

SCHOOL CERTIFICATE

CHEMISTRY 5070/123

GRADE 10 TERM ONE

Introduction;

Chemistry is the study of matter and its properties. The properties are either chemical or physical. Furthermore, chemistry deals with the composition of matter which a gain affects the properties of matter.

Branches of chemistry:

Branches of chemistry include; organic chemistry, inorganic chemistry, physical chemistry, analytical chemistry and environmental chemistry. Each of these branches could be sub-divided further

Importance of chemistry:

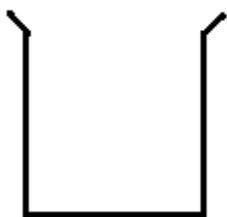
Chemistry affects our lives daily whether we are aware of its effects or not; for example every day we use detergents for cleaning, we cook food, we put salt in relish, we use sauce pans for cooking, we apply fertilizers to our crops in the gardens, we use drugs to treat illnesses, electricity is a product of electrons and other charged particles of matter. Generally chemistry is the key to science and technology.

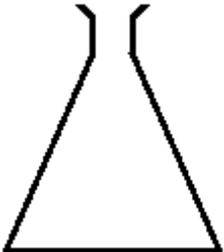
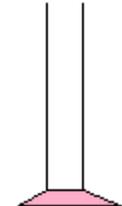
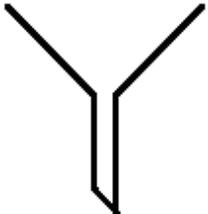
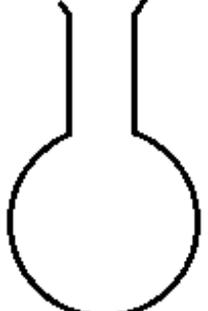
Careers in Chemistry;

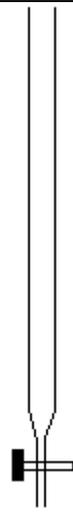
There are many careers in chemistry which include;

- a. Medicine
- b. Pharmacy
- c. Metallurgy
- d. Environmentalist
- e. Petrol chemical industry
- f. Chemical engineering

Some common apparatus in chemistry

Apparatus name	Diagram	Use
Beaker		For holding liquids

Test tube		For holding small amounts of liquids
Evaporation dish		For evaporating solvents
Conical flask		For holding liquids during titration
Measuring cylinder		For measuring volumes of liquids
Funnel		For holding filter paper when filtering solution
Volumetric flask		For making standard solution

<p>Burette</p>		<p>For adding volume of titrant to the titrand into the conical flask</p>
<p>Pipette</p>		<p>For transferring a known volume of liquids to a conical flask</p>
<p>Thermometer</p>		<p>For taking temperature reading</p>

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MATTER AND THE KINETIC THEORY:

Matter is anything that has mass, volume and shape. Matter exists mainly in three states namely; solid, liquid and gas. These three states are made up of tiny particles which could be atoms, molecules or ions. The three states are distinguished on the basis of the arrangement of the particles in each state.

SOLID:

In a solid; the particles are closely packed together such that they are not free to move about. The particles only vibrate (shake) at their fixed position. This is because the particles have very little kinetic energy and very high intermolecular force of attraction. As a result solids have a fixed shape and a fixed volume

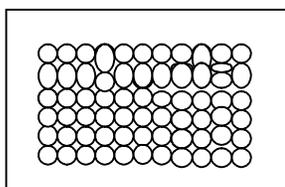
LIQUID:

In a liquid particles are close to each other but are free to move by sliding and falling over each other. This is because the particles in a liquid have a higher kinetic energy compared to the particles in the solid state. Furthermore, the intermolecular forces of attraction between the particles in a liquid are weaker hence the higher degree of movement in liquid. Consequently liquids have a fixed volume but not a fixed shape; instead liquids take up the shape of the container in which they are put

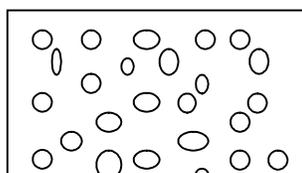
GAS:

In a gas the particles are far apart and are always in a constant random motion. This is because gas particles have a lot of kinetic energy and there is no intermolecular force of attraction between gas particles. Since the gas particles are always in a constant random motion they often collide with each other and the walls of the vessel in which they are put and in their collisions no energy is lost. Since gas particles are far apart; it explains why the atmosphere looks transparent and we see through the air. As a result gases have no fixed shape and no fixed volume. And since the particles in a gas are far apart gases can easily be compressed

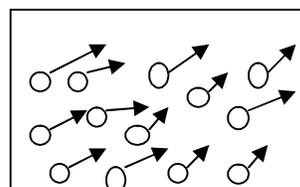
The diagrams below illustrate the arrangement of particles in the three main states of matter



Solid



Liquid

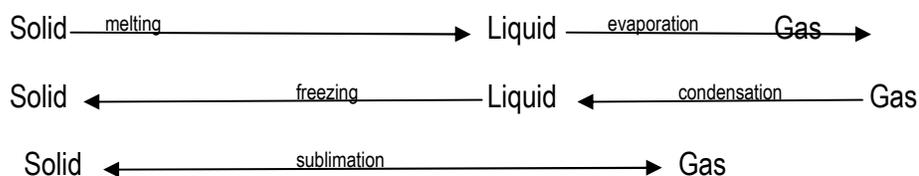


Gas

PHYSICAL AND CHEMICAL CHANGE

Physical change

Matter can undergo either a physical change or a chemical change. A physical change takes place when matter changes state from a solid state to a liquid state or from a liquid state to a gas state. A physical change also takes place when matter changes state from a gas state to a liquid state or from a liquid state to a solid state. There is another type of a physical change this takes place when matter changes state from a solid state to a gas state and from a gas state to a solid state directly without passing through a liquid state. A physical change is therefore defined as process by which matter changes state without producing a new substance and it is reversible. The diagram below illustrates processes involved in physical change:



Chemical change

A chemical change takes place when matter undergoes a permanent change producing a new substance and the process is not reversible. The following are examples of chemical change.

- Rusting of iron
- Decaying of matter
- Burning of wood
- Roasting a chicken

The table below contrasts chemical change from a physical change

Physical change	Chemical change
No new substance is formed	New substance is formed
It is reversible	It is non reversible
Little or no heat is absorbed or released	A lot of heat is absorbed or released
There is no change in weight	There is a change in weight

MELTING AND BOILING POINTS;

Melting point;

A pure solid substance has a fixed melting point. In a pure solid substance, the particles have kinetic energy because they are vibrating. When heat is applied to a pure solid substance the temperature of the substance rises until it reaches a point where it remains constant. It will remain constant until all the solid

particles change into a liquid state. This temperature is called the melting point of the pure substance. In nature there are no two pure substances that share the same melting point i.e. each pure substance has its own melting point.

When heat is being applied to a pure solid substance it is observed that the temperature reading on the thermometer rises until it becomes constant. When it becomes constant some of the particles in a solid that have gained enough kinetic energy change into a liquid state while those that have not gained enough kinetic energy remain part of the solid. The temperature will remain constant while heat is being supplied until all the solid particles change into a liquid state. When the temperature remains constant; the heat being supplied is used to overcome the intermolecular forces of attraction between the particles in the solid. Melting point is therefore defined as the lowest constant temperature at which a pure solid substance change from a solid state into a liquid state.

At melting point two states of matter exist namely the solid state and the liquid state. The melting point for a pure substance is also the freezing point except that in freezing the liquid is changed into a solid. Melting point of a pure solid is affected by dissolved impurities. Dissolved impurities lower the melting point of a pure solid

Experiment 1:

TITLE: Melting point

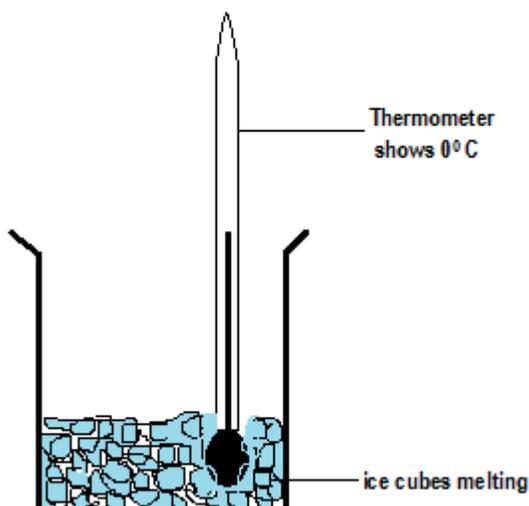
AIM: To determine the melting point of pure ice

APPARATUS: Ordinary mercury in glass thermometer, funnel, conical flask

MATERIALS: pure ice

Method: The reading on the thermometer was taken. Pure ice is placed in the funnel. The funnel is then put onto the conical flask and a thermometer inserted into the ice as shown below.

Diagram:



Observations: The temperature reading on the thermometer was seen to drop from 24°C to 0°C. The temperature remained steady at 0°C until all the ice had changed to liquid water.

Experiment 2:

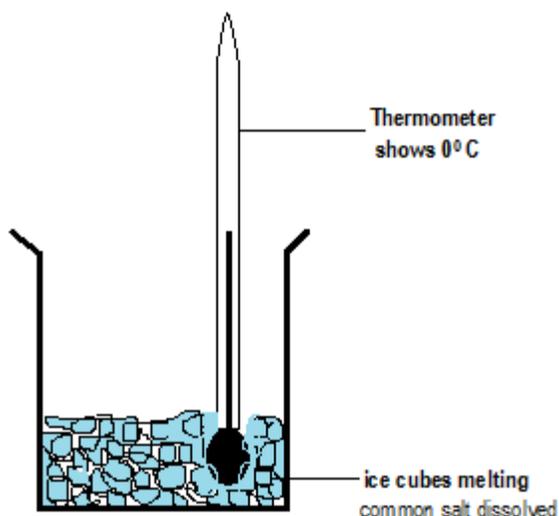
TITLE: Melting point

AIM: To investigate the effect of dissolve impurities on melting point of pure ice

APPARATUS: Two laboratory thermometers, two funnels and two conical flasks

MATERIALS: Pure ice and common salt

Method: The reading on the thermometer was taken. Pure ice is placed in the funnel. The funnel is then put onto the conical flask and a thermometer inserted into the ice. Another similar experiment was set up where table salt was sprinkled on the ice as shown in the diagrams below



Observations: The experiment which had pure ice had its lowest constant temperature at 0°C while the experiment which had salt sprinkled on ice had its lowest constant temperature at -2°C

Conclusion: The impurity (common salt) added to pure ice lowered the melting point of ice. Therefore dissolved impurities lower the melting point of the solid.

Application in life: in cold countries when snow forms; municipal workers sprinkle salt on roads to lower the freezing point of water.

Boiling point;

In a liquid state; the particles already have a much higher kinetic energy than they had in a solid state. However, they do not have enough kinetic energy to get into a gas state. The reasons for this are that there still exists the intermolecular force of attraction between the particles in a liquid. The other reason is due to the atmospheric pressure that exerts a force on top of the liquid surface preventing the particles in a liquid from escaping into the air.

Heating enables the particles in a liquid state to gain more kinetic energy which make them to overcome the intermolecular force of attraction. In addition kinetic energy enables particles to gain more speed and move faster, create more pressure and overcome the atmospheric pressure above the liquid. A liquid begins to boil when bubbles of the gas phase of the liquid begin to form within the liquid and rising to the top while heat is being applied. At this point the liquid's saturation vapour pressure is equal to the atmospheric pressure. The boiling point of a pure substance is the highest constant temperature at which at which the liquid's saturation vapour pressure equals that of the atmospheric pressure. At boiling point two states of matter exist namely the liquid state and the gas state.

The boiling point of a pure substance is affected by dissolved impurities and altitude of the place. Altitude is height above sea level. Dissolved impurities increase the boiling point. At higher altitudes pure liquids boil at a lower temperature due to lower atmospheric pressure while at lower altitudes pure liquids boil at a much higher temperature. For example pure water boils at 100°C at standard temperature and pressure i.e. at sea level. At Lwitikila Girls' Secondary School in Mpika pure water boils at 96°C. On top of a mountain pure water boils at a much lower temperature. Cooking food on top of a mountain poses a challenge as the food may be under cooked. To avoid the food being under cooked mountain climbers carry pressure cookers.

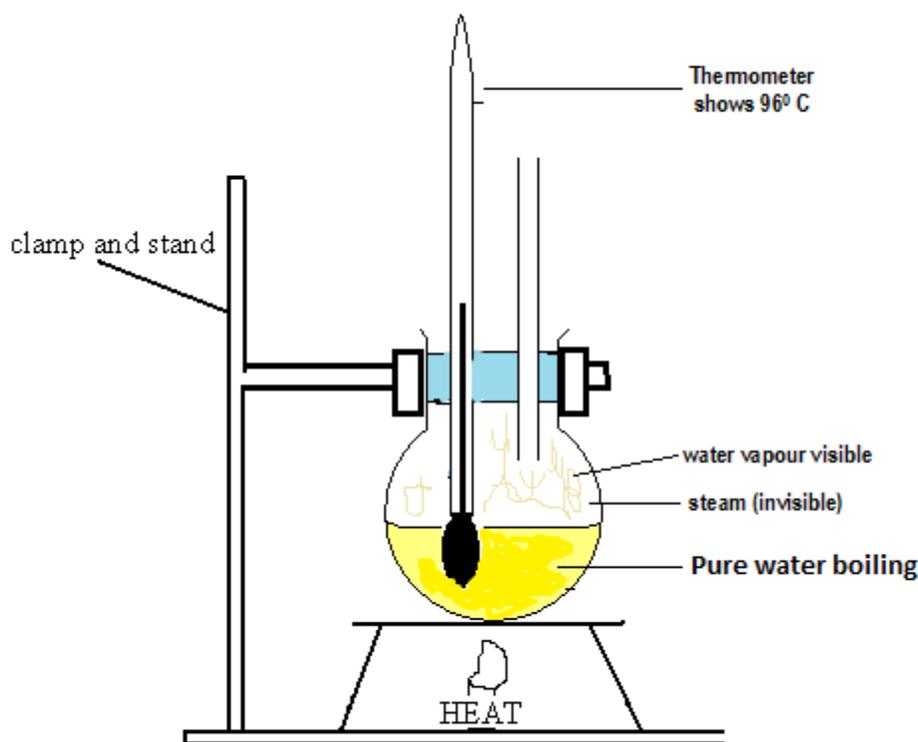
Experiment3: Boiling point

Aim: To determine the boiling point of pure water at Lwitikila Girls' Secondary School

Apparatus: Safety tube, round bottomed flask, laboratory thermometer, Bunsen burner, clamp stand and a rubber stopper

Materials: distilledWater and matches

Method: Distilled Water was put in the round bottomed and clamped. The thermometer and the safety tube were fitted onto the rubber and inserted into the mouth of the flask. The burner was lit and placed under the flask. The experiment was set up as shown in the diagram below.



Observations: When the Bunsen burner flame was placed under the flask; the temperature reading on the thermometer rose steadily from 23°C to 96°C. The temperature remained at 96°C as the water boiled for several minutes. At 96°C steam was seen escaping through the safety valve as the water was boiling.

Conclusion: The boiling point of pure water at Lwitikila Girls' Secondary School was 96°C.

Explanations: Distilled water is used because it is pure water. The safety tube is fitted to remove excess pressure as the water boiled otherwise the flask would burst. The boiling point was 96°C which is less than the 100°C which we know; this is because Lwitikila Girls' Secondary School is at a higher altitude. Water boils at 100°C at standard temperature and pressure (S.T.P.) which is at sea level. The standard temperature is taken to be 0°C and standard pressure is taken to be one (1) atmosphere or 760mmHg

Experiment:

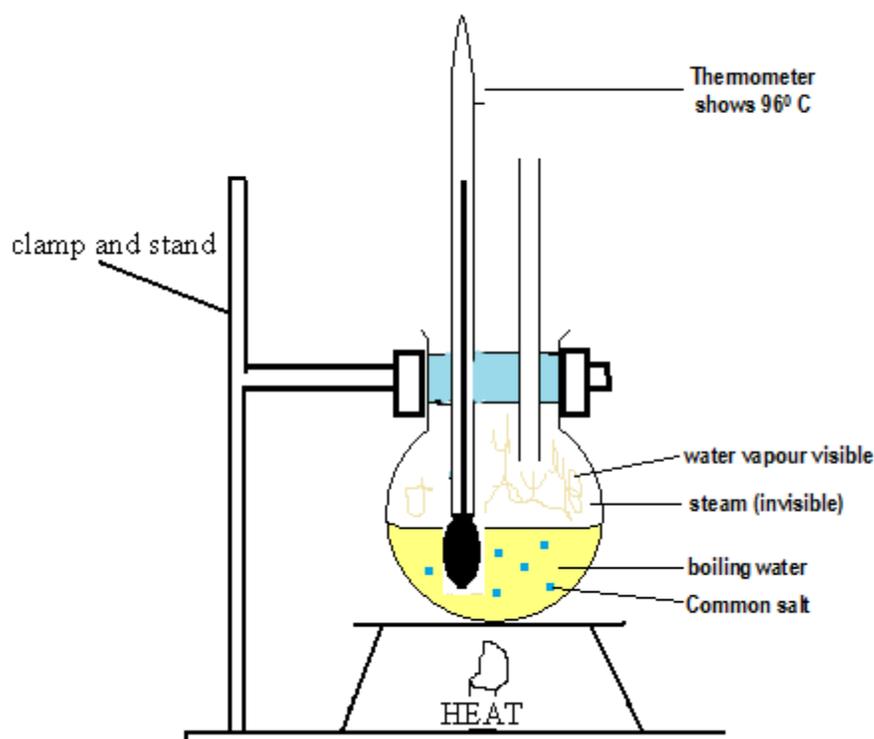
Title: The effect of dissolved impurities on the boiling point.

Aim: To investigate the effect of common salt on the boiling point of water.

Apparatus: Safety tube, round bottomed flask, laboratory thermometer, Bunsen burner, clamp stand and a rubber stopper

Materials: distilled water, table salt and matches

Method: Salt was added to the distilled Water and stirred. The solution was put in the round bottomed and clamped. The thermometer and the safety tube were fitted onto the rubber stopper and inserted into the mouth of the flask. The burner was lit and placed under the flask. The experiment was set up as shown in the diagram below.



Observations: When the Bunsen burner flame was placed under the flask; the temperature reading on the thermometer rose steadily from 24°C to 105°C. The temperature remained at 105°C as the water boiled for several minutes.

Conclusion: The boiling point of water with dissolved impurities at Lwitikila Girls' secondary school was 105°C.

Explanations: Liquids with dissolved impurities boil at much higher temperatures because the impurities absorb some of the heat that is being supplied to the system.

Application in life: We put salt in beans after it has cooked and we put salt in meat just before starting to cook. Explain why?

Experiment:

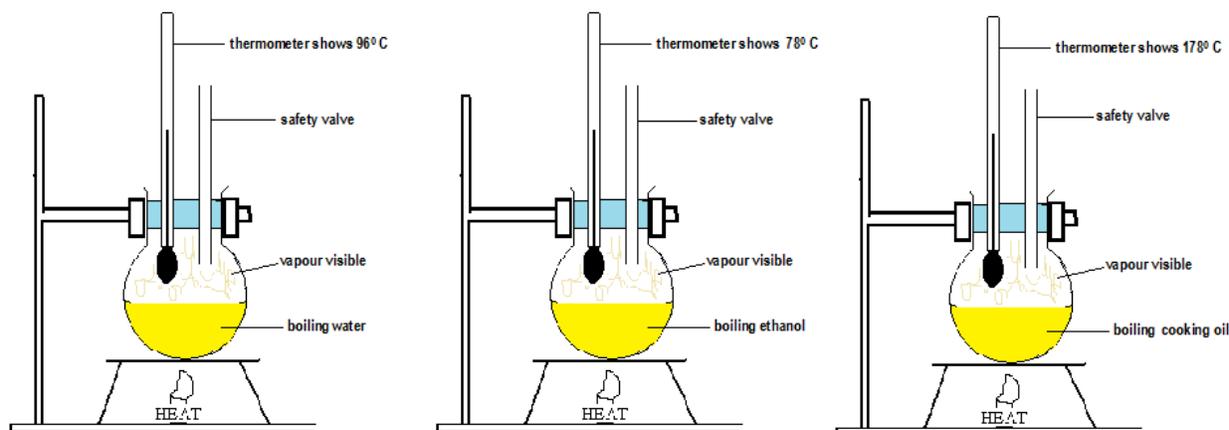
Title: Boiling points of Different pure substances

Aim: To show that different pure liquids have different boiling points

Apparatus: Safety tubes, round bottomed flasks, laboratory thermometers, Bunsen burners, clamp stands and a rubber stoppers

Materials: pure water, cooking oil and ethanol

Method: The experiment was set up as shown in the diagram below



Results; Water boiled at 96°C; ethanol boils at 78°C and cooking oil boiled at 178°C

Conclusions: Each pure substance has its own constant boiling point

EVIDENCE THAT MATTER IS MADE OF TINY PARTICLES OF MATTER:

Matter is made up of tiny particles. In a gas phase these particles are in a continuous random motion. The evidence to show that matter is made up of tiny particles is demonstrated by diffusion and Brownian motion

DIFUSION:

Diffusion is the process by which substances move from their region of higher concentration to the region of their lower concentration along a concentration gradient. The movement continues until the particles are evenly distributed.

Experiment:

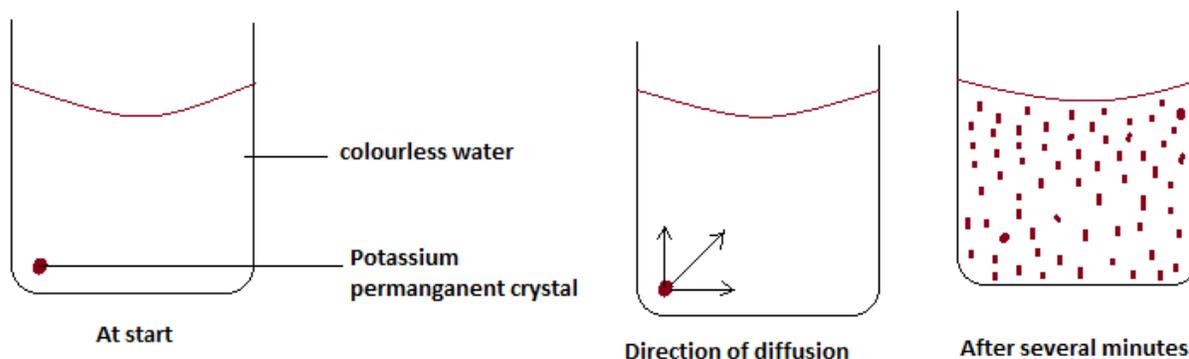
Title: diffusion

Aim: to demonstrate diffusion of potassium permanganate in water

Apparatus: a beaker

Materials: Water and potassium permanganate

Method: Water is put in a beaker and a crystal of potassium permanganate dropped into one corner of the beaker and left to stand for several hours as shown in the diagram below



At the start of the experiment

Observation:

After several hours; the purple colour spread to all parts of the liquid as shown in the diagram above

Factors affecting the rate of diffusion;

The rate of diffusion of any substance is affected by; concentration, temperature, particle size and the molecular mass of the substance.

- a. **Concentration:** concentration is the quantity of matter in a given volume. The higher the concentration in one place, the greater the rate of diffusion.
- b. **Temperature:** Temperature is the degree of hotness. When the temperature of a substance is high the particles have higher kinetic energy and this makes them move faster hence diffuses quickly
- c. **Particle size:** substances with smaller particles diffuses faster than bigger ones.
- d. **Molecular mass:** particles with smaller molecular masses diffuse faster than particles with bigger molecular masses.

Experiment:

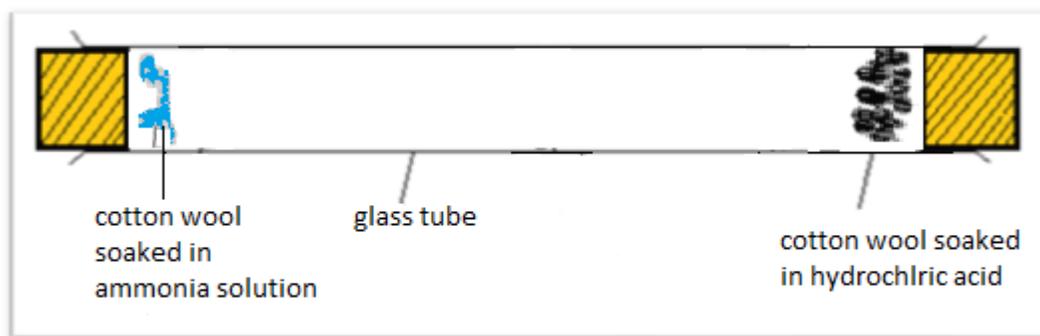
Title: Rate of Diffusion

Aim: To show that substances with smaller molecular mass diffuse faster than those with larger molecular masses

Apparatus: glass tube, corks

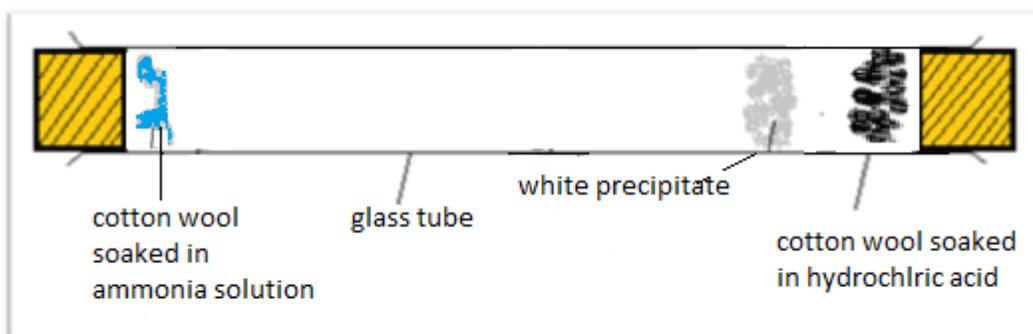
Materials: Hydrochloric acid, ammonia solution and cotton wool

Method: Ammonia is put on the cotton wool and placed on one side of the glass tube and hydrochloric acid is put on the cotton wool and placed on the other side of the tube as shown below:



At start

Observations. After 10 minutes a white fume of ammonium chloride was seen near the hydrochloric acid cotton wool as shown in the diagram below:



After several minutes

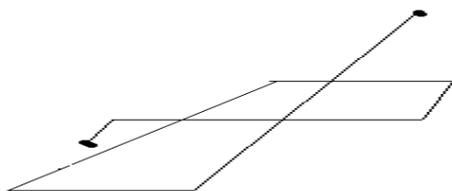
CONCLUSIONS: Ammonia diffuses faster than hydrogen chloride.

