, then the compound is soluble. If there are two layers then it is insoluble.

Method

- 4) Put about 15 of water in a test tube.
- 5) Add 2 to 3 drops of the test liquid.

Expected results

- When the test liquid is soluble in water, there will be only one layer in the test tube.
- When the test liquid is insoluble in water, two layers are formed.

IDENTIFYING ORGANIC COMPOUNDS USING FLOW DIAGRAM

- ❖ A flow diagram is like a guide a chemist follows in order to identify different organic compounds.
- ❖ A flow diagram is a kind of a map a chemist needs to follow in order to analyse unknown organic compounds. To analyze means to separate and detect.
- ❖ A flow diagram is established by carrying out tests one after the other. The tests carried out are; solubility test, acid test, Fehling's, Brady's and bromine on the available families to identify.

How to work out a flow diagram

The names of the compounds to be identified are written on the left side of the flow diagram, as in the figure below 1. You then carry out the following steps to test one property after the other;

Step 1: Solubility test

- ❖ This is the first test that is carried out. This test will enable you to separate soluble compounds from insoluble ones. The soluble compounds give one layer when mixed with water (positive result), however two layers means the compound is insoluble (negative result). The soluble compounds are written on the top of the flow diagram while the insoluble ones are written on the bottom.

Step 2: acid test

- ❖ This test usually comes after solubility test. It identifies a carboxylic acid from the soluble compounds separated in the first step. The organic compounds that give negative results are written down and identified further.

Step 3: Fehling's test

- ❖ This test identifies aldehydes. It can come after the acid test. The positive result of this test is the formation of the red precipitate when the aldehyde is oxidized.

Step 4: Brady's test

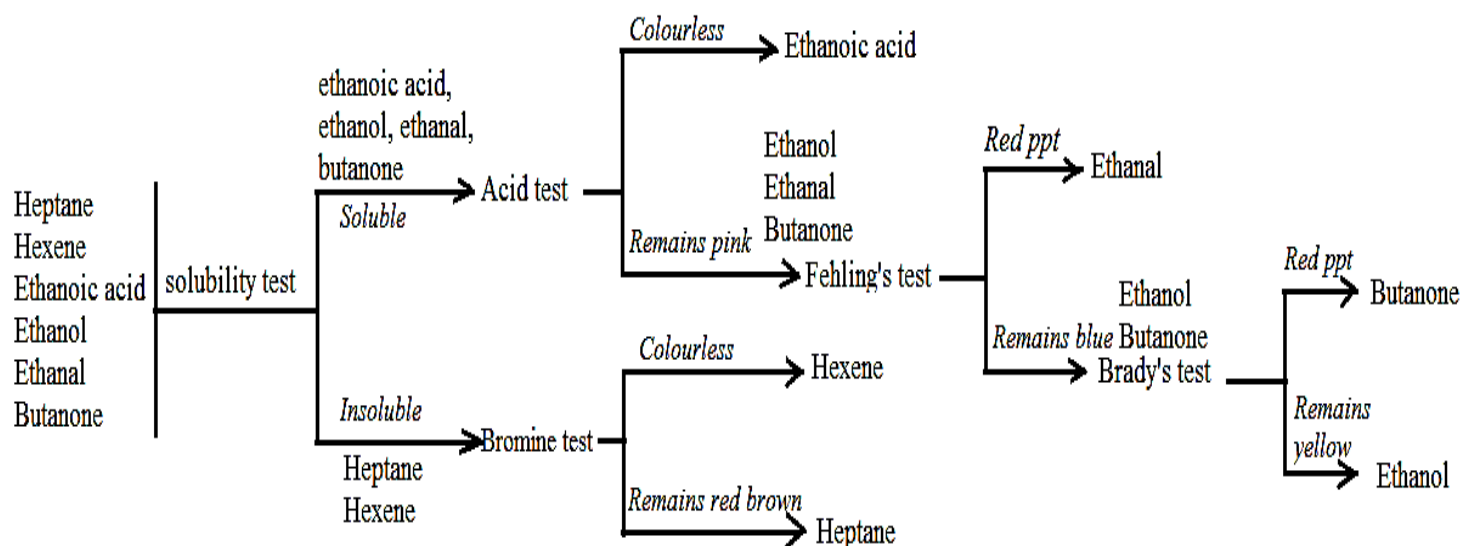
Brady's test can come after Fehling's test. Brady's test is used to identify ketones.

Step 5: Bromine test

The insoluble liquids that were separated in the solubility test can now be identified using bromine test. Usually this test is used to distinguish an alkene from an alkane

Example 1

You are provided with 6 beakers that are not labeled. Each of the beaker contains one of the following compounds: heptane (alkane), Hexene (alkene), ethanoic acid (carboxylic acid), ethanol (alkanol), propanol (aldehyde) and butanone (ketone). Draw the flow diagram that you can use to identify the compounds.



Note:

- 1) The kind of test you start with and the number of tests you carry out depend largely upon what homologous series of organic compounds you have. In example 1, the first test was solubility because some organic compounds are soluble while others are insoluble.
- 2) The tests carried out included solubility, acid, Fehling's, Brady's and bromine tests because the organic compounds to be identified were so many and required all these tests.

Example 1

Your chemistry teacher gave you three unlabeled beakers; each containing hexene, hexane and ethanoic acid respectively. Draw a flow diagram you can use to identify them.

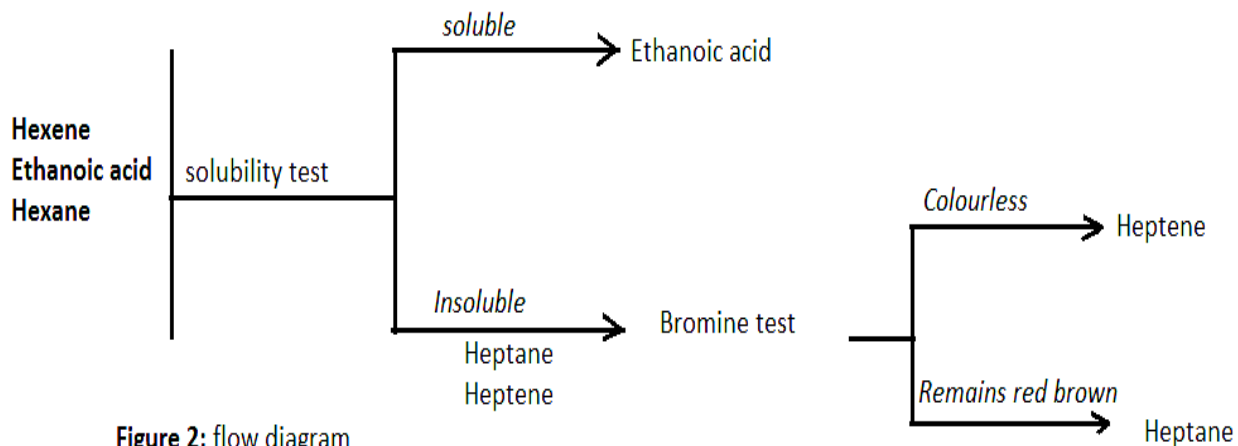


Figure 2; flow diagram

- ❖ In the second example, there are only two tests that can be carried out. The first test is solubility test. This is because of the few organic compounds given.