EXPLANATIONS: ammonia has a molecular mass of 17g and hydrochloric acid has a molecular mass of 36.5g. The appearance of the white precipitate indicates that ammonia with a molecular mass of 17g diffused faster because it has a smaller mass compared to hydrochloric acid which has a molecular mass of 36.5g. The white precipitate is ammonium chloride which is a result of the reaction between the hydrochloric acid and the ammonia.

Law of diffusion: This law states that the rate of diffusion of a substance is inversely proportion to the square root of its density

BROWNIAN MOTION;

Brownian motion is another evidence to prove that matter is made up of tiny particles and that these particles especially those in liquids and gases are always in a constant random motion. In 1927; Robert Brown was using a microscope to study pollen grains suspended in water, and observed that tiny pollen grains were constantly wondering about. He sketched the movement of pollen grain as shown below;



The actual explanation came some years later when it was suggested that the motion was caused by vibrating water molecules bumping into pollen grains. This motion became known as the Brownian motion.

The other Brownian motion is observed in smoke particles which are always being bumped into by air particles causing them to move in a zigzag motion.

MIXTURES, COMPOUNDS AND ELEMENTS

Matter can exist as mixtures, compounds and elements. There are many substances around us. Some are pure and others impure. It is therefore necessary to classify the substances into various categories. One way to classifying the substances is to classify them as either mixtures, compounds or elements.

A mixture is a substance made up of two or more components physically combined together in any proportion. Examples of mixtures include; air, sugar solution, chlorophyll, soil and blood

A compound is a pure substance made up of two or more substances chemically combined together. Examples of compounds include; water, copper (II) sulphate and sodium hydroxide.

An element is a pure substance made up of atoms with the same atomic number. Examples of elements include carbon, oxygen, calcium and copper. The table below compares a compound, an element and a mixture

Mixture	Compound	Element
Impure substance	Pure substance	Pure substance
components physically combined	components chemically combines	Atoms chemically combined
components are in any ratios	components are in fixed ratios	-

little or no heat is absorbed or released during formation of mixture	Heat either absorbed or released when compounds form	-
Can be separated by physical methods	can be separated by chemical methods	can not be separated into smaller units by any known chemical methods

CRITERIA FOR PURITY

Any pure substance must conform to the set standards. In chemistry the standards are set by the International Union of Pure and Applied Chemistry (IUPAC). The set standards include; melting point, boiling point, refractive index and density. There are no two pure substances that share the same melting point, boiling point, refractive index and density. For example;

- a. The boiling point of water is 100°C at STP
- b. The boiling point of ethanol is 78°C at STP
- c. The melting point of naphthalene is 80°C
- d. The melting point for ice is 0 °C
- e. The melting point of naphthalene is 80°C
- f. The density of water is 1000kg/m³
- g. The density of iron is 7800kg/m³
- h. The density of human flesh is 1070kg/m³
- i. The density of gold is 19300kg/m³
- j. The refractive index for glass is 1.5
- k. The refractive index for water is 1.33

MELTING POINT AND BOILING POINT

Melting point can be used to find out whether the substance is pure or not if the melting point of a substance is less than what is recommended at STP then the substance is not pure

Experiment:

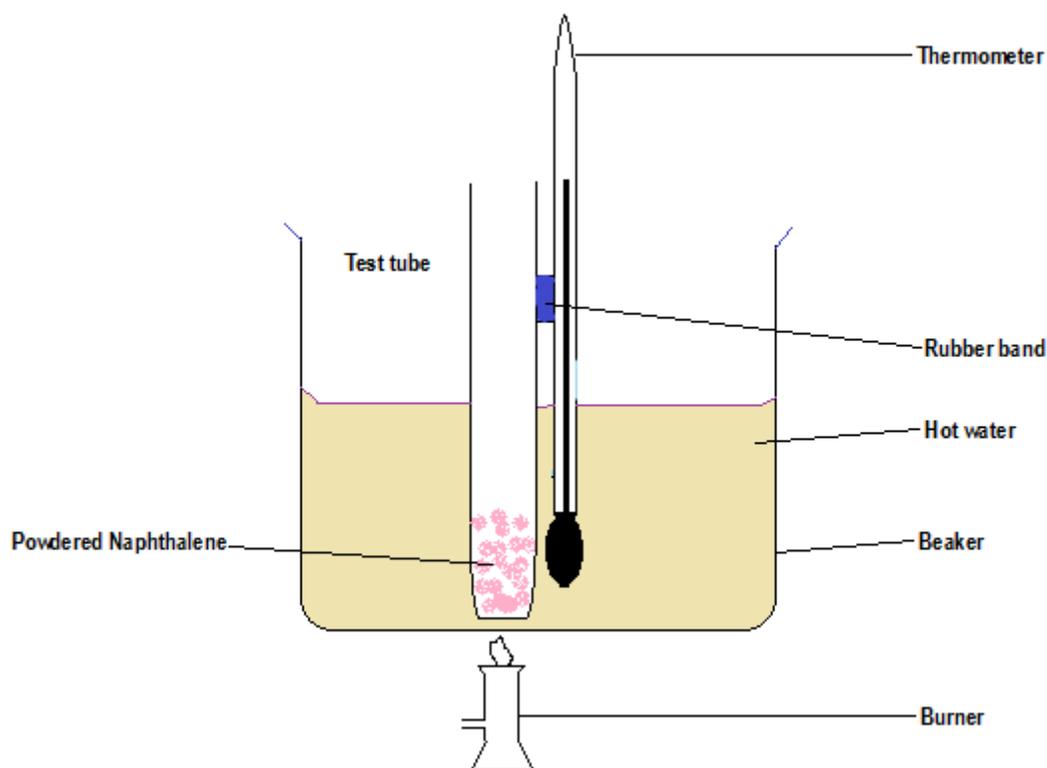
Title: Melting point

Aim: To determine the melting point of naphthalene

Apparatus: Beaker, thermometer, glass tube, Bunsen burner, and rubber band

Material: water and naphthalene

Method: Solid or powdered naphthalene is put into a boiling tube the boiling tube is then attached to the bulb of the thermometer by a rubber band. The capillary tube and the thermometer are partially immersed in water which is rapidly heated. The diagram below shows how the experiment was set up.



Observations: As the water was being rapidly heated, the temperature on the thermometer rose from 28°C to 46°C, to 62°C to 71°C and came to remain steady at 80°C. At 80 °C naphthalene started to melt. The temperature remained constant at 80°C until all the naphthalene had melted. When all the naphthalene had melted the temperature started to rise again.

Conclusion: The melting point of naphthalene was found to be 80°C

Experiment:

Title: Solidifying point:

Aim: to determine the temperature at which Naphthalene changes to a solid state by graphical method.

Apparatus: Thermometer, capillary glass tube, beaker and rubber band

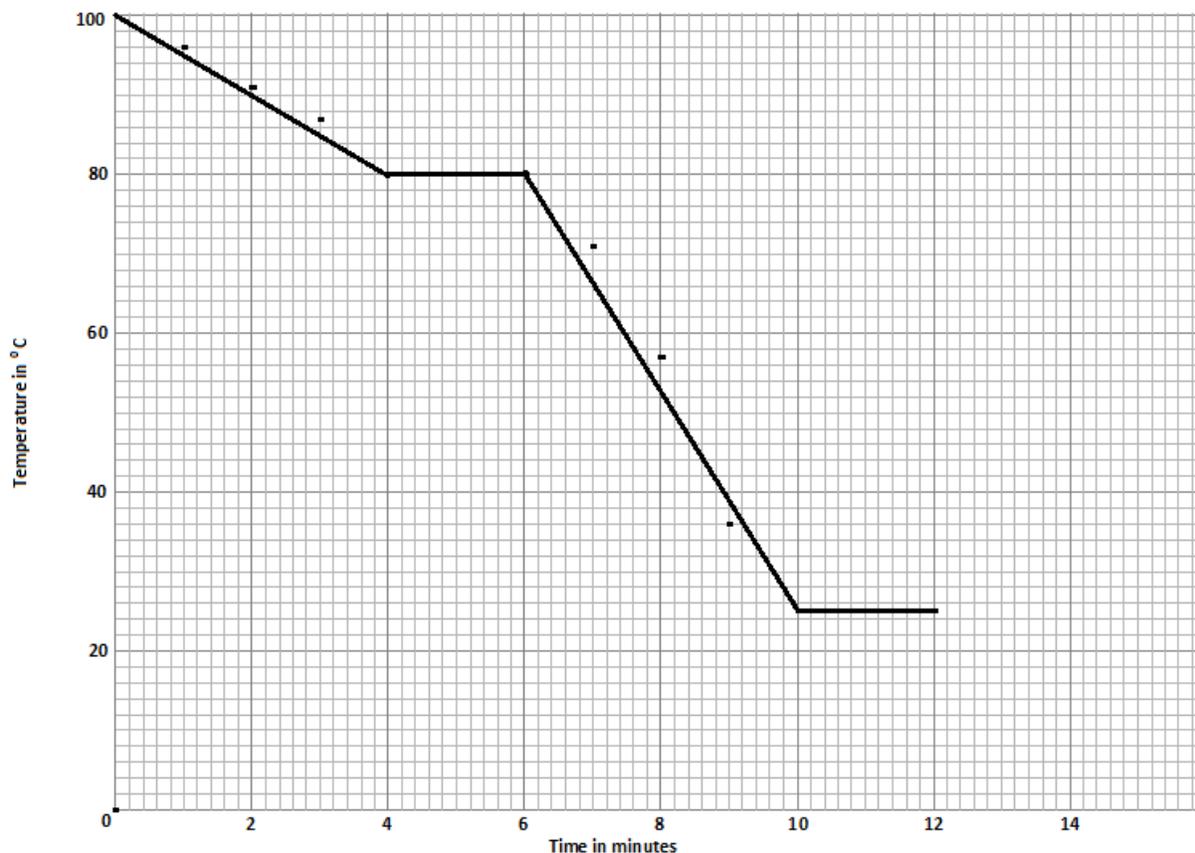
Materials: Naphthalene at 100°C

Method: The liquid naphthalene at 100°C was allowed to cool slowly and the temperature reading was taken at one (1) minute interval the temperature and time readings were recorded.

Results: the result from the above experiment were recorded as shown as shown in the table below

Temp in °C	100	96	91	87	80	80	80	71	56	36	25	25	25
Time in Minutes	0	1	2	3	4	5	6	7	8	9	10	11	12

From the above table a graph of temperature against time was plotted as shown below



Conclusion: The melting point of pure naphthalene is 80°C and the temperature at which pure naphthalene liquid changes from liquid state to solid state is 80°C. If the melting point was less than 80°C then we would conclude that the naphthalene was not pure. From the graph when the temperature reading on the thermometer remained constant or when the line became horizontal at 80°C there was a change of state from liquid state to solid state.

Explanations: From the graph it is observed also that the temperature continued to drop after the liquid naphthalene had completely changed into a solid state. This was due to the fact that the temperature at which naphthalene changed state from the liquid state to the solid state was much higher than the room temperature. Therefore the temperature was dropping to get to room temperature.

SEPARATION OF MIXTURES (separation techniques).

Since components in a mixture are physically combined they can be separated into their pure components by physical methods or physical techniques. Usually, when separating mixtures a combination of techniques are employed. These techniques include:

- i. Filtration
- ii. Evaporation
- iii. Distillation
- iv. Crystallization
- v. Chromatography
- vi. Fractional distillation
- vii. Decantation
- viii. Flootation
- ix. Magnetism
- x. Sedimentation
- xi. Centrifuge

.DISSOLUTION, FILTRATION, EVAPORATION AND CRYSTALLIZATION

Usually the four above techniques are combined to separate components in certain mixtures

Experiment:

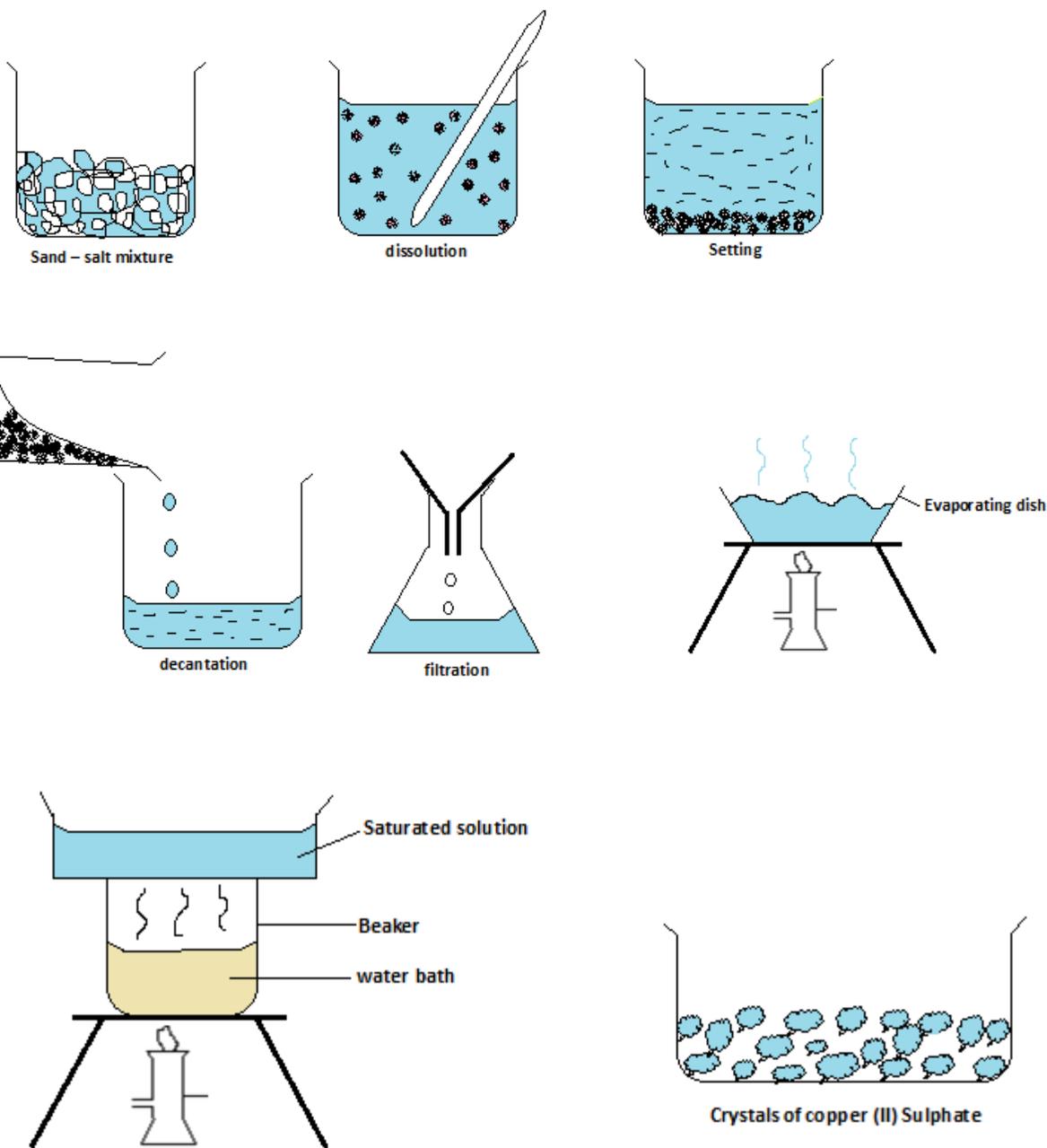
Title: Separation technique

Aim: To separate a mixture of sand and copper (II) sulphate

Apparatus: Bunsen burner, two beak, conical flask, funnel, stirring rod tripod stand, wire gauze, evaporation dish.

Materials: copper (II) sulphate, sand, water, matches and filter papers

Method: sand and copper (II) sulphate mixture was put in a beaker. Water was added to the mixture and stirred strongly using a stirring rod. After stirring the mixture was allowed to settle. After settling the mixture was decanted. The decanted solution was then filtered. The filtrate was put in an evaporation dish and heated strongly at first then gently using a water bath. The saturated solution was removed from the water bath and allowed to cool slowly. The crystals were then removed and dried. The diagrams below show the stages



Observations:

when water was added to the sand- copper sulphate mixture the mixture appeared dirt- bluish. When the mixture was allowed to settle, dense solid particles collected at the bottom. When the mixture was decanted the decanted solution contained some suspended impurities. When the solution was filtered the filtrate was clear and blue in colour. Some solid particles remained on the filter paper as residues. When the filtrate was heated strongly the excess water evaporated leaving a saturated solution. When the hot saturated solution was left to cool slowly over several hours blue crystals of copper (II) sulphate formed in the evaporation dish.

Conclusion:

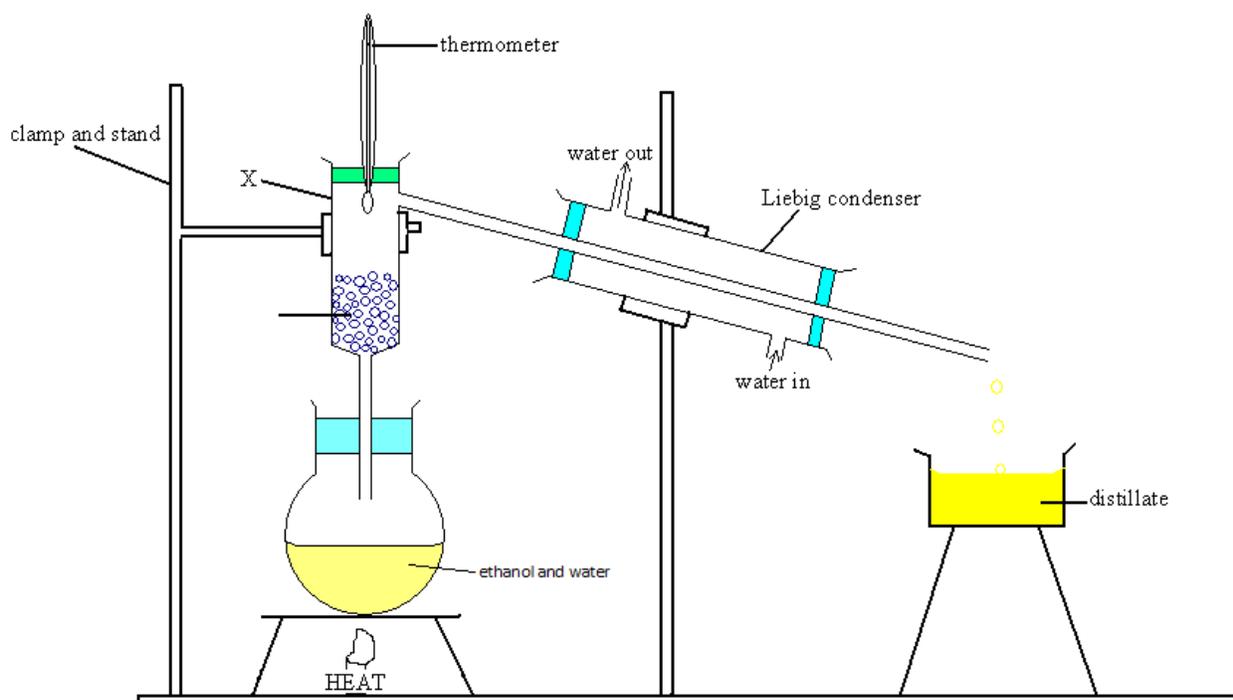
The mixture of copper (II) sulphate and sand can be separated by a combination of the following techniques; dissolution, decantation, filtration, evaporation and crystallization.

Explanations:

water was added to the mixture to dissolve copper (II) sulphate because copper (II) sulphate is soluble in water while sand is insoluble in water. Stirring was done to speed up the dissolution of copper (II) sulphate. The mixture was allowed to stand for some time so that the dense insoluble particles of sand could settle at the bottom. Decantation was done to separate the dense sand particles from the liquid and the suspended impurities. Filtration was done to separate the suspended impurities from the liquid; this is because the pores of the filter paper are too small to allow larger particles to pass through but big enough to allow water molecules, copper (II) ions and sulphate ions to pass through. The filtrate was heated strongly to drive out excess water. The solution was heated gently to avoid destroying the shape of the crystals. The saturated solution was allowed to cool slowly so that larger crystals are formed and to have good geometrical shape.

FRACTIONAL DISTILLATION

Fractional distillation is a separation technique that is used to separate two or more liquids which mix uniformly i.e. miscible liquids such as alcohol and water, air and crude oil. Components in the mixture can be separated by fractional distillation because they have different boiling points. The apparatus for fractional distillation is shown below:



Some points on the apparatus;

1. The thermometer is used to check the boiling point for the fraction being distilled
2. The thermometer is used to check the temperature of the fraction being distilled
3. The glass beads condense the fractions whose boiling point has not yet reached so that they drop back into the distillation flask.
4. The fractionating column offers a large surface area for condensing less volatile vapours.

5. The Liebig condenser holds a cold solvent e.g. water to condense the water back into the liquid state.
6. A coolant e.g. water moves into the condenser from the lower end and out of the condenser from the upper end so that it can sufficiently cool the vapours.
7. In case of azeotropic mixtures; it is not possible to separate components that have close boiling points without having traces of one component in the other

CHROMATOGRAPHY

Chromatography is a separation technique used to separate complex mixtures such as dyes in ink, colours of chlorophyll, amino acids in proteins, monosaccharide in starch, poisons in blood etc. chromatography separates components in a mixture because components have different solubility in a given solvent in which they dissolve. The solvent in which the components or solutes dissolve is called the mobile phase. The material upon which the mobile phase moves and upon which the separation of the components takes place is called the stationary phase. Stationary phase absorbs the solutes due to its capillarity property. The solutes stick to the stationary phase at different intervals due to their density and solubility in the solvent the less soluble solute sticks closer to the starting point while the more soluble ones stick near the end of the stationary phase. The band of colours that is developed as a result of the solutes sticking to the stationary phase is called chromatogram. If the stationary phase is a paper then the chromatography is described as paper chromatography and if the separation takes place on a solid it is described as solid chromatography.

Experiment:

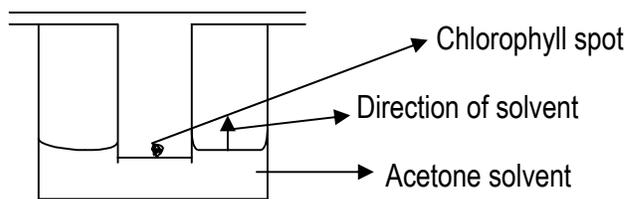
Title: Paper chromatography:

Aim: to separate and identify the components of chlorophyll

Apparatus: Means of pounding leaves, small beakers, thin dropper,

Materials: Fresh green leaves of tomato, acetone, filter paper,

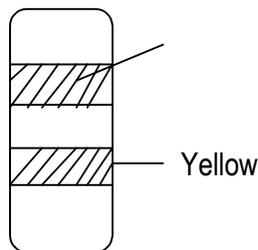
Method: the leaves were pounded and chlorophyll squeezed out into a beaker. The chlorophyll was concentrated by removing excess water. A thick concentrated spot of chlorophyll was made on the strip of the filter paper. Acetone was put in a beaker and the end of the strip was made to just touch the acetone with the spot of chlorophyll just a few millimeters above the acetone. As shown in the diagram below



Observation:

As the acetone solvent moved up the strip, two colours namely yellow and blue were seen on the strip, yellow colour was closer to the solvent and blue colour was above the yellow colour as shown in the diagram below

blue



Conclusion:

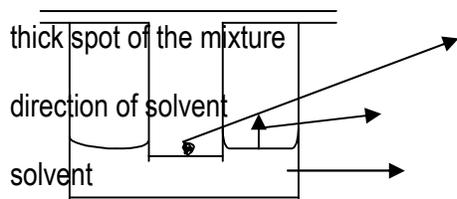
Chlorophyll is made up of a blue pigment and a yellow pigment. This also shows that chlorophyll is not a pure pigment

Explanation:

the solvent acetone moved upwards along the filter paper strip by capillarity action. The most soluble solute of chlorophyll which is blue dissolves in acetone most readily and it was carried the furthest and its colour appeared near the end. The less soluble solute which is yellow formed closer where the spot was placed. This experiment shows that chlorophyll is made up of mainly the colours from the blue region and the yellow-red region of the visible light. This experiment also shows that photosynthesis is most active in the two regions of the visible light.

Methods of carrying out paper chromatography

a. Ascending method;



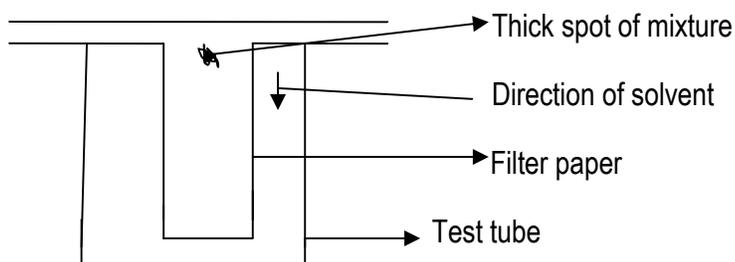
Advantage;

Colours are clearly seen and distinct

Disadvantage;

It takes time for colours to be seen as solvents move against gravity.

b. Descending method;



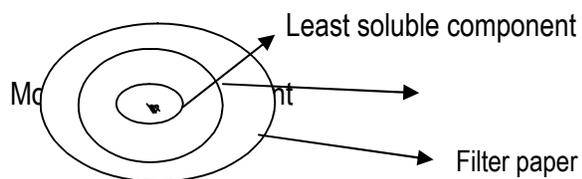
Advantage;

Colours form fast due to the pull of gravity

Disadvantage;

Component colours may overlap making it difficult to distinguish them.

Radial method;

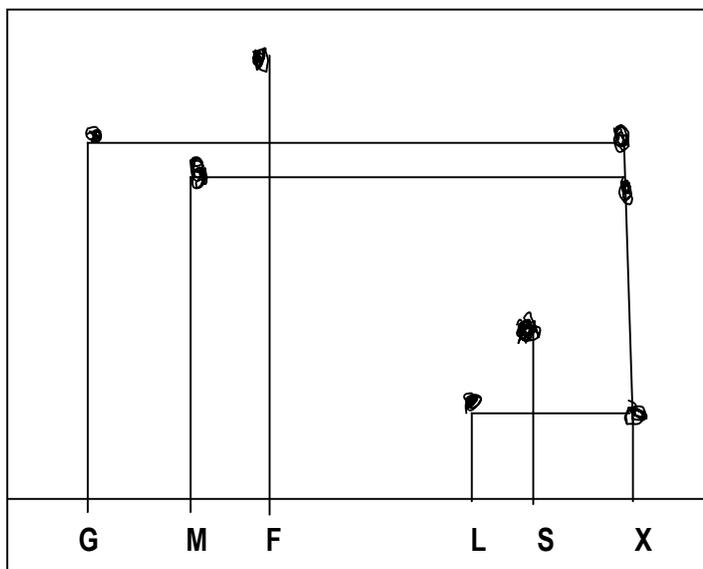


Uses of chromatography;

1. Used to separate and identify monomers in foods
2. Used to detect and identify poison in blood
3. Used to detect and identify various inks used in forged documents.

INTERPRETING CHROMATOGRAMS

Several different solutions are already plotted on a standard chromatogram. These solutions are known and their solutes have fixed position. The unknown solution to be analyzed is made to flow on the stationary phase. The solutes in the unknown solution stick to the stationary phase different positions. These positions are then compared with the positions of the known solutes. For example the solutes in the unknown solution X can be separated, analyzed and identified by carrying out chromatography and the position to which the solutes will stick on the chromatogram will be compared with the position of the already known solutes. The diagram below shows how this can be done



KEY:

- G → Galactose
- M → Maltose
- F → Fructose
- L → Lactose
- S → Sucrose

The components in the mixture X are:-

- (1) Lactose
- (2) Maltose
- (3) Galactose

IMPORTANCE OF PURITY IN PRODUCTS;

Manufacturers of various products such as pharmacists manufacturing drugs, soap manufacturers and manufacturers of food stuffs have to ensure that their products meet certain set standards of purity. Impurities in products may lead to consumers experiencing an unpleasant reaction. For example impurities in drugs may lead to some users developing rash.

LANGUAGE OF CHEMISTRY

a. Word equations:

A word equation is a way of describing a chemical reaction where the reactants are on the left hand side and the products on the right hand side separated by an arrow. For example the reaction between calcium metal and water is written as follows;



Word equations have disadvantages which include the following;

- a. They do not tell us the amounts of both the reactants and the products.
- b. The writings are too long and some people are forced to abbreviate.

CHEMICAL FORMULAE AND CHEMICAL EQUATIONS

When writing reactants and products, there is often a need to abbreviate the names of the reactants and products. In chemistry these abbreviations must be the ones that have been internationally accepted so that any scientist is able to interpret what has been written. In chemistry each element has been given its own unique symbol that is used to identify it. The symbol is known as the chemical symbol or chemical formula. The chemical symbol of the element is its first letter of its English or Latin name. The letter is always written as a capital letter. However, there are 104 elements known against only 26 letters of the English alphabet. Consequently some elements have two letters for their symbols and also to distinguish them from other elements that have the same starting letter. The second letter is always a small letter.

The elements are usually named:

- a. according to the place where it was first discovered.
- b. According to its chemical properties
- c. After the person who discovered it

The table below gives names and symbols of the common elements

Name of the element	Symbol
Lithium	Li
Hydrogen	H
Boron	B
Carbon	C
Oxygen	O
Fluorine	F
Sodium (Natrium)	Na
Magnesium	Mg
Aluminium	Al
Phosphorous	P
Chlorine	Cl
Potassium (Kalium)	K
Calcium	Ca
Zinc	Zn
Copper (Cuprum)	Cu
Silver (Argentum)	Ag
Iron (Ferrous)	Fe
Nitrogen	N
Helium	He
Gold	Au
Lead (plumbium)	Pb

Argon	Ar
Iodine	I

ATOMS, MOLECULES AND RADICALS:

ATOMS;

An atom is the smallest indivisible neutral particle of matter that can take part in a chemical reaction. In writing the symbol of an element on its own represents one atom of that element. E.g. Al, O, Cl, C, Cu etc. All these represent one atom of their respective elements. A symbol of an element with a number in front of the symbol, the number represents how many atoms of the element are there e.g.

Al \longrightarrow one atom of aluminium

3O \longrightarrow three atoms of oxygen

It is important to note that these atoms are not chemically combined but that they are independent of each other.

MOLECULES;

A molecule is the smallest particle of matter that can exist independently in a free state. Molecules are made up of atoms of non metals. There are two types of molecules and these are the molecules of non metal elements and the molecules of non metal compounds. Molecules of non metal elements are distinguished from atoms because they carry a small subscript number below at the back of the symbol. For example molecules of non metal elements are written as follows; H₂, O₂, S₈, Cl₂, I₂.

The above examples show one molecule of the respective element. The subscript number means that the atoms are chemically combined together.

Molecules of non metal compounds are formed as a result of chemical reactions between nonmetals.

Examples of molecules of nonmetal compounds include; H₂O, CO₂, NH₃, CH₄ and HCl.

The number in front of a molecule represents the total number of molecules in that chemical formula. For example

i. 3CO₂ \longrightarrow 3 molecules of carbon dioxide

ii. 2O₂ \longrightarrow 2 molecules of oxygen

iii. 4H₂O \longrightarrow 4 molecules of water

The total number of atoms in a molecule can be calculated by adding up the number of the individual atoms. For example

a. Cl₂ \longrightarrow 2 atoms of chlorine chemically combined

b. H₂O \longrightarrow 2 atoms of hydrogen and 1 atom of oxygen giving a total of 3 atoms chemically combined

c. 3H₂O \longrightarrow 3 molecules of water not chemically combined together, each molecule is made up of 2 atoms of hydrogen and 1 atom of oxygen chemically combined together.

RADICALS;

A radical is a group of atoms that exist together and cannot exist independently. The table below shows some common radicals.

Radical name	Radical symbol
Hydroxide	OH
Sulphate	SO ₄
Nitrate	NO ₃
Carbonate	CO ₃
Phosphate	PO ₄
Ammonium	NH ₄
Hydrogen carbonate	H ₂ CO ₃

VALENCY

Valency is a combining power of an element or radical. In other words when elements form compounds, they do so in certain fixed ratios. The table below gives the valencies of some common elements and radicals.

Element	Symbol	Valency
Metals		
Sodium	Na	1
Potassium	K	1
Magnesium	Mg	2
Calcium	Ca	2
Zinc	Zn	2
Aluminium	Al	3
Copper	Cu	1 and 2
Iron	Fe	2 and 3
Silver	Ag	1
Lead	Pb	2
Barium	Ba	2
Nonmetals		
Hydrogen	H	1
Chloride	Cl ⁻	1
Fluoride	F ⁻	1
Iodide	I ⁻	1
Oxygen	O	2
Sulphur	S	2
Nitrogen	N	3
Carbon	C	4
Phosphorous	P	3
Silicon	Si	4
Radicals		
Hydroxide	OH	1
Nitrate	NO ₃	1
Ammonium	NH ₄	1
Chloride	Cl ⁻	1
Oxide	O ²⁻	2
Sulphate	SO ₄	2
Carbonate	CO ₃	2
Phosphate	PO ₄	3
Hydrogen carbonate	H ₂ CO ₃	1
Ammonia	NH ₃	1

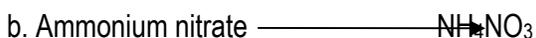
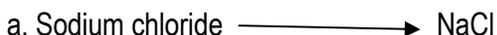
Chemical formulae

A chemical formula is a short hand way of representing a chemical compound, or a molecule or an atom. A chemical compound is derived from a chemical combination either between a metal and a nonmetal, or a metal and a radical or between two nonmetals. Usually we start with the symbol of a metal and end with the symbol of a nonmetal or radical. Before a chemical formula of the compound is written, certain rules must be considered. The rules include the following;

1. Always start with the symbol of the metal, or hydrogen or ammonium and end with the symbol of a non metal or radical e.g.



2. Write the correct symbols of the elements or element and radicals combined in the compound e.g.



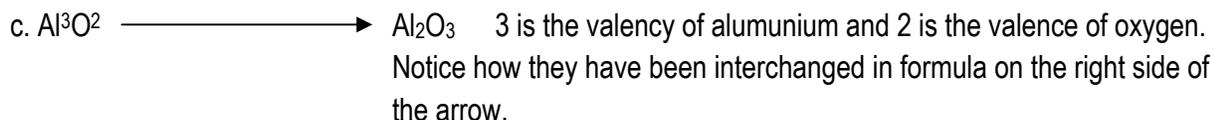
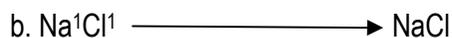
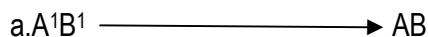
3. Write the valency number for each element or radical on the top of the respective element or radical. For example if element A and B were the two elements in the compound then the two elements and their radicals would be indicated as follows;

A^1 and B^3 the meaning is that element A has a valency of 1 and element B has a valency of 3.

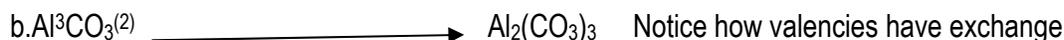
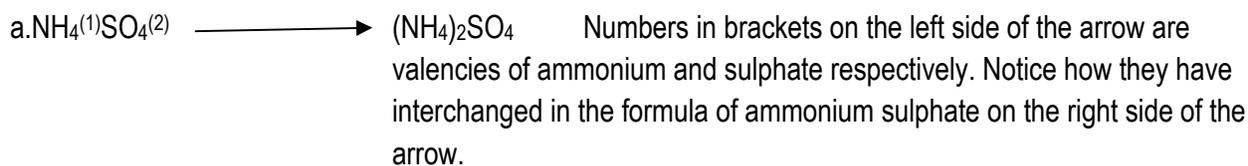
4. If the valency numbers for the two elements or an element and a radical can be divided by a common factor to give the simplest ratio, they must be simplified to the simplest ratio as shown in the examples below:



5. Do not divide the valency numbers if they do not have a common factor that can divide into both without leaving a remainder.
6. Interchange the valencies of the two chemical symbols involved and write the correct chemical formula of the compound with the interchanged valency numbers as small subscript numbers behind the chemical symbols as shown in the examples below:



7. If the compound involves an element and a radical or two radicals and their valences cannot be simplified into 1 to 1 ratio; radicals made up of two elements must be enclosed into brackets and then exchanged valencies written outside the brackets as shown in the examples below.



EXERCISE

1. Write the correct chemical formula for the following;

- i. Copper (II) nitrate
- ii. Aluminium carbonate
- iii. Calcium chloride
- iv. Aluminium oxide

2. Give the names of the following chemical formulae;

- i. $Cu(NO_3)_2$
- ii. $Mg(OH)_2$
- iii. CO_2
- iv. $NaCl$
- v. $Fe_2(SO_4)_3$

Balancing chemical equations: (law of conservation of matter)

When a chemical reaction takes place there is re-arrangement of atoms on the product side. The total number of atoms on the reactant side and product side must always be equal. The same atoms that were

on the reactant side must also appear on the product side. This is the law of conservation of matter i.e. matter can neither be created nor destroyed. Because of this law of conservation of matter, any chemical equation must be balanced. Balancing chemical equation is done by writing whole numbers greater than one (1) in front of the chemical formula of either the element or the compound. We can check if the equation is balanced by counting the number of atoms of each element on the reactant side and the product side. If the numbers are equal on both sides then the equation is balanced. Examples below show some balanced equations;

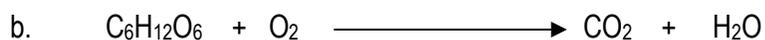


un balanced



Elements	Reactant side	product side
H	4	4
O	2	2

Balanced and notice where the numbers for balancing have been placed



un balanced

Elements	Reactant side	Product side
Carbon	6	1
Hydrogen	12	2
Oxygen	8	3

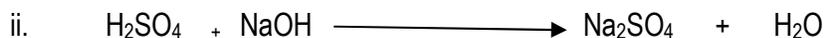


Element	Reactant side	Product side
Carbon	6	6
Hydrogen	12	12
Oxygen	18	18

Balanced and notice where the numbers for balancing have been placed

Exercise.

Balance the following equation;

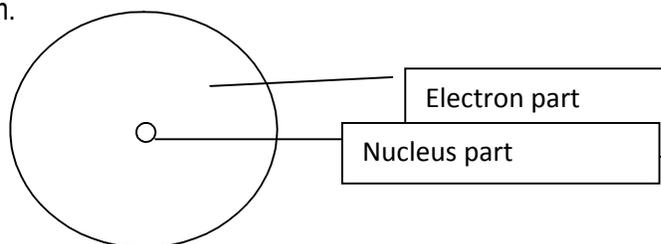


There is no specific formula for balancing chemical equations, instead all we do is to carefully look at the valencies and by inspection we can easily find the whole numbers that can be put in front of each chemical formula in the equation to balance its atoms

GRADE10 TERM 2

ATOMIC STRUCTURE

An atom is the smallest neutral particle of matter that can take part in a chemical reaction. An atom is made up of two main parts namely the nucleus and the electron part. The electrons move around the nucleus in a manner similar to the way planets move around the sun. The nucleus contains mainly neutrons and protons which are collectively known as nucleons meaning particles found in the nucleus. The simplified diagram below shows the structure of an atom.



Perceived structure of an atom

Volume of an atom;

The volume of an atom is taken to be the space occupied by the electrons around the nucleus. The volume of the nucleus is so small that it contributes almost nothing to the total volume of the electron.

Mass of an atom;

The mass of an atom is taken to be the mass of the nucleus. The mass contributed by the electrons is negligible. The mass of one (1) proton is taken to be 1 unit and the mass of one (1) neutron is taken to be 1 unit. And the mass of one (1) electron is taken to be 1/1836 the mass of one proton. This shows that the mass of one electron is far too small compared to the mass of either the proton or the neutron. Therefore for all practical purposes the mass of an atom is taken to be the mass of the nucleus.

When determining the mass of an atom, we simply add the neutrons and the protons. i.e. if the atom has 8 protons and 9 neutrons then its mass is 17. Therefore the mass number of an atom is the sum total of the neutrons and the protons. The protons are also known as the atomic number.

Thus; $8+9=17$

Atomic symbol, atomic mass and atomic number;

The atomic symbol, the atomic mass and the atomic number; all three are printed together in the periodic table. Thus if the symbol of an atom is A_ZX of an atom, A is the mass number and Z is the atomic number. The three are written together as shown below;

A_ZX where; X is the atomic symbol or the chemical formula of an element, A is the mass number or the sum total of the protons and neutrons and Z is the atomic number or the number of protons.

In a neutral atom the number of electrons is equal to the number of protons. If the number of protons (atomic number) is 8 the number of electrons is also 8. The number of neutrons is found by subtracting the atomic number from the mass number, thus:

$$A-Z= N$$

Where; A is the mass number, Z is atomic number and N is the neutron number.

Charge of an atom

An electron has an electric charge of negative one (1^-) and a proton has an electric charge of positive one (1^+). Therefore an atom has a net charge of 0 or an atom is electrically neutral.

Isotopes;

Some atoms have more than one atom. The atoms have the same atomic number but different mass numbers. The difference in mass numbers is caused by the number of neutrons that varies in their nucleuses. Such atoms with the same atomic number but different mass numbers are called isotopes. Isotopes therefore are atoms of the same element with the same atomic number but different mass numbers. Below are some examples of elements and their isotopes.

Carbon -12, Carbon-13 and carbon 14; written as

Chlorine-35 and Chlorine-37 written as

Oxygen-15, oxygen-16, oxygen-17 and oxygen -18 written as

Radio Active Isotopes:

Radioactive isotopes decay and when they decay their atomic mass may reduce. They decay because they release one or two particles. The particles that they release are the gamma particles or alpha particles. The rays that they release are called beta rays. Radioactive decay therefore is the spontaneous release of alpha particle or gamma particles or beta rays from the nucleus of the radioactive isotope. Because of the

properties of gamma, alpha and beta; radioactive isotopes have found many uses in medicine, agriculture, archeology and industry.

Uses of radioactive isotopes;

Radioactive isotopes decay and the mass they have reduces with time. They decay because they emit one or two even three of the radioactive particles rays. The particles that they release are called rays because of the properties of beta, gamma and alpha radioactive isotopes have found many uses in medicine agriculture, archeology and other Industries

Use of radio Active Isotopes

1. Medicinal use:

- Can be used to treat cancer.
- Can be used to check for blood levels in the body
- Can be used to locate damaged tissues and locate blood clots.

2. Agricultural use:

- Phosphorous is used to check the amount of fertilizer taken in by plants.

3. Archeological use:

- Carbon 14 is used for Carbon dating to determine the age of fossils.

4. Uranium:

- Uses in mineral exproleton
- Tracing leakages in oil pipes. Prospecting petroleum deposits other radioactive substances are used to check the thickness of materials such as papers.

RELATIVE ATOMIC MASS

Although isotopes of elements exists with different mass numbers, only one figure is quoted as the atomic mass of the element. This figure is very close to the atomic mass of one of the isotopes and in some cases it is the atomic mass of the isotope. The figure or number that appears as the atomic mass of an element is as a weighed number of baised number calculated from the relative abundances of the isotopes most abundant. For example oxygen exists is oxygen 15, Oxygen 17 and Oxygen 18. The atomic mass that appears on the atomic mass that of the oxygen element is 16. This is the most abundant in nature. The relative abundances of oxygen many appear as follows:

Oxygen 15 ₈	¹⁵ O	2.01%
Oxygen 16	¹⁶ O	96.60%

8

Oxygen 17 ^{17}O 1.20%

8

Oxygen 18 ^{18}O 0.19%

8

From the above relative abundances, the relative atomic mass can be calculated as follows:

$$\text{Relative Atomic Mass} = \frac{2.01}{100} \times \frac{15}{100} + \frac{96.60}{100} \times \frac{16}{100} + \frac{1.2}{100} \times \frac{17}{100} + \frac{0.19}{100} \times \frac{18}{100}$$

$$= 0.3015 + 15.4360 + 0.204 + 0.342$$

$$= 15.9957$$

$$= \underline{16.0}$$

Actually the atomic mass is called relative atomic mass meaning the atomic mass is calculated from a biased average meaning that the atomic mass of an element favours the isotope most abundance in nature and from the above example it favours Oxygen-16.

ELECTRONIC CONFIGURATION

This is the system of arranging electrons in the shells of an atom.

The principles behind this are:

- (1) The shell with the lowest energy value is filled first followed by the shell with second lowest energy value and so on and so forth i.e.

- K 1st shell
- L 2nd shell
- M 3rd shell
- N 4th shell

- (2) The maximum number of electrons in a shell of position (n) from the nucleus is determined by the formula $2n^2$

Shell	N	Number of electrons	
K	1	$2 \times 1^2 = 2$	
L	2	$2 \times 2^2 = 8$	
M	3	$2 \times 3^2 = 18$	
N	4	$2 \times 4^2 = 32$	

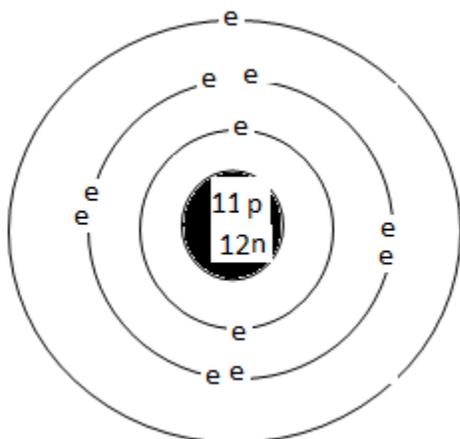
1. Duplet rule: This is the rule which states that maximum number of electrons in the first and the only shell is 2.
2. Octet rule: The maximum number of electrons in the last shell is 8.

Electrons in an atom are arranged in order of increasing energy levels. These energy levels are called shells. Each shell has orbitals. Electrons move around the nucleus in the orbitals just like planets move around the sun. The shells in order of their increasing distance from the nucleus are assigned the letters K, L, M and so on. The Shell that has the least amount of energy is K because it is closest to the nucleus and its electrons are strongly attracted by the positively charged protons in the nucleus. The shell further from the nucleus has the greatest amount of energy

because its electrons are weakly attracted by the protons in the nucleus; the electrons are weakly attracted by the protons because of shielding by the inner electrons.

The first shell, which is the closest shell to the nucleus, can only accommodate a maximum of two electrons. The second shell can accommodate a maximum of eight electrons and the third shell can accommodate a maximum of 18 electrons according to the formula

For example sodium with the atomic number 11 can have its electron configuration as shown below;



Electron configuration of sodium: 2, 8,1

EXERCISE;

Draw the electronic structure of calcium, (Ca-40 ; 20p), Oxygen (O-16; 8p) and Chlorine (Cl-37; 17p)

CHEMICAL BONDING

A bond is a link or a force that keeps two particles of matter together. Elements react because they are initially unstable so they react in order to gain stability. The stability is gained by attaining the noble gas electronic configuration of two (duplet) or eight (Octate) electronic configuration in the outer most shell. Duplet applies only to the first shell for example lithium and Beryllium attain the electronic configuration of Helium by losing the two electrons in its second shell, to remain with two electrons in the first shell which becomes the only shell. Sodium, Magnesium and Aluminium and together with the rest of the elements in the period obtain stability by attaining the electronic configuration of Argon

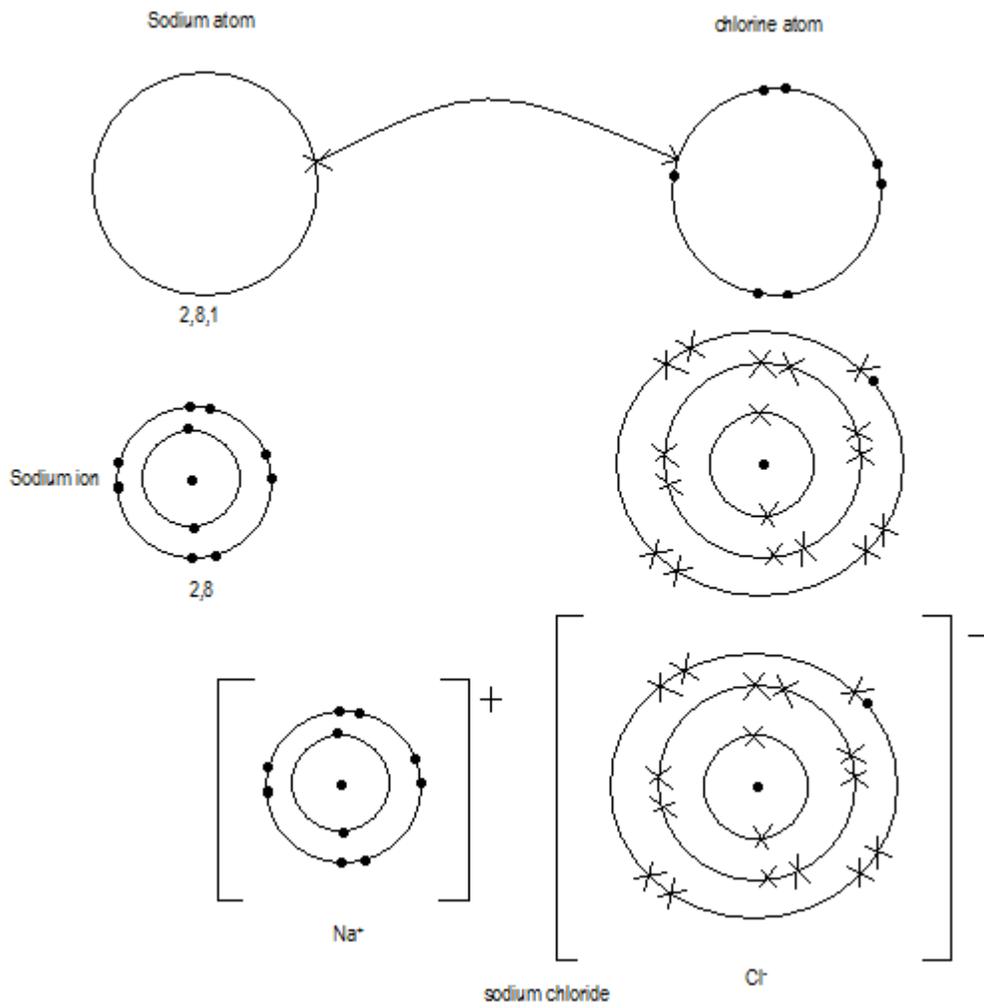
IONIC BONDING

Under Ionic Bonding elements attain stability by completely transferring valence electrons to another element or by gaining the electrons that have been transferred to it. The losing and gaining of electrons creates charged particles of matter. The charged particles of matter are called **ions**. The atom that loses the electrons becomes positively charged because the number of protons become more than the number of electrons. A positively charged particle (ION) is called a **Cation**. The gaining of electrons creates negatively charged ions because the number of electrons becomes more than the number of protons. A negatively charged ion is called an Anion.

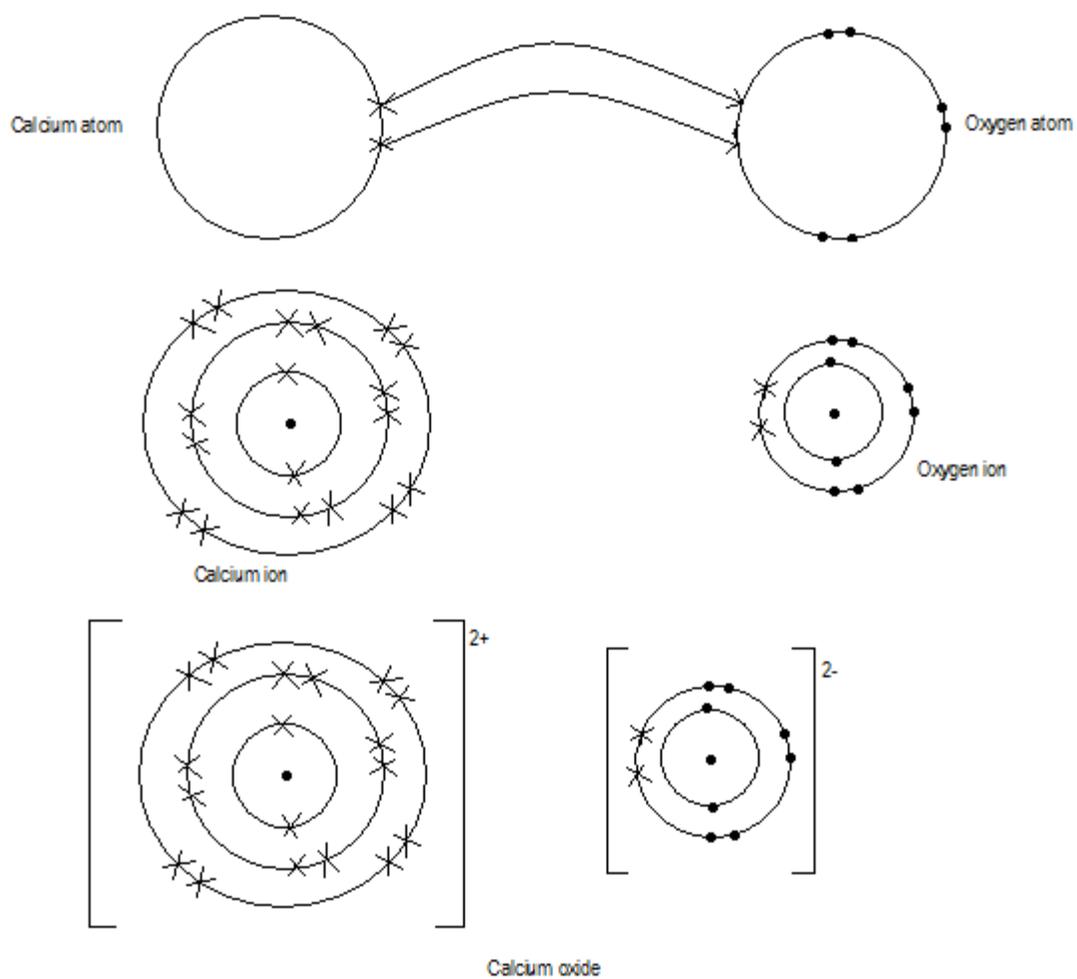
There is an attraction between the positively charged ion and a negatively charged ion. This attraction forms a bond between the two ions and since the bond is formed between ions is called an **Ionic Bond** or **Electrostatic Bond** because it is formed from ions that are electrically charged.

The diagram below shows Ionic Bonding on the valence shell

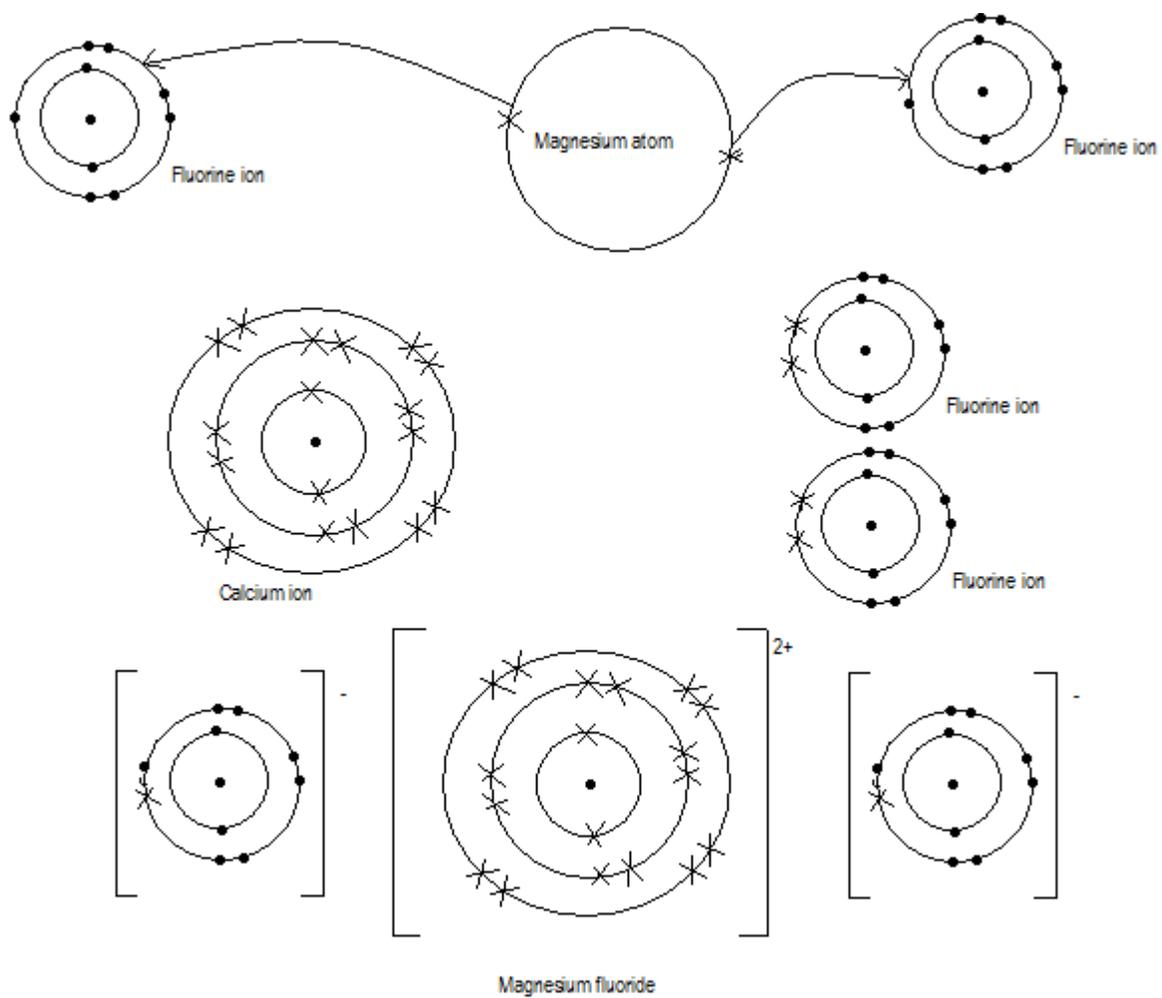
(a) Sodium Chloride, the ionic bond is between Sodium and Chloride.



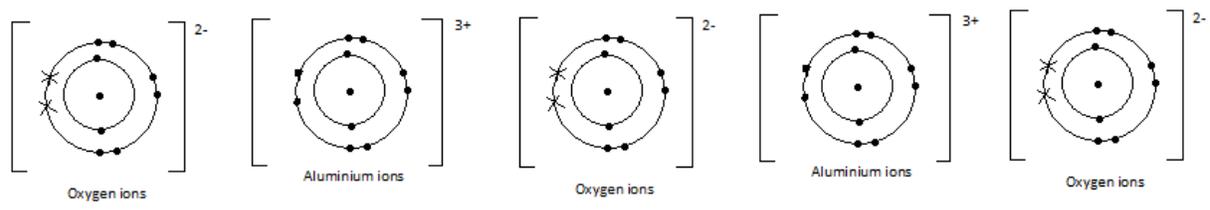
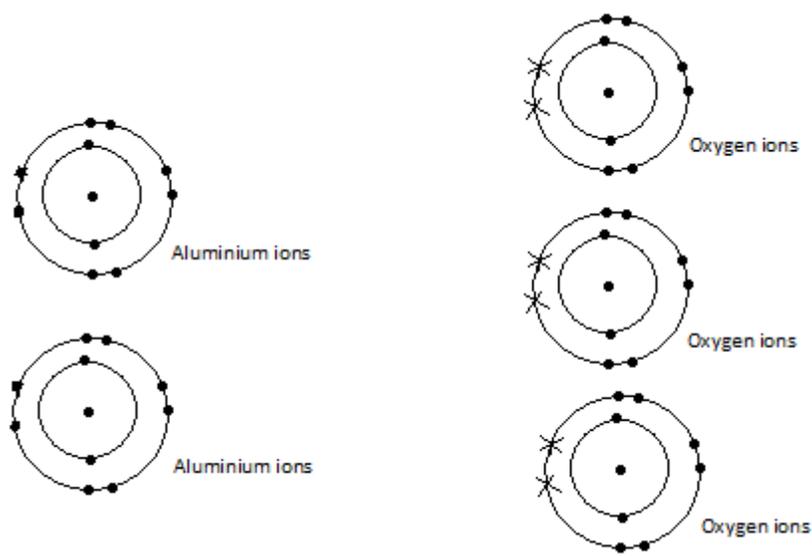
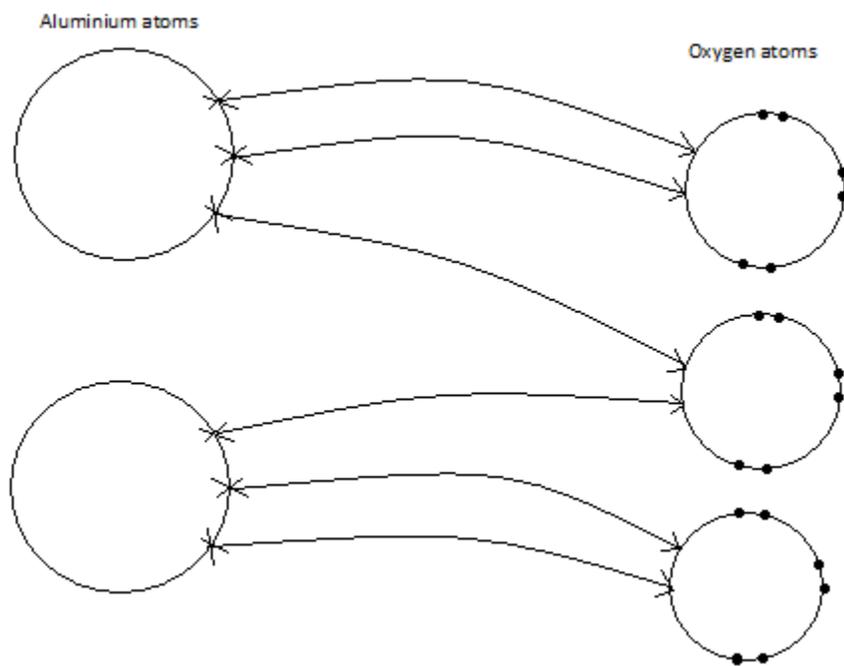
(b) Calcium oxide: Ionic bond between calcium and oxygen.



(c) Magnesium fluoride, Bond between Magnesium and fluorine.



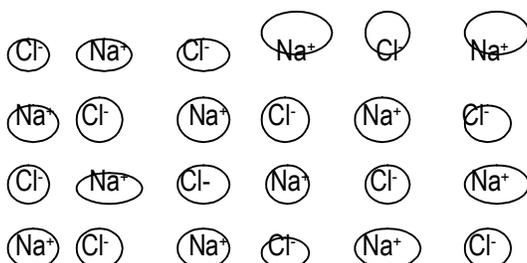
(c) Aluminium Oxide, the bond is between aluminium and Oxygen.



IONIC LATTICE

Sodium chloride, Magnesium oxide, and Aluminium Oxide are examples of electrovalent or Ionic compounds. In these compounds, Ionic form very big structure or Lattice rather than molecules. For examples in the Sodium

Chloride lattices, each Sodium ion is surrounded by six Chloride ions and each chloride ion is surrounded by six Sodium ion as shown below:

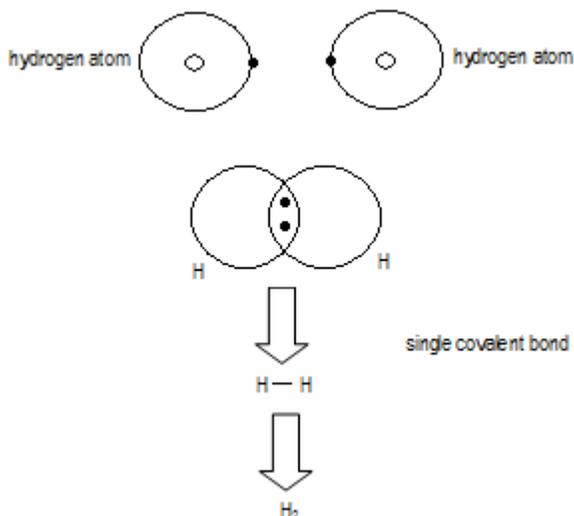


Characteristics of Ionic compounds.

- Have very high melting and boiling points e.g. Sodium Chloride melts at 647°C and boils at 1227°C. This is due to strong electrostatic forces of attraction between the negatively and positively charged ion. To break such bonds requires a great deal of energy.
- Electrovalent compounds are non-volatile. They are generally solids at room temperature.
- Electro-valent compounds are soluble in water. This is because water breaks the electrostatic forces of attraction.
- They are insoluble in organic liquid such as cooking oil, paraffin and diesel.
- Ionic compounds conduct electricity when they are in a solution state or molten state. This is because the negatively and positively charged ions are separated.
- Ionic compounds are formed mostly between metals and non- metals.

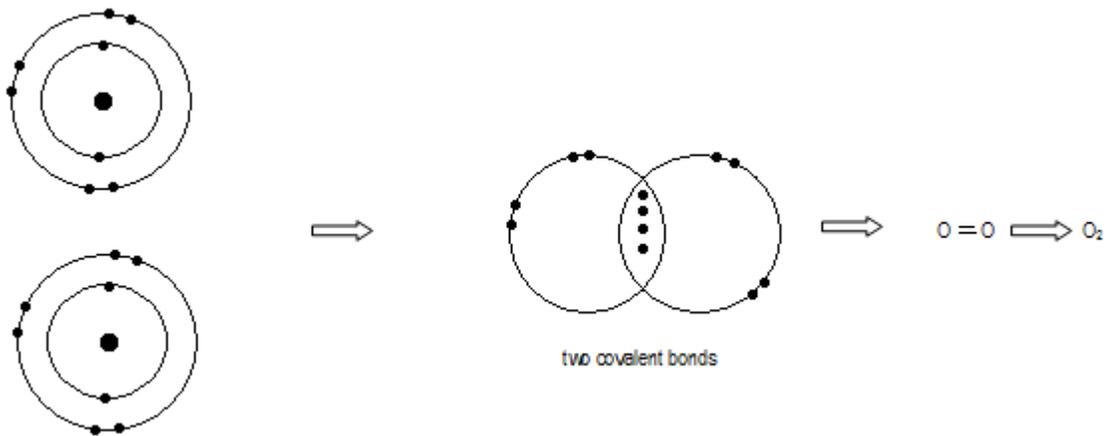
CONVALENT BONDING

Under covalent bonding elements or atoms obtain stability by sharing electron also known as molecular bond. This is made up of a pair of electrons. It is formed between non-metallic of elements which share one or more of electrons. Compounds that are formed through sharing of electrons are known as molecules. Below are some covalent

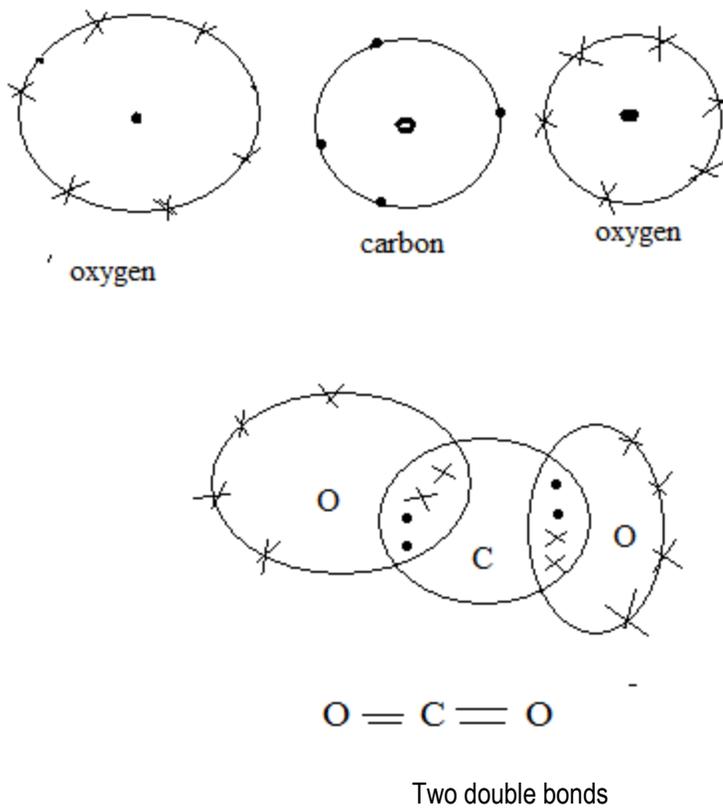


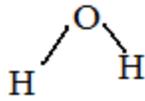
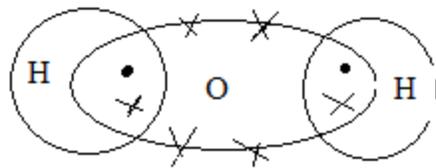
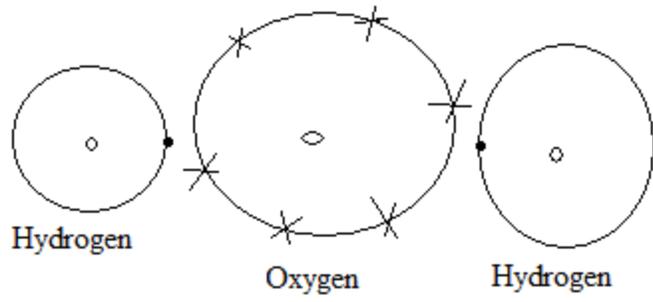
molecules and how they are formed. Hydrogen molecule

Oxygen molecule



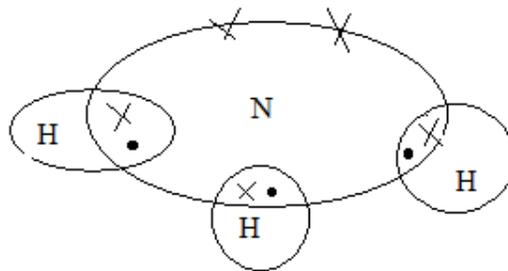
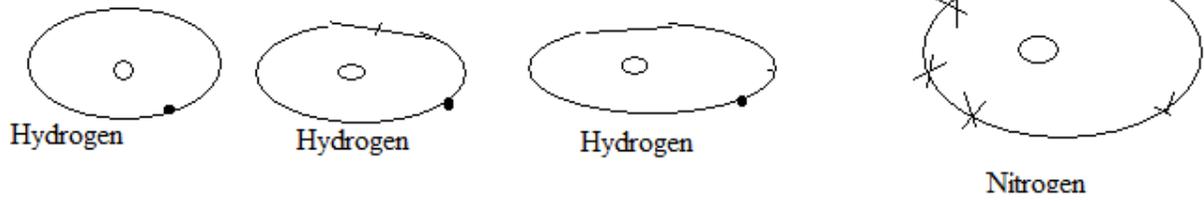
Carbon dioxide molecule



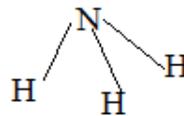


Two single covalent bonds

Ammonia molecule



Three covalent bonds

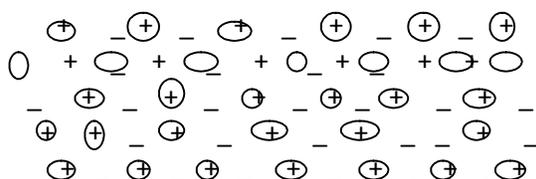


CHARACTERISTICS OF CONVALENT COMPOUNDS

- Formed between non -metals.
- They have low melting and boiling due to weak intermolecular forces of attraction.
- They are discrete molecule.
- They are insoluble in polar solvent like water but soluble in organic solvents like petrol.
- They are volatile (they can easily evaporate) because of the volatility some covalent compounds like chlorine are poisonous
- Do not conduct electricity because they have neither free electrons nor ions
- Since they are volatile; some compounds such as methane, petrol and some fine powders are very inflammable and this causes rapid explosion.

METALIC BONDING

Atoms of metals obtain a full outermost shell by losing their valency electrons, e.g. Sodium atoms lose a single electron from outermost shell. When a large number of sodium atoms lose these electrons, the results are many free electrons moving among sodium ions. The Sodium ion which is positively charged repels each other but they are held together by the negatively charged free electrons between them. The attractive forces between the electrons and the Sodium ions are greater than the repulsive forces between the Sodium ions. The electrons are free to move anywhere in the metallic lattice. In this case the electrons are said to be Delocalised. The attraction between the electrons and the positively charged ions of a metal is called Metallic Bonding.



Sodium Metal Lattice

Key

○ Electron

⊕ Sodium ions

Properties of Metals

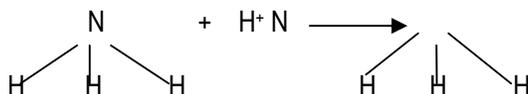
- Shiny when polished.
- They are malleable i.e. can be hammered into a large sheet of metal.
- They are ductile i.e. they can be drawn into a fine wire.
- Good conductors of electricity because of the delocalized electrons.
- High melting and boiling point therefore at room temperature pressure they are solids.

Co-ordinate or Dative bonding

Another variety of covalent bond that is electron sharing has been given a name co-ordinate or dative bond. This bond is characterized by the fact that the two shared electrons are both supplied by one of the participating atoms.

A co-ordinate bond is formed when one of the atoms possesses a lone pair of electrons which is donated to an atom which needs them to complete an electronic configuration of duplet or octet. The Ammonium molecule possesses such a lone pair as shown below

H⁺



Ionic Equations

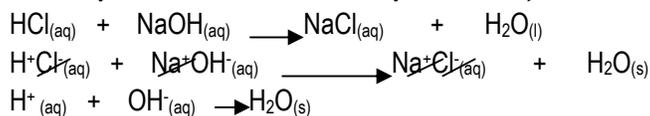
Ionic equations are usually derived from chemical equations. Ionic equations represent the actual ions (chemical specie) taking part in a chemical change to produce respective products. Ionic equations can tell us exactly where a product is coming from (especially solid products formed from aqueous solution). However, it is not every chemical equation which can be converted to a ionic equation. The following rules must be observed when converting equations to an ionic equation.

- Compounds in aqueous state must be ionized.
- Spectator ions and the ionized compounds must cancel each other in equality or the reactant and product.
- Express the chemical equation including state of the compound.
- Include the state symbol on the ionic equation formed.

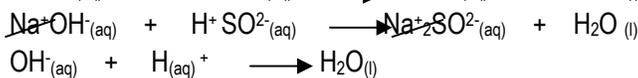
Examples:

Convert the following chemical equations to their ionic equations.

- (a) Dilute hydrochloric acid + Sodium hydroxide \longrightarrow water + aqueous Sodium Chloride



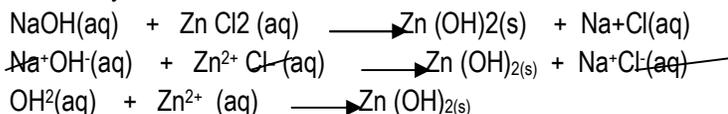
- (b) $\text{NaOH}_{(\text{aq})} + \text{H}_2\text{SO}_{4(\text{aq})} \longrightarrow \text{Na}_2\text{SO}_{4(\text{aq})} + \text{H}_2\text{O}_{(\text{l})}$



- (c) $\text{K}_2\text{CO}_{3(\text{aq})} + \text{H}_2\text{SO}_{4(\text{aq})} \longrightarrow \text{K}_2\text{SO}_{(\text{aq})} + \text{CO}_{2(\text{g})} + \text{H}_2\text{O}_{(\text{l})}$



- (d) Sodium hydroxide and Zinc Chloride solution?



THE PERIODIC TABLE

If one were to study each of the 107 elements individually, the task could be very tedious. To reduce this big task elements are put into groups. By studying the characteristics of one member of the group; the characteristics of other members could be predicted. In the study of the halogens (Group 7) elements for example it is better to study chlorine in detail and then study the other halogens i.e. fluorine, Bromine and Astatine by comparison. The halogens are found to be very similar in their chemical properties to chlorine. Knowledge of one of the halogens enable us to predict with considerable accuracy the chemical behaviour of others. The same is true of the members of the nitrogen family and the alkaline metals.

Early attempt to classify element into groups for easy study were made. John Newlands for example noted that there were many pairs of similar elements. In particular he noted that in each pair the atomic weights differed by a multiple of eight and he put up the law of octaves; meaning an element behaves like the eighth one following it in the table. Many scientists attempted to classify the elements. Some of the classifications were;

1. Dividing elements into metals and non metals.
2. Dividing elements into base formers and acid formers.
3. Arranging elements according to atomic weights.

But none of these attempts made things easier. It was the Russian Dmitri Mendeleev who is credited with the arrangement of the elements in the periodic table. He arranged elements according to atomic weights and similar chemical behavior. Mendeleev's arrangement left gaps for the undiscovered element and on this basis the Periodic Table was constructed. The discovery of atomic numbers in 1911 by Ernest Rutherford brought changes to the Mendeleev's arrangement. Today the periodic table is described as the arrangement of elements according to their increasing atomic numbers. The periodic table is made up of two main parts namely: **Groups** and **Periods**. Groups or families increase from left to right and lie vertically while periods increase from top to bottom and lie horizontal.

PERIODS;

Period one has only two elements namely hydrogen and helium. Periods two and three have eight elements each and are known as short periods. Periods four, five and six have eighteen elements each and are known as long periods. Along the period from left to right;

1. There is a gradual change in characteristics from the metallic character to metalloid characteristics and finally to a nonmetal characteristics.
2. The number of shells around the nucleus remains constant while the numbers of electrons increase. Thus if we know the period number then we know the number of shells.
3. The atomic radii of atoms decrease this is due to an increase in the number of protons in the nucleus.

Groups or Families;

Groups are made up of elements that have similar chemical properties. A group or family is a vertical column

Group 1

This group is called the alkali metals and is made up of lithium, sodium, potassium, rubidium, cesium and francium. Their electronic configurations are:

${}_{3}\text{L}$	2, 1
${}_{11}\text{Na}$	2, 8, 1
${}_{19}\text{K}$	2, 8, 8, 1
${}_{37}\text{Rb}$	2, 8, 18, 1

Physical properties of group one elements;

1. They are soft and can easily be cut with a knife
2. They are solids at room temperature and pressure

3. They have low melting points decreasing down the group from lithium at 180⁰C to cesium at 28⁰C
4. Their densities increase down the group starting with lithium at 0.53g/cm³ to cesium at 1.87g/cm³
5. All the alkali metals are good conductors of heat both in a solid state or fused state
6. Their hydroxides are solids at room temperature and pressure
7. Their hydroxides are very stable. They do not decompose when heated
8. Their chlorides have very high melting points

Chemical properties of group one elements;

1. They easily react with oxygen to form an oxide of the metal hence they are stored under oil which is inert

$$2\text{Na} + \text{O}_2 \longrightarrow \text{Na}_2\text{O}$$
2. They have one electron in the valence shell hence very reactive
3. They are very reactive and their reactivity increase down the group; i.e. rubidium is more reactive than potassium and potassium is more reactive than sodium. Sodium reacts vigorously in water while potassium burns when dropped in water.
4. They are found in nature in ionic compounds
5. They dissolve in water to form an alkali solution called the hydroxide of the metal e.g. sodium hydroxide.

Group 2

This group is called the earth alkali metals and is made up of the following elements; beryllium, magnesium, calcium, strontium, barium and radium. Their electronic configurations are:

₄Be 2, 2

₁₂Mg 2, 8, 2

₂₀Ca 2, 8, 8, 2

₃₈Sr 2, 8, 18, 8, 2

Chemical properties;

1. They form positive 2+ ions
2. The ease with which they lose electrons is only second with group one
3. Like group one; they react with cold water except magnesium which requires warm water
4. Their oxides and hydroxides are more basic
5. They are less reactive than group one elements hence they are not stored in oil
6. They burn in air to form an oxide of the metal

Physical properties;

1. They are soft and light

2. They have high melting and boiling points
3. They are good conductors of heat and electricity
4. The hydroxide of calcium is sparingly soluble while the hydroxide of magnesium is far less soluble.
5. Their compounds are more hydrated but the hydration decrease down the group. i.e. they have a high amount of water of crystallization

GROUP 7 (SEVEN)

This group is called the halogen group. It is made up of fluorine, chlorine, bromine, iodine, and astatine. The group is called the halogen which means salt formers because members react with metals to form salts e.g. sodium chloride. Their electronic configurations are;

${}_{9}\text{F}$ 2, 7

${}_{17}\text{Cl}$ 2, 8, 7

${}_{35}\text{Br}$ 2, 8, 18, 7

${}_{53}\text{I}$ 2, 8, 18, 18, 7

Physical properties;

1. Fluorine and Chlorine are gases at room temperature and pressure, Bromine is a liquid at room temperature and pressure, Iodine and Astatine are solids at room temperature and pressure
2. Chlorine is greenish-yellow, bromine is brown, iodine is dark-grey, and astatine is black
3. They have an evil smell and are very poisonous.
4. Bromine is very volatile and diffuses very fast
5. Iodine sublimates when heated.
6. Their boiling points increase down the group;

Fluorine -188°C ,

Chlorine -33.5°C ,

Bromine 59°C ,

Iodine 184°C ,

The table below summarizes the physical properties of the halogens

Halogen	State at r.t.p.	Colour	Melting point	Comment
Fluorine	Gas	Greenish.	-188°C	Poisonous
Chlorine	Gas	Greenish-yellow	-33.5°C	Poisonous
Bromine	Liquid	Brown	59°C	Poisonous
Iodine	Solid	Dark-grey	184°C	Sublimes and poisonous
Astatine	Solid	Black		-

Chemical properties;

1. They are the most electro negative elements.
2. They have 7 electrons in the valence shell.
3. They are all nonmetals

4. They are very reactive and their reactivity decrease down the group
5. They easily form covalent compounds with other nonmetal
6. The element above displaces any element below it in the group.
7. At room temperature and pressure they exist as diatomic molecules

Importance of the Halogens

1. They are used to manufacture insecticides
2. Used to manufacture drugs in pharmaceutile industry
3. When subjected to ultra-violet rays they can produce CFC_s which depletes the ozone layer.
4. Iodine is added to sodium chloride to prevent goiter development

Group 8 or group 0

This group has six gaseous elements which are not reactive and therefore they are called noble gases or inert gases. They were discovered after Mendeleev had constructed his periodic table. Their electronic configurations are;

${}^2\text{He}$	2
${}^{10}\text{Ne}$	2, 8
${}^{18}\text{Ar}$	2, 8, 8
${}^{36}\text{Kr}$	2, 8, 18, 8
${}^{54}\text{X}$	2, 8, 18, 18, 8

Chemical properties

1. They have 8 electrons in the valence shell except helium which has 2 electrons in the first and only shell.
2. They are very inert and hence they do not form any compounds

Physical properties

1. They are all gases at room temperature and pressure
2. Their boiling points increase down the group

Uses of group 0 elements

They are mostly used in hot environments to provide an inert atmosphere. For example;

1. Helium is used in hot air balloons because it is light and un reactive
2. Neon is used in advertising signs because of its colour
3. Argon is put in florescent tubes to prove un reactive atmosphere

Transition elements;

These elements are made up of 30 elements; all of them are metals and are inserted as a rectangular block between group two and group three in the periodic table. They are sometimes known as group D elements.

Physical properties;

1. They have very high melting and boiling points
2. They have very densities and hence they are called heavy metals
3. They are good conductors of heat and electricity
4. They have high tensile strength.

Chemical properties;

1. They have more than one valence e.g. iron (II) and Iron (III).
2. They form coloured salts e.g. copper (II) sulphate is blue, iron (II) is green and cobalt chloride is pink.
3. They are electro-positive
4. They form complex compounds through coordinate bonding.

Uses of transition elements

Many transition elements are used as catalysts e.g iron is used in the Haber process to produce ammonia

GRADE 10 CHEMISTRY 5070/123

TERM 3

STOICHIOMETRY AND MOLE CONCEPT

Stoichiometry is a branch of chemistry which deals with the measurements of quantities of elements and compounds in Chemical reactions.

The mass of an atom is very small and it is referred to as the atomic mass unit (a.m.u). It is determined by using an instrument called mass spectrometer. Since the instrument is very rare; for all practical purposes in schools and colleges the atomic mass unit is expressed in grams and kilograms. Therefore the scientific electronic balance or the beam balance can be used to measure the mass of a substance for example:

- a. The mass of sodium metal is 23g
- b. the mass of aluminium is 27g
- c. The mass of oxygen atom is 16g

To determine the relative mass of a compound we simply add up the relative atomic masses of the atoms (elements) in the chemical formula and after considering the valences. For example the relative molecular mass of sodium chloride is calculated as follows:

NaCl;

Na = 23.0g

Cl = 35.5g

58.5g

Relative molecular mass of water:

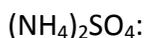


$$\text{H} = 1 \times 2 = 2\text{g}$$

$$\text{O} = 16 \times 1 = \underline{16\text{g}}$$

$$18\text{g}$$

Relative molecular mass of ammonium sulphate:



$$\text{N} = 14 \times 2 = 28\text{g}$$

$$\text{H} = 1 \times 4 \times 2 = 8\text{g}$$

$$\text{S} = 32 \times 1 = 32\text{g}$$

$$\text{O} = 16 \times 4 = \underline{64\text{g}}$$

$$132\text{g}$$

One mole is an amount of a substance corresponding to the formula mass or formula weight or molecular mass or molar mass or atomic mass of a pure substance. or more technical the mole is defined as the amount of the substance which contains the same number of particles as there are in 12g of Carbon – 12 isotope. The carbon-12 isotope is chosen because it measures exactly 12 a.m.u. on the mass spectrometer. Or the mole is defined as any amount of substance that contains the Avogadro's number of particles. The terms formula weight, formula mass and molar mass are used interchangeably and they mean the same thing. The actual number of particles in one mole of any substance is 6.02×10^{23} particles. The particles could be atoms or molecules or electrons or protons or neutrons or ions. The table below shows one mole of each substance

Name of the substance	Chemical formula	Formula mass/ weight	Number of moles
Hydrogen atom	H	1g	1 mole of hydrogen atoms
Hydrogen molecule	H ₂	2g	1 moles of hydrogen molecules
Carbon dioxide	CO ₂	44g	1 mole of carbon dioxide
Ammonium sulphate	(NH ₄) ₂ SO ₄	132G	1mole of ammonium sulphate

We can calculate any number of moles in any given quantity of a named substance because we can always calculate the relative molecular mass of the substance from the chemical formula provided the chemical formula is correct. Examples below show how this could be done;

1. The number of moles in 10g of sodium hydroxide;

Solution: formula mass of NaOH = 23+16+1 =40g

$$1 \text{ mole of NaOH} \rightarrow 40\text{g}$$

$$X \text{ moles of NaOH} \rightarrow 10\text{g}$$

$$40x = 10$$

$$X = \frac{10}{40}$$

$$X = 0.25 \text{ moles}$$

2. The number of moles in 80g Of NaOH can be calculated as follows:

Solution: formula weight for NaOH = 23 + 16 + 1 = 40g



$$40X = 80$$

$$X = \frac{80}{40}$$

$$X = 2 \text{ moles of NaOH}$$

If we have the number of moles and the formula mass of the substance, we can calculate the mass that gives us that number of moles. The example below illustrates this;

1. What mass of sulphuric acid gives 0.5moles of sulphuric acid?

Solution: the formula weight of H₂SO₄ = (1X2) + (32) + (16X4) = 98g

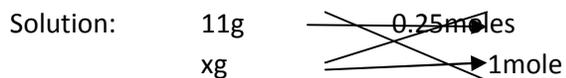


$$x = 0.5 \times 98$$

$$x = 49 \text{ g}$$

If we have the number of moles and the mass, we can also calculate the formula mass or formula weight of the compound. The example below illustrates this:

1. 11g of a pure substance gives 0.25 moles of the substance. Determine its formula weight.



$$0.25x = 11 \text{ g}$$

$$x = \frac{11}{0.25}$$

$$x = 44 \text{ g of the substance is 1mole}$$

The Avogadro's law

This law states that one mole of any substance contains 6.02×10^{23} particles. Hence if we know the number of moles we can calculate the actual number of particles contained in that number of moles. The example below shows how it is done;

How many particles are in 0.25 moles?

Solution:

1mole	→	6.02X10²³
0.25mles	→	x

x	=	6.02X10 ²³ X 0.25
x	=	1.5X10 ²³ particles.

Exercise

- The formula mass for sulphuric acid (H_2SO_4) is 98g
 - how many moles are in 60g of sulphuric acid
 - how many molecules of sulphuric acid are in 49g of sulphuric acid.

MOLAR GAS VOLUME

The molar gas volume of gases depends on pressure and temperature i.e. the volume occupied by any gas depends on pressure and temperature. Under stated conditions gases occupy fixed volumes. The conditions mostly used in science are; Standard Temperature and Pressure (STP) and the Room Temperature and Pressure (RTP). At Standard Temperature and Pressure; one mole of any gas occupies 22.4dm^3 . The 22.4dm^3 is called the molar gas volume at STP. At Room Temperature and Pressure; one mole of any gas occupies 24dm^3 or 24000cm^3 . The 24dm^3 is the molar gas volume at Room Temperature and pressure.

CALCULATION OF VOLUME OCCUPIED BY GASES:

Examples:

- What volume is occupied by 4 moles of Nitrogen gas (N_2)?
 - At r.t.p
 - at s.t.p

SOLUTION: at r.t.p

(a) 1mol	→	= 24 dm³
4 moles	→	= x
x		= 96 dm ³ of N_2

(b) at s.t.p;

1mole	→	22.4dm³
4moles	→	x

X	=	22.4 x 4
X	=	89.6dm ³

COMPOSITION BY MASS AND BY PERCENTAGE

A compound is a pure substance made up of two or more elements chemically combined in fixed ratios. Since the elements are in fixed ratios their composition by mass and percentage are also in fixed ratios. The composition by mass and by percentage of any given quantity can be determined.

The examples below show how this can be done

(iii) Moles of water, H₂O produced.



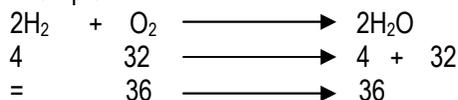
0.5 moles of Zinc reacted. Calculate:

- Mass of ZnCl₂ produced.
- Moles of HCl used.
- Volume of H₂ produced measured at r.t.p.

REACTING MASSES

The Law of Conservation of mass is always observed in a balanced chemical reaction. Law conservation of mass states that the total mass of reactants is equal to the total mass of products in a balanced equation.

Example:



CALCULATION OF REACTING MASSES

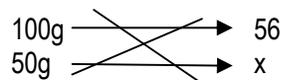
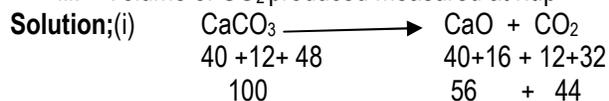
Example:

If 50g of calcium Carbonate, CaCO₃ was used in the reaction as follows:-

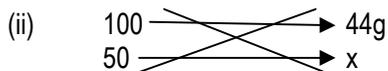


Find Mass of

- CaO produced
- CO₂ produced
- Volume of CO₂ produced measured at r.t.p



$$\begin{array}{rcl} 100x & = & 50 \times 56 \\ X & = & \underline{28\text{g}} \end{array}$$



$$\begin{array}{rcl} 100x & = & 44 \times 50 \\ X & = & \underline{22\text{g}} \end{array}$$

(iii) Moles of CO₂ produced = $\frac{\text{mass}}{\text{Formula mass}}$
= $\frac{22\text{g}}{44\text{g}}$
= $\underline{0.5\text{moles}}$

1 mole of any gas at rtp occupies \longrightarrow 24dm³

0.5 moles of any gas at rtp occupies \rightarrow x

$$X = 24 \times 0.5$$

$$X = \underline{12\text{dm}^3} \text{ volume occupied by CO}_2 \text{ at rtp.}$$

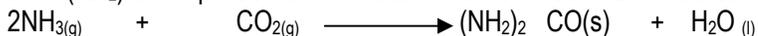
Exercise

1. Magnesium metal reacts as follows:-



- Mass of Magnesium used.
- Mass of H_2SO_4 used.
- Volume of H_2 produced at s.t.p.

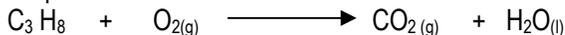
2. Urea (NH_2)₂ CO is produced when ammonia reacts with carbon dioxide as follows:



If 68 tonnes of ammonia was used, calculate:

- Mass of Urea (NH_2)₂ CO produced.
- Volume of CO_2 used at r.t.p.

3. Propane $\text{C}_3 \text{H}_8$ burns as follows:



- Balance the equation
- If 88g of Propane was used, find:
 - Mass of water produced
 - Volume of CO_2 produced

EMPIRICAL AND MOLECULAR FORMULA

Empirical Formula:

It is a formula which shows the simplest ratios of atoms of elements present in a compound e.g. the empirical formula of $\text{Al}_4 \text{O}_6$ is $\text{Al}_2 \text{O}_3$

Molecular formula:

It is a formula which shows the actual number of atoms of each element present in a compound e.g. Al_4O_6

CALCULATIONS OF EMPIRICAL FORMULA

STEPS:

- Divide each percentage/mass by the relative atomic mass of the element.
- Divide each of these answers by the smallest value of them.
- The numbers you get are the ratios which represent numbers of atoms of elements present in a compound.
- Make sure you round off to the whole number
e.g. In a black Copper Oxide, there is 79% Copper and 12% of Oxygen. Find the empirical of Copper oxide.

Cu

O

3. A compound of Carbon hydrogen and Oxygen contains **40.0%, 6.6%** and **53.4%** by mass respectively. It's molar mass is **90g/mol**.
 - (a) Calculate empirical formula
 - (b) Molecular formula.
4. An Oxide of hydrogen has a percentage composition by **Mass H = 5.9%** and **O = 94.1%**. Its molecular mass is **34**. Find the molecular mass of the Oxide.
5. A compound with relative mass of **88** consists of **54.55% Carbon, 9.09% Hydrogen** and **36.365 Oxygen** by mass. Calculate its molecular formula.
6. **2g** of organic compound gave upon complete combustion **4.86g CO₂** and **2.03g H₂O**. The compounds contains **C, H, O** only. What is the empirical formula of the compound

SOLUTIONS AND CONCENTRATIONS

A solution is a uniform mixture of a solute and a solvent. A solvent is a substance that dissolves a solute and a solute is a substance that dissolves in a solvent. The concentration of a solution depends on the amount of solute that has dissolved in that solution or concentration is the amount of solute in a given solution. A dilute solution has less solute and more solvent. A concentrated solution has much more solute than a dilute solution. If more and more solute is being added, the solution becomes saturated. A saturated solution is one that has dissolved a maximum amount of solute at that temperature and any extra solute will not dissolve and instead it will collect at the bottom of the container. Solute is usually expressed in number moles.

COMMON UNITS OF CONCENTRATION

- I. g/dm³(grams per cubic decimeter)
- II. mol/ dm³ (mol per cubic decimeter)
- III. Grams per cubic decimeter (Solid solutes)
- IV. Parts per million
- V. Volume/volume
- VI. Mass/mass
- VII. Percent/percent

Molarity and grams per cubic decimeter are mostly used in chemistry to express concentration. To find concentration in grams, we use the relationship.

$$\text{Concentration} = \frac{\text{Mass in grams}}{\text{Volume in dm}^3}$$

Note: 1dm³ = 1 000cm³

Example; What is the concentration of dissolving 8g of Sodium hydroxide, Na OH in 200cm³ of water?

$$\text{Con} = \frac{\text{Mass in grams}}{\text{Volume in dm}^3}$$

$$\frac{8}{200} \div 1000 = \frac{8}{0.2} = \underline{\underline{40\text{g/dm}^3}}$$

Molarity.

This is the concentration expressed in mol/dm³. It is the number of moles of a solute in 1 dm³ of solution. The unit of molarity is **mol/dm³** or **M**.

A solution which contains 1 mole of a solute in 1 dm³ of solution is a molar solution.

$$\text{i.e. } \frac{1 \text{ mol}}{1 \text{ dm}^3} = 1 \text{ mol/dm}^3 \text{ or } 1\text{M}$$

A solution which contains 2 moles of a solution in 1 dm³ is 2 molar solution i.e.

$$\frac{2 \text{ mol}}{1 \text{ dm}^3} = 2\text{M}$$

Molarity is the same as Molar Solution.

$$\text{Molarity} = \frac{\text{moles}}{\text{dm}^3}$$

e.g. Calculate molar solution of 0.1 mole Na₂CO₃ in 250cm³ of water.

Exercise

- What is the concentration of a solution in g/dm³ made by dissolving.
 - 10.6g of sodium carbonate (Na₂CO₃) in 250cm³ of water?
 - 80.0g of NaOH in 200cm³ of water?
 - 10g of Sodium thiosulphate in 800cm³
- Find the Molarity thiosulphate in 800cm³?
 - 0.05 mol of NaOH in 200cm³ solution.
 - 2g of NaOH in 200cm³ solution
 - 80g of NaOH in 2dm³

Conversion from: g/dm³ to Mol/dm³

$$\text{Mol/dm}^3 = \frac{\text{grams}}{\text{dm}^3} \div \text{molar mass} = \frac{\text{mol}}{\text{dm}^3}$$

Example:

Convert 10g/cm³ of NaOH to Mol/dm³

- Calculate molarity of 2g of NaOH dissolved in 250cm³ of water.

DILUTION:

This is a process of making the solution weaker by adding more solvent to the solution. The weaker the solution has low concentration but large volume, while the stronger solution has high concentration but small volume when both contain same amount of a solute (Moles).

LAW OF DILUTION:

It states that the number of moles in the dilute solution is equal to the number of moles in the concentration solution if the same amount of solute is used.

Number of moles in conc = no. of moles in dilute

$$M_c V_c = M_d V_d$$

Note that mole = V x M

Find the concentration of a dilute solution of volume 2dm³ prepared from a high concentration solution of volume 0.5 dm³ 10mol/dm³.

exercise

1. Find the volume of 12M Sulphuric acid concentrated used to prepare dilute 0.2M Sulphuric acid of volume 5000cm³.
2. Find the moles contained in 25cm³ of 0.1000 mol/dm³ of NaOH. Leave the answer to 4 significant figures.

DETERMINATIONS OF MOLARITY AND VOLUME IN TITRATION PRACTICALS:

Titration is an analytical method used to determine the unknown concentration of a titrant when one of the two solutions has a known concentration. Titration deals with exact volumes and concentration.

1. Find the molarity of 40cm³ of HCl needed to neutralize exactly 20.0cm³ of 1m NaOH.

The equation for neutralization is



- (2) 25cm³ of Sulphuric acid, H₂SO₄ is neutralized by 27cm³ of 0.1m Sodium hydroxide, NaOH. What is the molarity of the acid?



- (a) Determine the molarity of 12.5cm³ of Sulphuric acid. H₂SO₄ used to titrate 25.0cm³ of 0.1000M NaOH.



- (b) Find the volume of 0.5M hydrochloric acid, HCl used to titrate 25.0cm³ of 0.1000m NaOH.



LIMITING AND EXCESS REACTANT:

In a chemical reaction involving two or more reactants, the amount of the product formed depends on the reactant that is in smaller amounts. The reactants that are in smaller amounts determine the amounts of the product to be formed and such reactants are called limiting reactants in other words a limiting reactant is a reactant upon which the chemical reaction depends. It is completely used up and the reaction stops. Excess reactant is a reactant which does not finish and part of it remains after the reaction.

DETERMINATION OF A LIMITING REACTANT:

The reactant which produces less amount of a product is the limiting reactant. You need to compare the two reactants in order to determine the limiting reactants through calculation. This is done by dividing the amount of each reactant by its formula mass i.e. the mass of its one mole. The reactant with a smaller number of moles is the limiting reactant.

Examples;

1. Hydrogen and oxygen are heated together to produce water vapour according to the equation



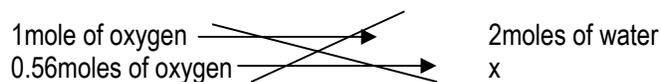
If 6g of hydrogen and 18g of oxygen were heated together. What was the mass of water vapour produced?

Solution;

$$\begin{aligned} \text{Number of moles of hydrogen in the reactant} &= \frac{6\text{g}}{2} \\ &= \underline{3 \text{ moles}} \text{ of hydrogen} \end{aligned}$$

$$\begin{aligned} \text{Number of moles of oxygen in the reactant} &= \frac{18}{32} \\ &= \underline{0.56 \text{ moles of oxygen}} \end{aligned}$$

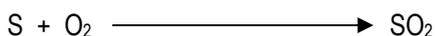
The smallest number of moles is on oxygen and therefore the mass of water was determined by the amount of oxygen.



$$\begin{aligned} X &= 0.56 \times 2\text{moles} \\ X &= \underline{1.12\text{moles}} \text{ of water produce} \end{aligned}$$

$$\begin{aligned} \text{Mass of water vapour produced} &= \text{number of moles} \times \text{formula weight of water} \\ &= 1.12 \times 18\text{g} \\ &= \underline{20.16\text{g}} \end{aligned}$$

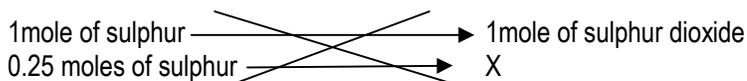
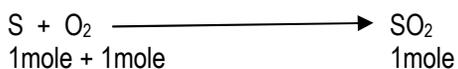
2. 8g of sulphur is heated using 16g of oxygen. Find the mass of sulphur dioxide produced. The chemical equation is written below;



Solution;

$$\begin{aligned} \text{Number of moles of sulphur} &= \frac{8}{32} \\ &= 0.25 \text{ moles of S} \end{aligned}$$

$$\begin{aligned} \text{Number of moles of oxygen} &= \frac{16}{32} \\ &= 0.5 \text{ moles of O}_2 \end{aligned}$$

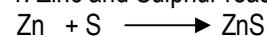


$$X = 0.25\text{moles of sulphur dioxide}$$

$$\begin{aligned} \text{Mass of sulphur dioxide produced} &= 0.25 \times 32 \\ &= \underline{16\text{g}} \text{ of sulphur dioxide was produced} \end{aligned}$$

Exercise

1. Zinc and Sulphur react to form Zinc Sulphide as follows:



In an experiment 12.0g of Zinc are mixed with 6.5g of sulphur and allowed to react.

a. Determine

- (i) Limiting reactant
- (ii) Reactant in excess
- (a) How many moles of Zinc were produced?
- (b) How many grams of the reactant in excess remained after the reaction?

2. 75g of Magnesium bromide, $MgBr_2$ was reacted with 150g Silver nitrate $AgNO_3$ according to the equation.



- (a) Determine
 - (i) Limiting reactant
 - (ii) Reactant in excess.
- (b) What mass of Silver Bromide was produced?.
- (c) What mass of the reactant in excess remained?
- (d) Calculate the moles of Magnesium nitrate $Mg(NO_3)_2$ produced in the reaction.

PERCENTAGE YIELD:

In all chemical reactions the amount of the product formed is almost always less than what we expect. This is due to a number of factors such as:

- a. Side reactions
- b. Wastage due to handling
- c. Incomplete reactions
- d. Not all reactants are transferred into the reaction vessels.

The expected amount of product is called the **theoretical yield** while the amount got is called is called the **actual yield**. In other words Percentage yield is the measure of the difference between the actual (practical) yield and the theoretical yield expressed in percentage.

- **Theoretical Yield:** This is the maximum amount of products calculated in a balanced chemical equation.

Calculations of the amount of products from the specified amounts of reactants using a balanced chemical equation are theoretical.

- **Actual Yield:** It is the amount of products obtained by performing a chemical reaction (practical or experiment).

The actual yield or percentage yield is calculated as follows;

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

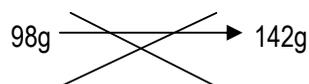
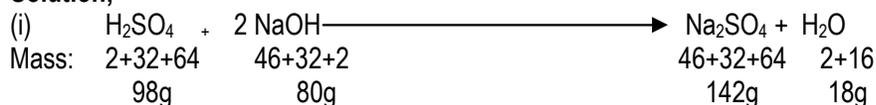
Example;

1.49g of sulphuric acid reacted with excess sodium hydroxide to produce sodium sulphate and water, according to the equation:



- (i) Determine the mass of sodium sulphate produced
- (ii) If the mass of sodium sulphate produced is 69.8g calculate the percentage yield

Solution;



$$49 \longrightarrow x$$

$$98x = 49 \times 142$$

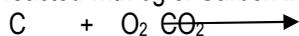
$$x = \frac{49 \times 142}{98}$$

$$x = \underline{71\text{g}}$$

$$\begin{aligned} \text{(ii) Percentage yield} &= \frac{\text{Actual yield} \times 100}{\text{Theoretical yield}} \\ &= \frac{69.8 \times 100}{71.0} \\ &= \underline{98.31\%} \end{aligned}$$

Exercise

- i. Oxygen reacted with 6g of Carbon in the Car engine, 20g of carbon dioxide were produced.



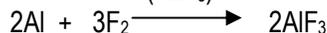
What was the percentage yield?

- ii. Dinitrogen Tetrafluoride, N_2F_4 was obtained from the reaction between NH_3 and 14.00g of Fluoride gas (F_2).

Calculate the percentage yield for this reaction if 4.80g of N_2F_4 was obtained:



- iii. Aluminium Fluoride (AlF_3) is obtained from the reaction of 13.4g of Aluminium in the reaction;



Calculate the percentage yield if 40.0g of AlF_3 were actually produced.

- iv. 95% of Copper (II) Oxide was produced when 6.4g of Copper were heated in the air:



What was the actual yield?

PURITY PERCENTAGE:

Purity percentage is a degree of how pure a substance is, expressed in percentage. Usually substances are not 100% pure, they contain some other materials. In the reaction (equation) we test how pure the reactant is, based on the actual products produced practically. Percentage yield can be regarded as a measure of purity.

$$\% \text{ Purity} = \frac{\text{amount of pure substance}}{\text{amount of impure substance}} \times 100$$

Exercise

1. Sample of black Copper (II) Oxide was in the form of thin pieces of wire. When ammonia was passed over 1.0g of the warm oxide, 0.68g of Copper metal was formed. What was the purity percentage of the oxide?



2. In titration 25.0cm³ of 1.0mol/dm³ NaOH reacted exactly with hydrochloric acid with a concentration of 1.0mol/dm³ which is 24.0cm³. What was the percentage purity of NaOH?
3. A 5.00g sample of impure barium carbonate was analyzed chemically in order to determine its purity percentage. 5.00g of impure barium carbonate was reacted with excess dilute nitric acid. 480cm³ of carbon dioxide gas measured at r.t.p. were formed. The equation for the reaction is given below:



Calculate;

- (i) The mass of barium carbonate present in the impure BaCO_3
 - (ii) Percentage purity of barium carbonate
 - (iii) The maximum mass of barium nitrate that can be crystallized from the resulting aqueous solution
- 4 In a titration, 25.0cm^3 of 1.0M NaOH reacted exactly with 11.5cm^3 of Sulphuric acid, H_2SO_4 with concentration of 0.1M .

What is the purity percentage of NaOH ?

GRADE 11 TERM ONE

Types of oxides.

Basically there are four types oxides namely acidic oxides, basic oxides, amphoteric oxides and neutral oxides.

Acidic oxides

These are oxides of nonmetals. They are formed when nonmetals burn in air. These oxides when dissolved in water they form acidic solutions. Examples of acidic oxides include sulphur dioxide and carbon dioxide.

Basic oxides

These are metal oxides and if soluble in water form an alkaline solution. Metals of group one and group two form basic oxides that are soluble in water. Some transition elements form insoluble basic oxides. Insoluble basic oxides include copper (II) oxide and the oxides of Iron.

Amphoteric oxide:

These are metal oxides that have both acidic and basic properties i.e. they behave as acidic oxides in a basic medium and behave like basic oxides in an acidic medium. Amphoteric oxides include aluminium oxide and zinc oxide.

Neutral oxides

These oxides are neither acidic nor basic. They include dihydrogen oxide or water and tin oxide.

ACIDS, BASES, ALKALIS AND SALTS

ACIDS:

A acid is a compound which when dissolved in water produces hydrogen ions as the only positively charged ions or it is a substance that donates a proton willingly or it is a substance that accepts an electron.

COMMON ACIDS:

S/NO.	NAME	FORMULA	IONS	
1	Nitric Acid	HNO ₃	H ⁺	NO ₃ ⁻
2	Hydrochloric acid	HCL	H ⁺	Cl ⁻
3	Sulphuric acid	H ₂ SO ₄	2H ⁺	SO ₄ ²⁻
4	Phosphoric (v) acid	H ₃ PO ₄	3H ⁺	PO ₄ ³⁻
5	Ethanoic acid	CH ₃ COOH	H ⁺	CH ₃ COO ⁻
6	Carbonic acid	H ₂ CO ₃	2H ⁺	CO ₃ ²⁻

- The process by which ions are formed is called ionization.
- Hydrogen ions give a characteristic of an acid.

STRONG AND WEAK ACIDS:

A strong acid is an acid which completely ionizes in aqueous solution. Examples are HNO₃ and HCl. A weak acid is an acid which partially ionizes in aqueous H₂CO₃, CH₃COOH, H₃PO₄ and all organic acids in general are weak acids.



PHYSICAL PROPERTIES OF AN ACID:

1. It has a sour taste.
2. It is corrosive (it can wear out a metal or skin)
3. It is soluble in water
4. It is poisonous.
5. It is colourless

CHEMICAL PROPERTIES

1. It shows an action on indicators e.g. It turns blue litmus paper to red or pink.
 - (a) It turns methyl orange to red.
 - (b) It is colourless in the presence of phenolphthalein.

2. It reacts with reactive metals to produce salt and hydrogen gas.



Ionically:



Note equation:



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

Example:



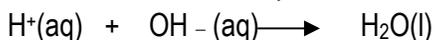
Ionicly:



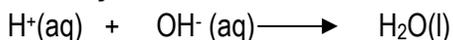
Not equation:



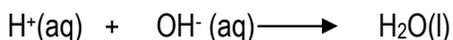
4. If reacts with a base to produce salt and water only.



Ionicly:



Not equation:



BASICITY OF AN ACID

Basicity is a number of hydrogen ions that can be formed from one molecule of the acid in solution.

BASICITY COMMON ACIDS

S/NO.	FORMULA	BASICITY	TYPE
1.	HCl	1	Mono basic acid
2.	HNO ₃	1	Mono basic acid
3.	CH ₃ COOH	1	Mono basic acid
4.	H ₂ SO ₄	2	Di basic acid
5.	H ₂ CO ₃	2	Di basic acid
6.	H ₃ PO ₄	3	Tri basic acid

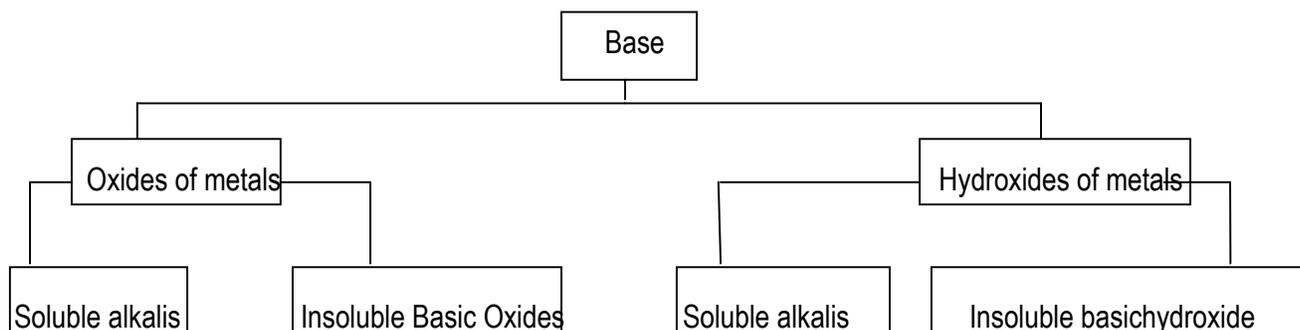
BASES AND ALKALIS

BASES:

- Bases are oxides, hydroxides of metals and ammonia
- A base is a Proton acceptor
- A base is an electron pair donor.

Soluble bases are called alkalis. An alkali is a compound which produces hydroxide ions (OH⁻) as the only negatively charged ions in solution.

NB: All alkalis are bases but not all bases are alkalis.



A base is a substance that when dissolved in water produces hydroxide ions (OH⁻) as the only negatively charged ions or a base is a proton acceptor. Generally bases are classified as metal oxides, metal hydroxides and Ammonium hydroxides. Other bases include Carbonates and Bicarbonate. Soluble bases dissolve in water to form hydroxide ions (OH⁻) as the only negatively charged ions and known as alkalis.

PHYSICA PROPERTIES OF BASES:

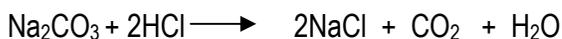
1. Some bases dissolve in water e.g Sodium hydroxide and Potassium hydroxide.
2. Bases when in solution feel soapy
3. Soluble bases have a bitter taste.
4. Soluble bases turn red litmus paper blue.
5. Soluble bases can corrode the skin.
6. Bases such as Sodium hydroxide can corrode glass ware.

The table below shows some common bases and their solubility in water.

NAME OF THE BASE	CHEMICAL FORMULA	SOLUBILITY IN WATER
Sodium oxide	Na ₂ O	Very soluble
Potassium oxide	K ₂ O	Very soluble
Sodium hydroxide	NaOH	Very soluble
Potassium hydroxide	KOH	Very soluble
Ammonia	NH ₃	Very soluble
Ammonium hydroxide	NH ₄ OH	Very soluble
Calcium oxide	CaO	Slightly soluble
Calcium hydroxide	Ca(OH) ₂	Sparingly soluble
Magnesium oxide	MgO	Slightly soluble
Magnesium hydroxide	Mg(OH) ₂	Slightly soluble
Copper oxide	CuO	Insoluble
Iron III hydroxide	Fe(OH) ₃	Insoluble
Copper hydroxide	Cu(OH) ₂	Insoluble

CHEMICAL PROPERTIES OF BASES

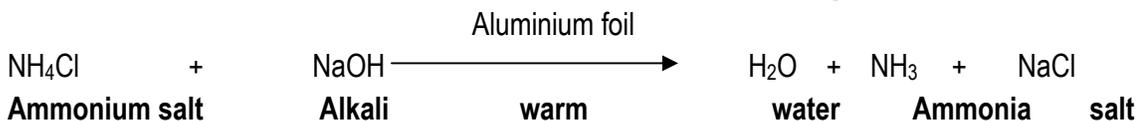
1. Bases react with acids to form salts, water and carbon dioxide e.g.



2. Alkalis react with acids to form a salt and water only the reaction is called Neutralization e.g.



3. Alkalis react with Ammonium salts to form Ammonia a salt and water e.g.



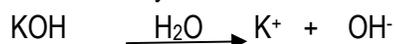
STRONG AND WEAK BASES

A strong alkali or base is one that when dissolved in water undergoes complete ionisation. Examples of strong alkalis are:-

(i) Sodium hydroxide;



(ii) Potassium hydroxide



They are strong alkalis because when dissolved in water they are crystal lattice break up completely to form ions.

Weak bases are alkalis on the other hand when dissolved in water only ionizes partially. For example ammonia gas produces few hydroxide ions when dissolved in water compared to a strong soluble of the same concentration.



Partially ionizes

Preparation of insoluble hydroxide

Insoluble hydroxides such as copper (II) hydroxide and iron (II) hydroxide can be prepared by precipitation method as shown in the chemical equation below;



In the above equation, iron (II) hydroxide is insoluble in water. The process is started with two soluble compounds; one to provide the hydroxide and the other to provide the metal which would combine with the hydroxide to form the insoluble hydroxide. The soluble compound is collected as a filtrate after filtering and the insoluble hydroxide remains on the filter paper, it is washed in distilled water and dried.

INDICATORS:

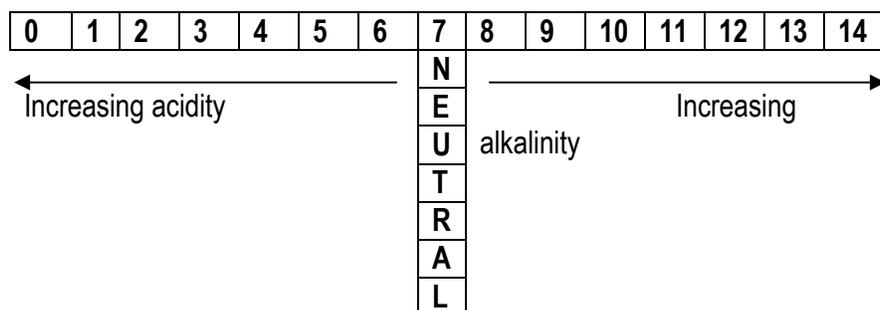
Indicators are substances that show by change whether the solution is acidic or basic. They are actually weak acids that can be used to distinguish acids from bases. Most indicators show a sharp, colour change in acids and bases except universal indicator that show the degree of acidity or alkalinity.

Some common indicators and their colour changes are shown in the table below.

INDICATORS	COLOUR CHANGE IN ACID	COLOUR CHANGE IN BASE
Litmus Paper	Red/Pink	Blue
Phenolphthalene	Colourless	Pink
Methylorange	Pink	Yellow
Universal indicator	Red at Ph ₄	Violet at Ph ₁₀

THE pH SACLE (POTENTIAL OF HYDROGEN)

The universal indicator is a mixture of indicators. The resultant colour change shows the degree of acidity alkalinity. Acidic substances have a pH less than 7 Neutral substances have a pH of 7 and basic or alkaline substances have a pH greater than 7. PH of the soil is very important in agriculture as each crop does well in a particular PH range. Lime is added to acidic soils to neutralize the acid. Quick lime or slaked lime is used because it is not very soluble in water hence it does not increase the PH of the soil or increase the alkalinity of the soil.



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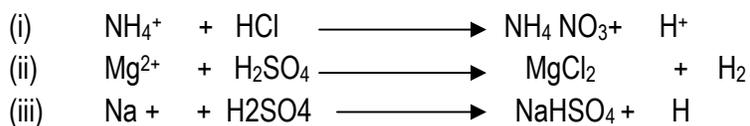
DILUTING AN ACID

There are many times we will be required to dilute acids to a certain morality or concentration. The procedure for diluting acids is that the acid shall be added to water first and not verse versa. After which the mixture can be re-diluted further.

When enough dilution has taken place water can now be added to the acid. The above procedure is applicable whenever we are diluting concentrated acid.

SALTS:

A salt is a compound formed when all the ionisable hydrogen of an acid are partially or wholly replaced by either the metallic ion or ammonium ion. Or a salt is a compound formed when part of the ionizable hydrogen ions have been replaced by a metallic ion or ammonium ion e.g.



The table below shows some acids and salts made from them:-

S/NO.	ACID	RADICAL	NAME OF SALT	EXAMPLE
1	Hydrochloric acid	Cl^-	Chlorides	MgCl_2
2	Sulphuric acid	SO_4^{2-}	Sulphates	Na_2SO_4
3	Nitric acid	NO_3^-	Nitrates	$\text{Zn}(\text{NO}_3)_2$
4	Carbonic acid	CO_3^{2-}	Carbonates	CaCO_3
5	Ethanoic acid	CH_3COO^-	Ethanotes	CH_3OONa
6	Phosphorus acid	PO_4^{3-}	Phosphates	Na_3PO_4

NAMING SALTS

When naming salts; the cations are stated first the Anion second. The Cation is either a metallic ion or an ammonium ion and the Anion is the radical from the acid.

TYPES OF SALTS

Generally there are two types of salts namely, Normal salts and acidic salts. Normal salts are formed when all the ionisable hydrogen ions of an acid have been replaced by a metallic ion in ammonium ion, On the other hand, acidic salts are salts formed when one part of the ionisable hydrogen ion of acid have been replaced by the metallic ions or ammonium ions.

The table below shows examples of normal salts and acidic salts.

S/NO.	NAME OF SALT	CHEMICAL FORMULA	TYPE OF SALT
1	Sodium Chloride	NaCl	Normal salt
2	Calcium Sulphate	CaSO_4	Normal salt
3	Zinc Nitrate	$\text{Zn}(\text{NO}_3)_2$	Normal Salt
4	Magnesium hydrogen Sulphate	MgHSO_4	Acidic salt
5	Potassium hydrogen Carbonate	KHCO_3	Acidic salt
6	Sodium hydrogen Carbonate	NaHCO_3	Acidic salt
7	Sodium hydrogen Sulphate	NaHSO_4	Acidic salt

BASIC SALTS

Basic salts are salts that contain OH^- ions e.g. $\text{Cu}_2(\text{OH})\text{CO}_3$ Copper hydroxide Carbonate. All basic salts can be bases as well as salts.

PREPARATION OF SALTS

There are many methods of preparing salts. But whatever method is to be used, it will depend on whether that salt is soluble or insoluble in water.

The table below summarizes soluble and insoluble salts.

Soluble salts	Insoluble salts
All common salts of sodium, potassium and ammonium	All carbonates are insoluble except those of Na, K and NH_4^+
All nitrates	-
All chlorides are soluble except those of silver and lead	Chlorides of silver, and lead
All sulphates are soluble except sulphates of barium, calcium and lead	Sulphates of lead, barium and calcium
All carbonates of sodium, potassium and ammonium	All carbonates are insoluble except carbonates of sodium, potassium and ammonium
All bi carbonates	-

Some of the methods of preparing salts are;

1. Direct combination
2. Displacement reaction
3. Neutralization
4. Action of an acid on an insoluble base
5. Double decomposition or precipitation
6. Action of an acid on a carbonate

A. DOUBLE DECOMPOSITION (PRECIPITATION)

An insoluble salt such as lead sulphate, barium sulphate and lead iodide can be prepared by double decomposition or precipitation method. In this method, two solutions of the soluble salts containing the desired radical and a cation are mixed to produce an insoluble salt and a soluble salt. The needed salt is the insoluble salt. The precipitated insoluble salt is filtered off, washed in distilled water and dried.

Experiment;

TITLE: *preparation of salts by double decomposition*

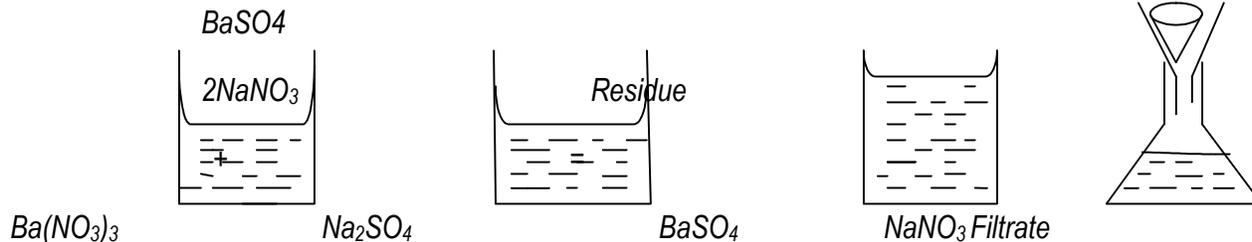
AIM: *To prepare Barium sulphate*

APPARATUS: *Two beakers, a funnel, a stirring rod and a conical flask*

MATERIALS: *Filter paper, barium nitrate, sodium sulphate, distilled water*

PROCEDURE: *Barium nitrate and sodium sulphate are dissolved in beakers of water separately by stirring the solutes in distilled water using a stirring rod. The two solutions are then mixed, stirred and filtered off. Distilled water is added to the residue on the filter paper so that any remaining soluble particles dissolve and drain into the conical flask. To ensure that all the nitrate and sodium ions have been washed out a test for nitrate and sodium ion is done i.e. sodium ions use a flame test and nitrate test for ammonia gas. (See notes on identification*

of ions). The insoluble salt on the filter paper is then dried and stored. The diagrams below illustrate the procedure.



Barium sulphate salt

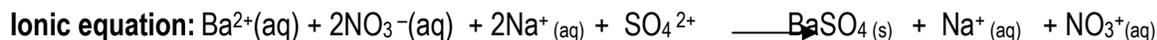
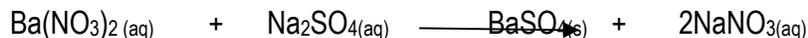
OBSERVATIONS: When barium nitrate and sodium sulphate were mixed a white precipitate was formed. After filtering an insoluble white precipitate called the residue remained on the filter paper and a colourless liquid called a filtrate collected in the conical flask.

CONCLUSION: Barium sulphate can be prepared by double decomposition method.

EXPLANATIONS: In double decomposition method the radicals exchange cations. However, the ions of the soluble salt do not chemically combine in solution but just remain as spectator ions. Many other insoluble salts such as lead iodide and silver chloride can be prepared by using similar methods

CHEMICAL EQUATIONS:

In all chemical reactions, chemical equations are used to describe and summarize what has taken place. In the above reaction; the chemical equation is:



A. NEUTRALIZATION OR ACTION OF AN ACID ON AN ALKALI

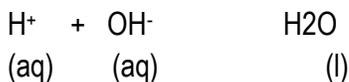
Preparation of salt by neutralization (Titration). In this case titration would mean to find the amount of the needed to neutralize a known amount of an alkali or vice versa by using an indicator. The indicator will show the end point by the colour change.

In the acid-base titration, the base is pipette usually 25cm³ of the base is pipette and put in the conical flask and the base is called the titrand. The acid is called the titrant. The acid (titrant) is slowly and carefully added to a measured volume of alkali (base) using a burette until the indicator;

usually phenolphthalein changes colour. An indicator is used to show when the alkali has just been neutralized completely by the acid at the end-point.

Once we have established the end point, we can add the same volume of the alkali, but this time without the indicator. The solution that is produced can then be evaporated slowly to obtain a pure salt i.e. a salt without indicator impurities.

This neutralization reaction involving the acid and the base can best be described using the ionic equation.



Experiment

TITLE: *Preparation of salts by neutralization.*

AIM: *To prepare sodium chloride.*

APPARATUS: *Two beakers, a funnel, a conical flask, white tile, burette, pipette, pipette filler, clamp stand.*

MATERIALS: Sodium hydroxide of known concentration, hydrochloric acid indicator.

Methods: 25cm³ of Sodium Hydroxide is pipetted into conical flask and few drops of phenolphthalein is pink in alkaline solutions and colourless in an acid.

The acid (hydrochloric acid) is added into the Sodium hydroxide solution in small volumes from the burette. The content of the conical flask must be swirled after each addition of the acid for thorough mixing. The alkali has been neutralized completely at the end point. This is shown by the pink colour of the indicator just disappearing, becoming colourless. The final volume reading on the burette is recorded. Further titration runs are carried out until consistent volumes and end point are obtained. Three (3) consistent results or values are used to calculate the average volume of the acid; this value is used to calculate the concentration of the acid. This diagram below shows how the experiment is done.

Observation:

At end point the colour of the solution in the conical flask changed from pink to colourless when equivalent volume of the acid and the base were mixed and water evaporated, crystals of sodium chloride formed.

Equivalent amounts of the acid and the base can now be taken and mixed to form a salt and water only. The water can then be evaporated and crystals of sodium chloride formed.

RESULT OF TITRATION

Range	1	2	3	4
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Final	15.8	31.5	46.4	30.8	46.5
Initial	0.00	15.8	31.5	15.0	30.8
Volume used	15.8	15.7	14.9	15.8	15.7

$$\frac{15.7 + 15.8 + 15.7}{3}$$

$$\frac{47.2}{2}$$

15,733

15.7cm³

Examples:-

Find the molarity of the acid if the molarity of the base is 0.2.

Number of moles of base = number of moles of acid

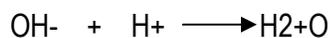
$$M_b \times V_b = M_a \times V_a$$

$$0.2 \times 25\text{cm}^3 = 15.7\text{cm}^3 \times \text{Molarity}$$

$$\frac{5\text{moles}}{15.7\text{dm}^3} = \text{Molarity}$$

Molarity = 0.32mole/dm³

2. Write the ionic equation for this reaction $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
(aq) (aq) (aq) (aq)



Is the salt formed acidic or basic. State the reason for your answer as it is slightly acidic because in the presence of phenolphthalein it is colourless in an acid media at the end the solution is colourless.

Preparation of salt by the action of an acid on an insoluble base

This method is mostly used mostly to prepare salts of un reactive metals such as copper, lead and oxides of iron because in this case it is not possible to use direct reaction of the metal with an acid so the acid is neutralized by using the particular metal oxide:

Experiment:

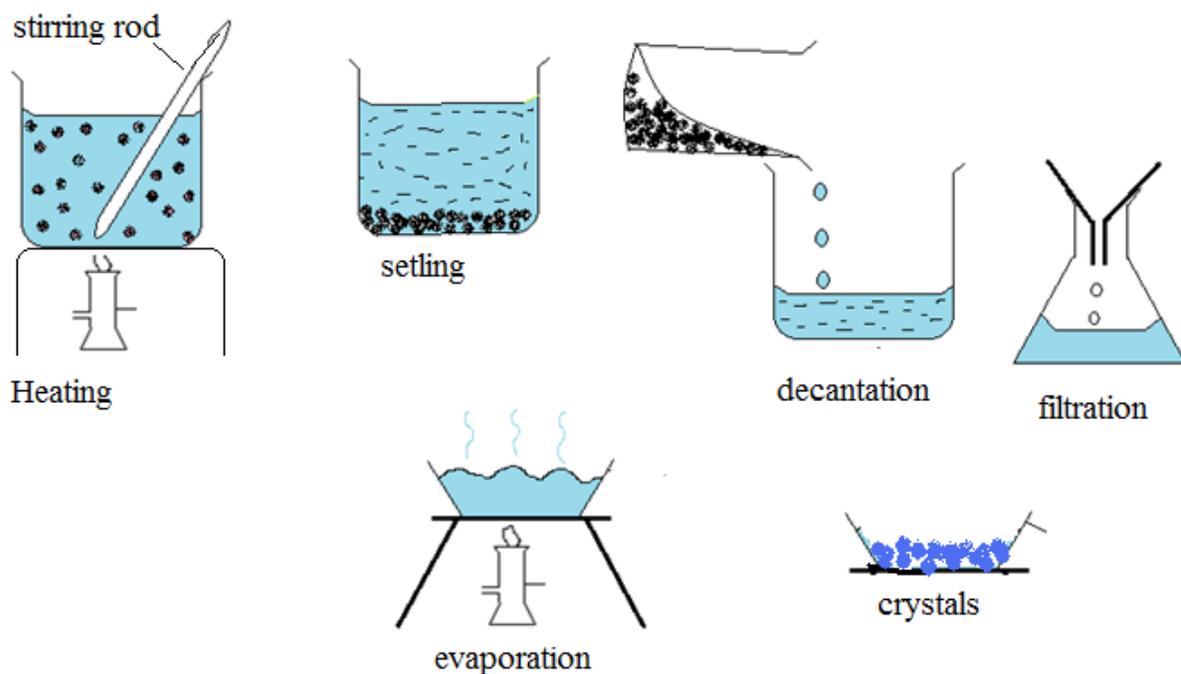
Title: Preparation of salts

Aim: To prepare copper (II) sulphate by the action of sulphuric acid on an insoluble copper (II) oxide.

App: two beakers, evaporation dish, conical flask, funnel, stirring rod, Bunsen burner, tripod stand and a spatula

Materials: distilled water, copper (II) oxide, sulphuric acid and filter paper.

Method; Sulphuric acid is heated and while heating the acid copper (II) oxide was being added slowly using the spatula. As the oxide was added the mixture was stirred to facilitate reaction. When no more copper (II) oxide could dissolve, the mixture was removed from the heat and allowed to settle. When the mixture had settled it was decanted and filtered. The distillate was heated in an evaporation dish to drive off excess water to produce a saturated solution. The hot saturated solution was allowed to cool slowly and form good crystals. The diagram below shows the staged in the preparation of copper (II) sulphate.



Observations

When copper (II) oxide was added to the heated sulphuric acid and stirred the solution turned dark-blue. When the mixture was filtered a clear blue solution was collected. When excess water was evaporated and the hot saturated solution allowed to cool slowly blue crystals of copper (II) sulphate formed in the evaporation dish.

Explanations;

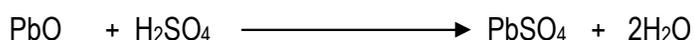
Sulphuric acid had to be heated because copper only reacts with hot acids and the stirring was done to speed up the reaction. Furthermore, copper oxide is insoluble in water so the stirring and heating helped

the copper (II) oxide to dissolve in the acid. The mixture was allowed to settle so that the unreacted dense particles of copper (II) oxide could sink to the bottom. Filtrate was done to separate residues from the clear solution called the filtrate. The filtrate was heated to drive off excess water and make the solution saturated. The saturated solution was allowed to cool slowly so that big crystals with good geometrical shapes could form.

In the above experiment, the chemical equation is:



Other insoluble bases that can be can react in a similar way include;



Preparation of salts by the action of an acid on a carbonate

This method can be used with any metal carbonate and any dilute acid provided the salt produced is soluble in water. The limitation to this method happens when the salt formed is insoluble in water because the insoluble salt formed coats around the un reacted carbonate and the reaction stops. To overcome this limitation continuous stirring should be done to expose the un reacted carbonate to the acid.

Experiment:

Title: Preparation of salts

Aim: To prepare copper nitrate by the action of nitric acid on copper carbonate

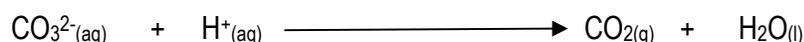
App: Beaker, and stirring rod, spatula, conical flask, funnel

Materials: copper carbonate, dilute nitric acid, filter paper, lime water

Method: Dilute sulphuric acid was put in a beaker and copper carbonate was slowly added to the acid using a spatula and stirring gently. When no more effervescence could be seen the solution was decanted and filtered. The filtrate was copper (II) nitrate. The filtrate was heated to evaporate excess water and the saturated solution allowed to cool slowly to form crystals.

Observation: When copper carbonate was added to the acid there was effervescence and a colourless gas was given out. The colourless gas turned lime water milky.

Explanation: Effervescence was due to the production of carbon dioxide. Metal carbonate contain carbonate ions (CO_3^{2-}). In this reaction the carbonate ions reacted with the hydrogen ions from the nitric acid to form carbon dioxide and water according to the equation:

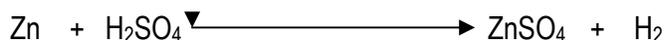


The over equation is:



Preparation of salts by the action of an acid on a metal

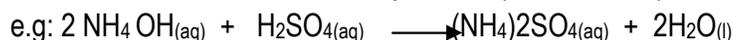
Certain salts can be prepared by reacting dilute acids with moderately reactive metals such as magnesium and zinc. In this method a fairly reactive metal displaces the hydrogen from the acid. In most cases the salt to be prepared should be soluble in water. For example;



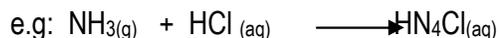
AMMONIUM SALTS

Ammonium salts can be prepared by:

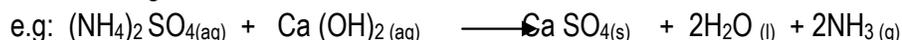
- (i) A reaction between ammonium hydroxide (a weak alkali) and a dilute acid.



- (ii) A reaction between ammonium gas and dilute acid.



Since ammonium salts are acidic they can react with alkalis to produce normal salt, water and ammonium gas.



Ammonium gas turns a damp red litmus paper to blue. This shows that ammonia is alkaline gas. NOTE: It is not advisable to apply ammonium sulphate fertilizer at the same time with calcium hydroxide. Fertilizer why?

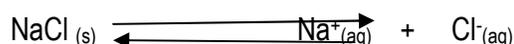
This is because the two will react to produce an insoluble salt and ammonia gas which will escape with Nitrogen into the atmosphere. The crops will remain without nitrogen, a very important crop nutrient.



Solubility of salts.

Salts are made up of positive and negative ions which are bound together by the force of attraction of their opposite charges. If the energy needed to break their ionic bonds is lower than the energy given off by

interaction of the ions with water the salt dissolves quickly. At the start the salt dissociates quickly and the reaction is generally going one way. As more and more salt is added to the solvent, the solution gets more and more saturated. Eventually so many ions are dissolved in the solution that the reverse reaction starts to take place i.e. some ions starts to re-make the bonds that had existed before the salt was added to water. The reverse reaction decreases the speed at which water dissolves the next particles of the salt. Finally the reverse reaction reaches the rate of dissociation of the salt. In other words the rate of the reverse reaction becomes equal to the rate at which the salt dissolves in water. When the two rates are equal the solution reaches equilibrium and the solution is saturated. As saturated solution is one that cannot dissolve any more solute and the extra solute added increases the reverse reaction where the extra solute in this case salt precipitates out of the solution and collects at the bottom of the vessel. In case of sodium chloride at equilibrium the situation is presented as follows:



Solubility of salts is defined as an amount of salt that needs to be added to distilled water for the solution to reach equilibrium. Solubility product constant is a simplified solubility constant (KSP) between the solid particles of a salt and their respective ions in the solution. Its value indicates the degree into which a salt dissociates in water. The higher the solubility product constant the more soluble the salt is. The KSP expression for a salt is the product of the concentration of the ions, with each concentration raised to a power equal to the coefficient of that ion in the balanced chemical equation for the solubility equilibrium. The general formula is written as:

$$\text{KSP} = \frac{[\text{A}]^x [\text{B}]^y}{1}$$

key x is the coefficient of A
y is the coefficient of B

Effects of acids on the environment

Acids have a PH less than 7. Sources of acids in the environment include;

Acidic oxides from the burning of nonmetals in air

- (a) Acidic effluents discharged into the environment by manufacturing industries
- (b) Acidic fertilizers applied on farmlands
- (c) And acidic liquids from school laboratories washed down sinks

Acids have the following adverse effects on the environment;

1. Acidic rain;

Rain is usually slightly acidic with a PH a little less than 7. This is because carbon dioxide in air dissolves in rain water to form a weak carbonic acid. Sulphur dioxide from the burning of fossil fuels such as coal and nitrogen dioxide from hotter parts of engines. When these oxides dissolve in rain water they form an acid rain which can have a PH of much less than 7. The acid rain can cause the following;

- a. Corrode iron roofs of buildings

- b. damage the plants by making the soil acidic. Furthermore, the acidic rain water washes out ions such as magnesium ions and calcium ions and the soil become short of these ions necessary for plant growth.
- c. acid rain may irritate people's skin leading to scratching and rash.
- d. Acidic rain corrode marble stone works

2. Acidic effluents;

Acidic effluents from manufacturing and extraction industries may get into rivers and lakes. These lower the PH of water. At a lower PH of 4 aquatic plants and animals are killed for example fish lose their slime layer and scales leading to sores on fish body.

3. Acidic fertilizers

Fertilizers are salts and hence there are normal fertilizers and acidic fertilizers. When acidic fertilizers are applied on a farmland they will make the soil acidic and kill microorganisms found in the soil. In addition if acidic fertilizers get washed into rivers and lakes they cause eutrophication.

Uses of acids

Despite the harmful effects that acids have on the environment, they are very useful and some of the uses are;

- a. sulphuric acid is used as a drying agent in the manufacture of fertilizers
- b. Acids such as sulphuric acid is used in car batteries where it is used as an electrolyte.the battery is known as lead-acid accumulator
- c. Sulphuric acid is used in larger amounts in the iron and steel industry to remove rust and oxidation
- d. weak acids such as lemon juice or vinegar are used to remove lime scales.
- e. Tartaric acid is added to baking powder to raise the dough and it acts as a preservative.
- f. Acids are used to treat wasp sting as was inject an alkali when they sting

Uses of alkalis

- 1. Slaked lime (calcium hydroxide) is used to neutralize acidic industrial waste
- 2. Alkalis are used in the manufacture of lubricants because they feel soapy and slippery
- 3. Alkalis are used in making soaps for example sodium hydroxide is used to make bathing soap and potassium is used to make shampoos

- 4 Alkalis are used to neutralize acids
5. Alkalis are added to washing detergents to provide a suitable PH for the enzymes to digest certain stains.
6. Alkalis are added to water in rivers and lakes where acid rain has lowered the PH of the water
7. Calcium carbonate is used in paper industries as a bleaching agent.

Uses of salts

The table summarizes the use of some salts

SALT	USE
Sodium Chloride (NaCl)	For food flavouring, making hydrochloric acid hospital saline
Ammonium Sulphate	Fertilizer for crops
Ammonium Chloride	In torch batteries as electrolyte.
Calcium Carbonate	Iron extraction, making cement and glass.
Calcium Chloride	Drying agent, extraction of Sodium.
Calcium Sulphate	Plaster casts for injured limbs.
Iron (II) Sulphate	In iron tablets
Magnesium sulphate	In mechanics.
Potassium Nitrate	In fertilizer and gas powder.
Silver bromide	In photography to darken the film.
Sodium carbonate	Glass making softening water making washing powder.
Iron (II) fluoride	Additive toothpaste.
super phosphate of ammonia	a very good fertilizer

Water of crystallization

Most salts form crystals which contain water. The water is loosely bonded to the crystals and it is always there in the same fixed ratio. The water of crystallization in salts of transition elements imparts colour. For example the crystals of copper (II) sulphate always form a blue chemical compound with the formula of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. Salts with water of crystallization are called hydrated salts and crystals that have lost their water of crystallization are called anhydrous salts. Crystals of hydrated salts should not be heated to dryness because they would lose their colour and shapes get distorted

Determination of water of crystallization

We can determine the water of crystallization by heating hydrated crystals of a salt. When hydrated crystals are heated their water of crystallization evaporate as steam. This steam can be made to condense into liquid water if the steam is passed through a condenser. The crystals lose their colour and become colourless which is the anhydrous form of the crystals. For example if we heat blue copper (II) sulphate ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$) which is hydrated it becomes whitish or colourless (CuSO_4) anhydrous salt. If we add water to the anhydrous whitish powder of copper(II) sulphate it becomes blue again. The addition of water to the

whitish powder and the subsequent turning of the whitish powder into a blue colour is proof that the liquid that was lost during heating was water. Alternatively we can test the liquid that condensed if it is water by dipping into it a dry blue cobalt chloride paper and if the blue cobalt chloride paper changes colour from blue to pink then we can conclude that the liquid that was lost by the crystals is water.

Experiment:

Title: water of crystallization

Aim: To determine the mass of water of crystallization in copper (II) sulphate

App: Beaker, Bunsen burner, Triple beam balance

Materials: Blue copper (II) sulphate crystals,

Method: mass of empty dry beaker = 26.4g

Mass of beaker and blue copper (II) sulphate = 47.6

Mass of beaker and anhydrous copper (II) sulphate = 40.2g

Mass of water lost = 47.6 – 40.2g

= 7.4g

Mass of anhydrous CuSO₄ = 40.2 -26.4

= 13.8g

Mass of water of crystallization in one mole of anhydrous copper(II) sulphate;

7.4g of water ~~→~~ 13.8g

X ~~→~~ 160g

$$13.8X = 160 \times 7.4$$

$$X = 85.8$$

Number of moles of water in 85.8g = $\frac{85.8g}{18}$

$$= 4.77 \text{ moles}$$

The theoretical value of the number of moles of water of crystallization that combines with copper (II) sulphate is 5 moles. However in this experiment we could not arrive at 5 moles because of factors affecting the percentage yield.

TERMS ASSOCIATED WITH WATER OF CRYSTALLIZATION

1. Hygroscopic:

These are crystals that absorb water from the atmosphere but do not dissolve in the absorbed water.

2. Efflorescence:

These are salts or crystals that lose water to the atmosphere

3. Deliquescence:

These are salts or crystals that absorb water from the atmosphere and dissolves in the absorbed water

RATE OF REACTION

Chemical kinetics is the study of the speed with which a chemical reaction occurs. The reaction can either be fast or slow or moderate. The rate of a chemical reaction can be measured by the rate at which a reactant is used up or the rate at which a product is formed. The rate of a chemical reaction depends on a number of factors which include;

1. Concentration of the reactants
2. Catalyst
3. Temperature
4. Particle size
5. Pressure
6. Medium
7. Light

Concentration of reactants;

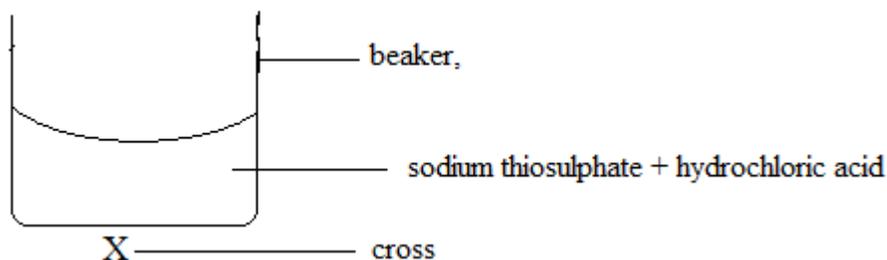
Concentration is the number of moles per cubic decimeter. This is the number of moles available for a chemical reaction to proceed. For any reaction to occur reacting particles must collide with each other. Increasing concentration increases the frequency of collision between the two reactants. A higher concentration of the reactants leads to more successful collision per unit time which leads to increased reaction rate.

The graph below shows the rates of reaction of different concentration of hydrochloric acid and excess sodium carbonate. Each reaction started with 2mols/dm^3 of hydrochloric acid.

Experiment

- Title:** Rate of reaction.
- Aim:** To investigate the effect of concentration on the rate of reaction
- App:** Three conical flasks, three beakers, a piece of white paper and a stop watch
- Materials:** Hydrochloric acid of 0.5mols/dm^3 and sodium thiosulphate of 1mol/dm^3 , 0.25mols/dm^3 and 0.063mols/dm^3 .
- Method:** 20cm^3 Sodium thiosulphate of 0.5mols/dm^3 , 0.25mols/dm^3 and 0.063mols/dm^3 were prepared and put into three separate beakers. A cross was made on a white piece of paper. The beaker with 1mol/dm^3 was placed on the white piece of paper such that the

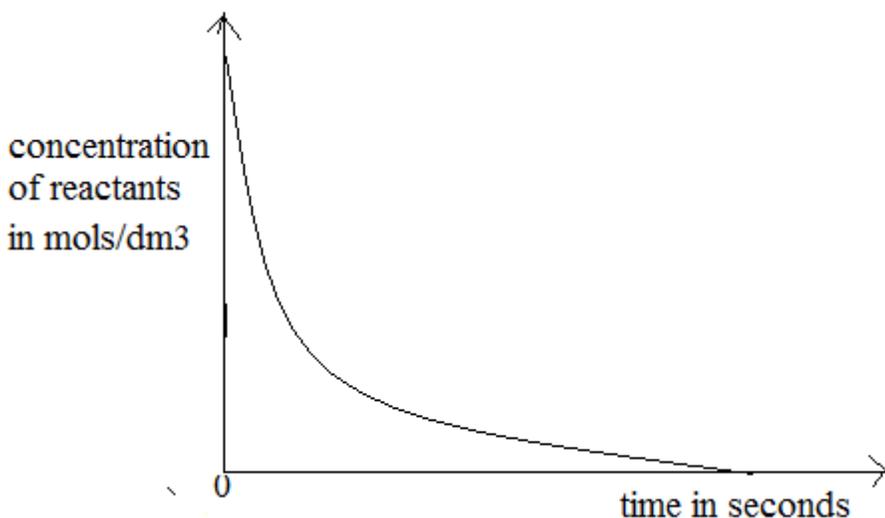
cross could be seen from above. 20cm³ Hydrochloric acid was added to each of the beakers containing sodium thiosulphate and at once the stop watch started. Immediately the cross disappeared, the stop watch is stopped. The time was taken for the cross to disappear for each concentration of sodium thiosulphate. The diagram below illustrates the method.



Observations: when the acid was added there was a yellow precipitate that formed and obscured the cross from above. The time taken for each concentration to obscure the cross was recorded as per the table below;

Concentration of Na ₂ S ₂ O ₃	1mol/dm ³	0.25mols/dm ³	0.065mols/dm ³
Time taken in seconds	16	35	57

From the above table a graph could be drawn.



Conclusion: The cross disappeared when enough sulphur had formed to blot out the cross. The solution with the highest concentration of sodium thiosulphate blotted out the cross first and the solution with the least concentration. Therefore, the higher the concentration the faster the rate of reaction.

Explanations: The solution of 1mol/dm³ sodium thiosulphate had more particles of sodium thiosulphate per cubic decimeter. There were more collisions between the hydrochloric acid particles and the sodium thiosulphate particles. These collisions

results resulted into more sulphur atoms being deposited at the bottom of the beaker. The equation below summarizes the chemical equation.



Surface area;

This is the space that a particle exposes for a chemical reaction to occur. Smaller particles react faster because they have a larger surface area compared to their volumes. Bigger particles react slowly because they have a smaller surface area compared to their volumes. For example the reaction between calcium carbonate powder with dilute hydrochloric acid is faster than the reaction between solid calcium carbonate (marble chips) with dilute hydrochloric acid of the same concentration.

Experiment:

Title: Rate of reaction

Aim: To determine the surface area that reacts faster

App: Two beakers, measuring cylinder and beam balance, stop watch

Materials: 0.25mol/dm³ of hydrochloric acid, powdered calcium carbonate and marble chips

Method: 5g of powdered calcium carbonate and 5g of marble chips were weighed and put into two separate beakers. 6cm³ of the Hydrochloric acid was added to the 5g powdered calcium carbonate at once and timed. Another 6cm³ of hydrochloric acid was measured using a measuring cylinder and added to the marble chips at once and timed. The time taken for effervescence to stop for each experiment was recorded as shown in the table below;

Results/Observations:

Surface area	Time taken for effervescence to stop
Powdered calcium carbonate	25seconds
Solid calcium carbonate(marble chips)	80secods

The powdered calcium carbonate reacted very fast and the reaction came to an end quickly. The solid calcium carbonate (marble chips) reacted slowly and took much longer time to finish

Conclusion:

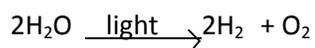
The reaction rate is faster in the powdered calcium carbonate than in the solid calcium carbonate.

Precaution:

Smaller particles offer a large surface area for a chemical reaction. The reaction is rapid and explosive especially if they have a low energy of activation. Particles that have large surface area should be handled with care and not to be exposed to naked fire especially if they have the particles of a material with low ignition point. Particles such as combustible gases in mines and flour particles in flour mills can easily catch fire and cause rapid explosion if exposed to naked flame. In flour milling plants people are not allowed to smoke inside the plant or carry matches into the plant. Similarly in underground mines where there are combustible gases such as methane gas, flames must be properly sealed so that gases do not come into contact with the flame

Light:

Some chemical reactions are affected by light. These include photosynthesis. Light is absorbed by green plants using chlorophyll. In green plants this light energy is used to split water molecules into hydrogen and oxygen as per equation below



When there is less light intensity the rate of photosynthesis proceeds slowly but when light intensity is high the rate of photosynthesis proceeds as a faster rate.

Pressure:

Pressure is due to a number of particles enclosed in a particular vessel. The rate of reaction increases when there are more particles enclosed in a vessel because they collide frequently. Frequent collision results into more successful reaction. Therefore increasing pressure increases the rate of reaction.

Temperature:

In physics temperature is the degree of hotness and temperature affects the kinetic energy that the particles possess. Warm or hotter reactants form products faster meaning that the rates of reactions are higher or they take less time to form products.

The reaction between Sodium thiosulphate ($\text{Na}_2\text{S}_2\text{O}_3$) and hydrochloric acid can also be used to study the effect of temperature on the rate of a reaction. In this experiment the quantities of reactants are kept exactly the same each time, only the temperature of the reactants is changed.

Experiment:

Title: Rate of reaction

Aim: To investigate the effect of temperature on the rate of reaction.

App: Five beakers, a stop watch and a thermometer.

Mat: Sodium thiosulphate and hydrochloric acid.

Method: Five beakers each containing 5cm³ of Sodium thiosulphate of 0.1molar solution were heated to different the temperatures: 20°C, 30°C, 40°C, 50°C, 60°C. Hydrochloric acid of 0.1molar was added quickly to each one of the them starting with the one with the lowest temperature and at the same time the stop watch was switched on. The stop watch was stopped at the moment the cross could no longer be seen from above.

Observation: As the reaction proceeded the cross became fainter and fainter, the cross eventually blotted out.

The table below shows the results of the effects of temperature on the rate of reaction.

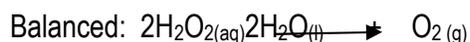
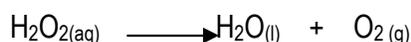
Temp (°C)	20	30	40	50	60	70
Time in seconds for the cross to blot	200	125	50	33	24	14

Conclusion: The higher the temperature the faster the cross disappeared. In other words the higher the temp the faster the rate of reaction.

Explanations: When the temperature at which the reaction is carried out is increased, the energy that the particles have is increased and the particles move faster. This increases the number of successful collisions of the reactants

Catalyst:

A catalyst is a substance which can change the rate of a chemical reaction without being chemically changed itself. In the laboratory, the effects of a catalyst can be observed by using the decomposition of hydrogen peroxide. Hydrogen peroxide decomposes into oxygen and water very slowly at room temperature and pressure according to the equation.



They are substances, however, that can speed up this reaction, one of them being Manganese IV Oxide.

EXPERIMENT:

AIM: To see if Manganese IV Oxide would speed up the decomposition of hydrogen peroxide in to water and oxygen.

APP: Spatula, two test tubes and rack.

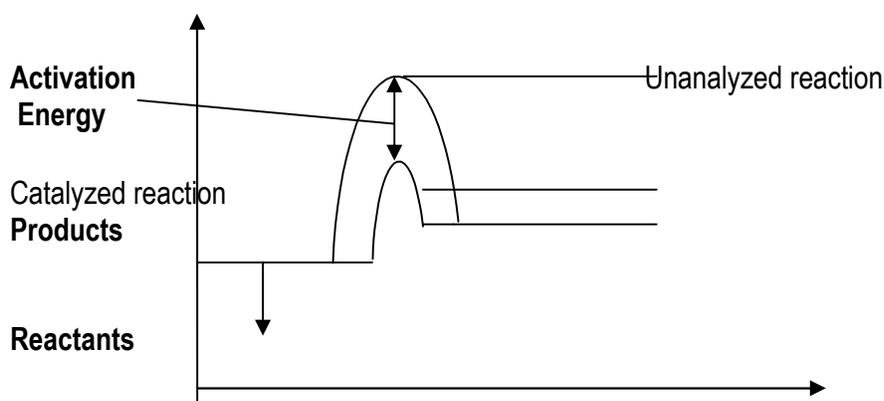
MATERIALS: Hydrogen peroxide and permanganate IV Oxide.

METHOD: Equal volumes of hydrogen peroxide are put in a test tubes and the rate of bubble formation described as either fast or slow or very slow. To one of the test tubes containing hydrogen Peroxide Manganese IV Oxide is added and the rate of bubble formation compared again.

OBSERVATION: The rate of evolution of a gas from the hydrogen peroxide where Manganese (IV) Oxide was added was very fast while where Manganese IV Oxide was not added the rate was very slow.

CONCLUSION: Catalysts generally speed up the rate of a chemical reaction.

EXPLANATIONS: A catalyst increases the rate of a chemical reaction by providing an alternative reaction path- way or a catalyst lowers the activation energy. The activation energy is the energy barrier which reactants should overcome when their particles collide to react successfully and form products. The graph below show the catalyzed and the un catalyzed reaction.



One group of naturally occurring catalysts are enzymes. Enzymes are biological catalysts that speed up hundreds of different chemical reactions taking place in our bodies. These biological catalysts are very specific. Each chemical reaction taking place in our bodies has a specific enzyme. For example hydrogen peroxide is a substance produced by the liver and it is a very poisonous substance. It is decomposed rapidly by the enzyme catalyse as shown by the equation below



Other enzymes are used in washing powder where they are added to remove stains of oils and proteins. In industry, enzymes are used to bring about reactions at normal temperature and pressure that would otherwise require very expensive conditions such as very high temperatures and very high pressures as well as very expensive equipment.

Successful processes using enzymes need to ensure that:

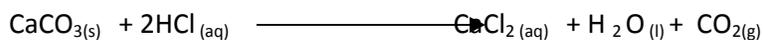
- The enzyme is able to function for a long time by optimizing the environment
- The enzyme is not lost by trapping it on the surface of an inert solid.
- Continuous process occurs rather than batch processes.

Industries employing enzymes include;

- i. Making bread, yogurt, cheese, and Chocolate
- ii. Making beer and wine
- iii. Tenderizing meat
- iv. Removing hairs and bristles from hides
- v. Biological detergents

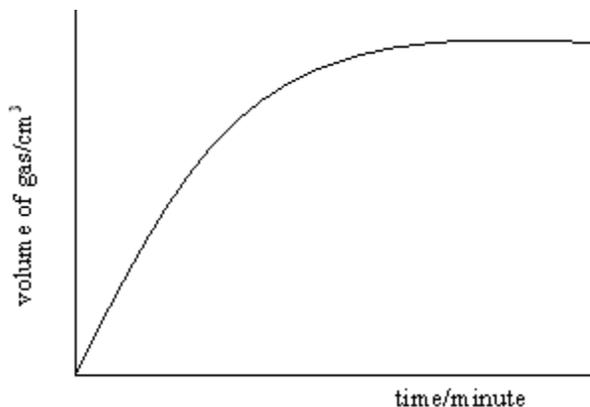
Exercises

- 1) Study the reaction between calcium carbonate and hydrochloric acid. The equation for this reaction is



The experiment was carried out to study the effect of the surface area of calcium carbonate in the form of small lumps or medium lumps or large lumps. In each experiment the same mass of calcium carbonate was used. The volume of the gas collected was measured in the gas syringe at intervals.

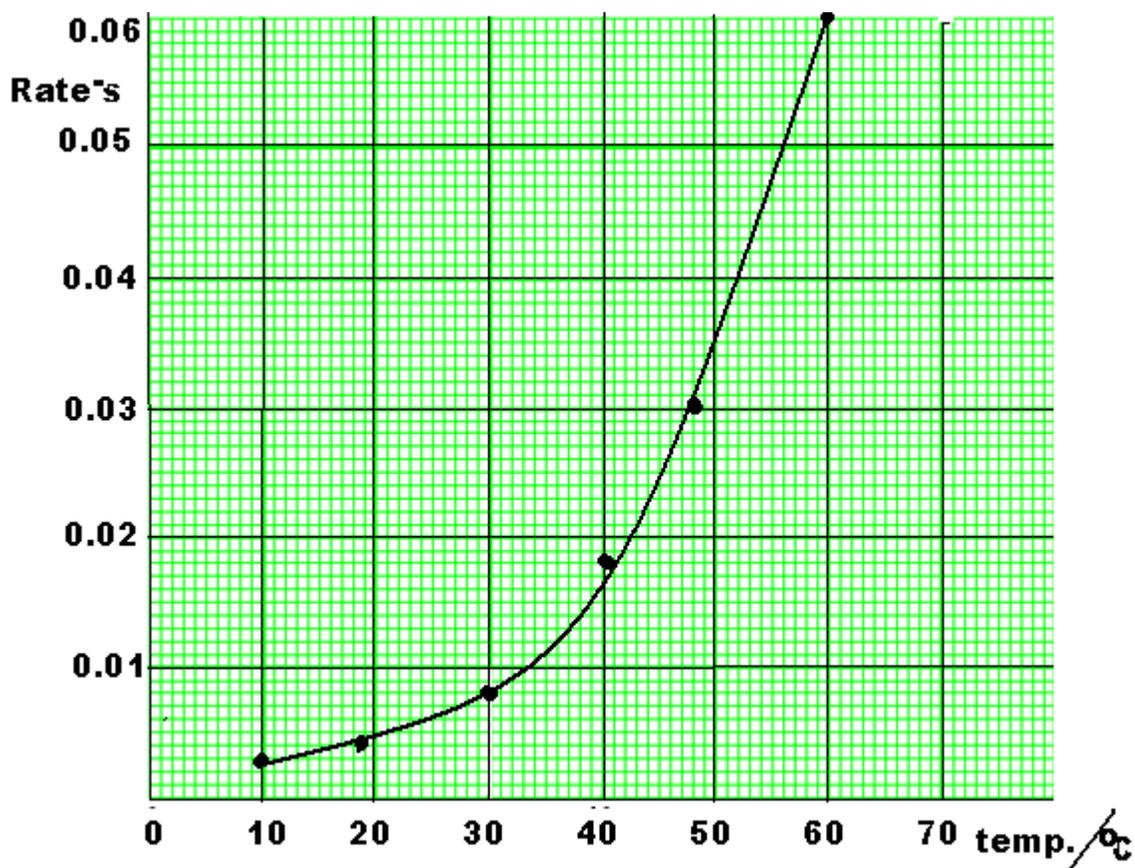
- (a) Write down two other things that should be kept constant in each experiment. [2]
- (b) In each experiment all the calcium carbonate had reacted within the five minutes. The graph for medium lumps is shown on the grid. Indicate on the same grid the graph lines for small lumps and large lumps. Label your lines S for small lumps and L for large lumps. [2]



- (c) How could you use the measurements of mass to follow the rate of reaction?

Total loss in mass/grams	0	0.32	0.60	0.83	1.04	1.13	1.19	1.20	1.20
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- (i) Plot the results on a graph and draw a smooth curve through the points.
 - (ii) Another student did the experiment again. The student used 50cm³ of Hydrochloric acid of half the concentration used by the first student. On the same graph, sketch another curve to show how the total loss in mass would change with time.
- (c) Why does the reaction eventually stop?
 - (d) State two ways the reaction could be made to go faster without using a catalyst or changing the concentration of the reagents
4. In an experiment to study the effect of temperature on the rate of a reaction, the following graph was obtained;



- (a) What is the reaction at;
 - (i) 25°C?
 - (ii) 45°C?
- (b) Calculate the temperature change (T) that Δ causes the rate of the above reaction to be doubled from 0.02 S⁻¹ to 0.03 S⁻¹
- (c) What two other factors can result in the increase of the rate of a chemical reaction?
- (d) Suggest a reason why most industrial processes are carried out at high Temperatures
- (e) State a disadvantage of high temperatures in the industrial processes.

5. (a) What do you understand by the phrase “ **rate of reaction?** [1]

(a) State the Collision Theory, in relation to rates of reactions. [1]

(b) Using the Collision Theory, explain briefly how pressure influences the speed of a reaction involving gaseous reactants. [3]

(c) What is a catalyst? [1]

(d) A student chemist investigated the effect of heating varying amounts of manganese (IV) oxide with a fixed mass (5.00g) of potassium chlorate (KClO₃). The potassium chlorate decomposed producing potassium chloride (KCl) and oxygen (O₂). The results are shown in the table below.

Mass of catalyst per 5.00g of KClO ₃	Time (t) required to produce maximum volume of gas (min)	1/t(min ⁻¹)
0.00	8	0.125
0.10	6	0.17
0.20	5	0.20
0.40	4	0.25
0.75	3	0.33
1.00	2.5	0.33

(i) Write a balanced equation for this reaction. [2]

(ii) Sketch the graph of 1/t against mass of catalyst used. [2]

(iii) Why did the student plot 1/t rather than t against the mass of catalyst used? [1]

(iv) How would the rate of evolution of oxygen alter if the manganese (IV) oxide used was granular instead of the powdered? [1]

(e) Apart from those already mentioned in the question, state one other factor that influences the rate of chemical reactions. [1]

REDOX (REDUCTION AND OXIDATION)

ReDox in terms of oxygen and hydrogen

Oxidation is the addition of oxygen to a substance or the removal of hydrogen from a substance. On the other hand reduction is the removal of oxygen from a substance or the addition of hydrogen to a substance.

Examples of Oxidation;



Copper metal has been oxidized because oxygen has been added to it.



Carbon monoxide has been oxidized to carbon dioxide because oxygen has been added to it.

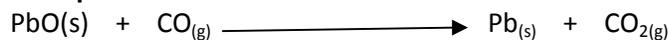


Hydrogen sulphide has been oxidized to sulphur because hydrogen has been removed from the hydrogen sulphide.



Sulphuric acid has been oxidized because hydrogen has been removed from the sulphuric acid.

Examples of Reduction



Lead oxide has been reduced to lead metal.



Iron oxide has been reduced to iron metal because it has lost the oxide.



Sulphur has been reduced because hydrogen has been added to it.

REDOX IN TERMS OF ELECTRONS

The two definitions of oxidation and reduction given above were used by earlier scientists because they considered hydrogen to have properties opposite to oxygen. However, not all reactions involve oxygen and hydrogen only. Reactions that do not involve oxygen or hydrogen are better described in terms of electron gain and electron loss. When a substance loses an electron it is said that the substance has been oxidized because it gains a higher charge. On the other hand when a substance gains an electron it is said that the substance has been reduced because it moves from a higher charge to a lower charge. Examples below illustrate oxidation and reduction in terms of electron gain and electron transfer.



Number of Protons on the left = 11

Number of electrons on the left = 11

Charge on sodium on the left = 0

Number of protons on the right = 11

Number of electrons on the right = 10

Charge on sodium on the right = +1

Bearing in mind that each electron carries a negative charge and each proton carries a positive charge. The difference gives the charge on an element or atom.

In the example of sodium above, the sodium atom has been oxidized to a sodium ion (cation) because it has lost an electron. The loss of the electron has made the sodium atom to move from an atom with a zero charge to an ion with a positive one (+) charge.



Iron has been oxidized from a cation of 2+ to a cation of 3+ because it has lost an electron.



Number of protons on chlorine atom on the left side = 17

Number of electrons on chlorine atom on the left side = 17

Charge on chlorine atom on the left side = 0

Number of protons on the chlorine atom on the right = 17

Number of electrons on the chlorine atom on the right = 18

Charge on chlorine on the right side of the arrow = 1⁻

Chlorine has been reduced from an atom with a charge of zero to an ion with a charge of negative one(1⁻)

Oxidation numbers:

- All elements in a free state i.e. uncombined with any other element have an oxidation number of zero. In other words all atoms have an oxidation number of zero.
- In the case of a simple ion; the element has an oxidation number with the same size as the charge on the ion. The charge is equal to the valence of the element. For example;

Ion	Oxidation number
Na ⁺	1 ⁺
Cu ²⁺	2 ⁺
O ²⁻	2 ⁻
Al ³⁺	3 ⁺
Cl ⁻	1 ⁻
Fe ²⁺	2 ⁺
Fe ³⁺	3 ⁺

- The oxidation number of oxygen is -2 except in peroxide
- The oxidation number of hydrogen is +1 except in metal hydroxides
- The sum total of all the oxidation numbers of the elements in a compound is zero because the number of the positive charges is equal to the number of the negative charges. The oxidation number of each element in the compound could be determined.

Examples:

- FeSO_4 ; Fe²⁺ oxidation number = 2+

S^X oxidation number = X

O₄²⁺ oxidation number = 4 x -2

The sum total of the oxidation numbers of the elements in a compound is Zero. Hence;

$$2+x+(4x-2) = 0$$

$$X = -2 + 8$$

$$X = 6$$

The oxidation number of sulphur is 6

- $\text{Cu}(\text{NO}_3)_2$; Cu²⁺ oxidation number = 2+

N oxidation number = 2x

O²⁻ oxidation number = -2x3x2

The sum total of the oxidation numbers of the elements in a compound is Zero. Hence

$$\begin{aligned} 2 + 2x + (-2 \times 3 \times 2) &= 0 \\ 2 + 2x + (-12) &= 0 \\ 2x &= 10 \\ x &= \underline{5} \end{aligned}$$

The oxidation number of Nitrogen is 5.

OXIDIZING AND REDUCING AGENTS

A reducing agent is one which causes a reduction in the other substance and is itself oxidized.

An oxidizing agent is one that causes oxidation in the other substance and is itself reduced.

For example;



In the above equation, sodium is a reducing agent because it has reduced chlorine from a chlorine molecule to chlorine ions and itself has been oxidized from a sodium atom to a sodium ion. On the other hand chlorine is an oxidizing agent because it has oxidized sodium from a sodium atom to a sodium ion i.e. a cation with a positive charge and has itself has been reduced from a chlorine molecule to chlorine ions i.e. anions with a negative charge.



Ionically;



Silver is a reducing agent because it has reduced sulphur from an atom of sulphur to an anion of sulphur while sulphur is an oxidizing agent because it has oxidized a silver atom to a silver cation.



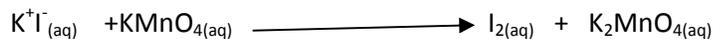
Ionically;



Chlorine is an oxidizing agent because it has oxidized iodine from iodine anions to iodine molecules with a charge of zero. Iodine has moved from a charge of negative one (-1) to a charge of zero (0). While chlorine is a reducing agent because it has reduced chlorine from chlorine molecules to chlorine anions

TEST FOR AN OXIDIZING AGENT

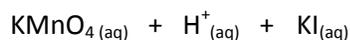
The addition of an oxidizing agent such as potassium permanganate to potassium iodide solution will cause a brown solution of iodine to be formed. The colourless potassium iodide ions are oxidized to iodine molecules. The chemical equation below illustrates the oxidation;



- Oxidizing agent manganate VII (KMnO_4) has oxidized iodide ions (I^-) to iodine molecules (I_2). The change of colour from colourless to brown is a test that potassium permanganate is an oxidizing agent.
- Manganate VII is reduced to manganate VI

TEST FOR A REDUCING AGENT

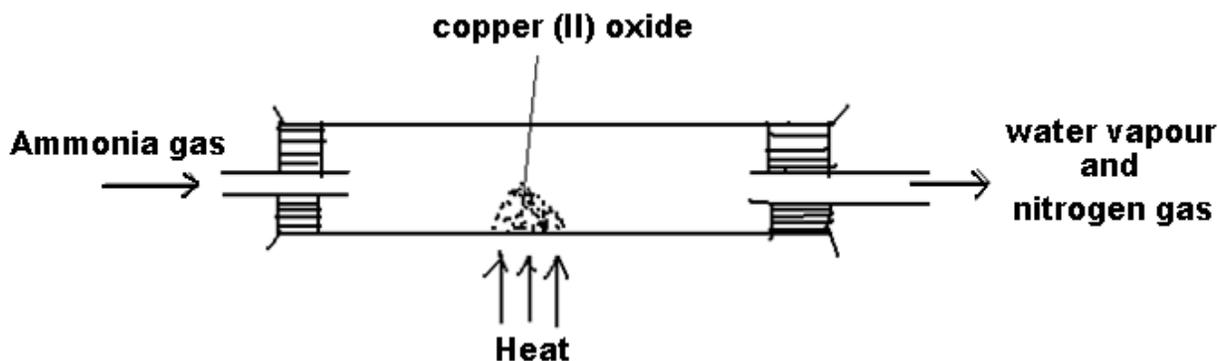
The addition of purple acidified potassium permanganate to a substance e.g. potassium iodide; the reducing agent causes the purple solution to decolourize. Also the addition of orange potassium dichromate solution to a reducing agent, the reducing agent turns the orange solution to green. The equations for the two reactions above are;



Materials that are often used as oxidizing agents include; oxygen, chlorine, dichromate and manganate. Materials that are often used as reducing agents include hydrogen, sulphur dioxide, carbon and carbon monoxide.

Exercise

1. Ammonia gas is passed over heated copper (II) oxide as shown in the diagram below:



The word equation for the reaction which occurs is:

Ammonia + copper (II) oxide \longrightarrow water vapour + nitrogen + copper

(a) Write the balanced chemical equation including state symbols for the reaction above.

(b) The above reaction is 'redox' reaction.

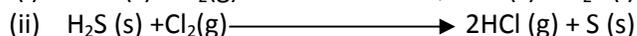
(i) What is meant by a 'redox' reaction?

(ii) State which substance is oxidized and which one is reduced. Give a reason for each answer. [4]

(iii) State how the oxidation numbers of nitrogen and copper change in this reaction. [2]

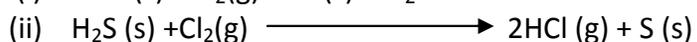
(c) What do you understand by the term reducing agent? [2]

(d) In each reaction below name the reducing agent



2. (a) What do you understand by the term oxidising agent? [1]

(b) In each reaction below name the oxidising agent

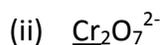
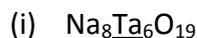


(c) Explain why the reaction below is a redox.



[2]

(d) Find the oxidation number of the underlined element in each of the following species:



(e) Describe how you would test a reducing agent.

ENERGETICS

All chemical reactions involve energy changes. Energy involved is either absorbed or released. Energy involved include; heat energy, light energy, electrical energy nuclear etc. A reaction in which energy is released is described as an exothermic reaction. Examples of exothermic reactions include combustion, neutralization, fusion, respiration etc. Reactions in which energy is absorbed from the surrounding are described as endothermic reactions. Endothermic reactions include photosynthesis, melting of ice, absorption of light in photography.

In exothermic reaction there is bond formation, where as in endothermic reaction there is bond breaking.

Experiment:

Aim: To observe change in temperature when a chemical reaction occurs.

App: 2 test tubes, spatula and a thermometer.

Material: Water, Sodium hydroxide and Ammonium chloride.

Method: Put 1cm³ of each of the compound into separate test tubes. Put 1cm³ of water to the compounds in the test tubes and insert the thermometer into the solution. Using the thermometer stir the solution while taking note of the temperature changes and also feel the test tube in the hand and observe the hotness or coldness

Observations:

When the test tubes were held in the hands the test tubes with a solution of Ammonium Chloride felt cold while the test tube with the Sodium Hydroxide felt hot. Furthermore the changes in temperature on the thermometers were recorded as shown in the table below:-

Solution	NH₄Cl	NaOH
Initial Temperature	27°C	22°C
Final Temperature	16°C	75°C
Difference	+11°C	-53°C

Conclusion: When a chemical reaction takes place energy in form of heat is either released or absorbed.

Explanation: Both the compounds NH_4Cl and NaCl are ionic solids. When the solids dissolve in water, the ions in the solid become separated from each other. The ionic bonds that hold the negative ions and the positive ions are broken. Secondly, the water molecules form new bonds with the ions. Energy is needed to break the bonds that existed between the cations and the anions; on the other hand, energy is released when new bonds are being formed.

If the energy released, when new bonds are being formed is greater than the energy needed to break the bond, reaction becomes exothermic and heat is released to surrounding. In our experiment the energy needed to break the bonds between Na^+ (Sodium ions) and the OH^- (hydroxide ions) was less than the energy that was released when Sodium ions were forming new bond with water molecules and hence the test tube in which the reaction was taking place became hot and the temperature rose from 22°C to 75°C meaning that heat was being released to the surrounding from the reaction. and the reaction was described as **Exothermic Reaction**. Other exothermic reactions include neutralization, respiration and combustion. In the reaction concerning ammonium chloride, the energy needed to break the bond between ammonium ions and chloride ions was far greater than the energy released when ammonium ions were forming new bond with water molecules and as a result the test tube felt cold, meaning that the reaction was absorbing heat from the surrounding. This type of reaction is called **Endothermic Reaction** and the temperature on the thermometer dropped from 27°C to 16°C . Sometimes the change in temperature is hardly noticeable. However, this does not mean that the reaction is thermal neutral. With a very sensitive thermometer; we could detect temperature changes.

BOND ENERGY AND ENTHALPY

When any chemical reaction occurs, chemical bonds in the reactants have to be broken down and this requires energy to be absorbed by the compound. The energy absorbed is used to breakdown the bonds. When the new bonds in the products are formed, energy is given out. The bond energy is the amount of energy in **kJ/mol** associated with the breaking or making of one mole of chemical bonds. In a molecular element or compound. Bond energy also tells us how much energy is needed to break a chemical a chemical bond and how much energy is given out when making a chemical bond.

The table below shows bond energies in kilojoules per mole.

Bond	Bond Energy (kJ/mol)
H - H	436
H - Cl	431
C = C	612
N \equiv N	409
Cl - Cl	242
C - H	412
N - H	163
O = O	497
C = O	803
H - O	464
C - O	358

C - C	348
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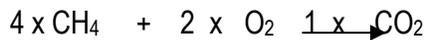
Exercise:

From the table above calculate energy change (enthalpy change) when 1 mole of methane is completely burned in air to give carbon dioxide, water and energy, according to the equation.



Bond breaking

Bond formation.



$$2642\text{kg} \longrightarrow 3462\text{kg}$$

Energy in \longrightarrow Energy out

Enthalpy = Energy need to break - energy given out when making bond.

$$2642 - 3462$$

$$\Delta = -820\text{kJ}$$

Exercise,

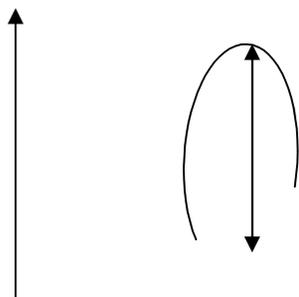
1. Calculate the enthalpy change for the complete combustion of carbon in air to yield carbon dioxide. According to the equation

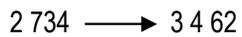
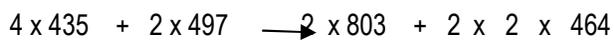
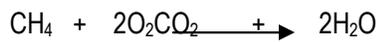
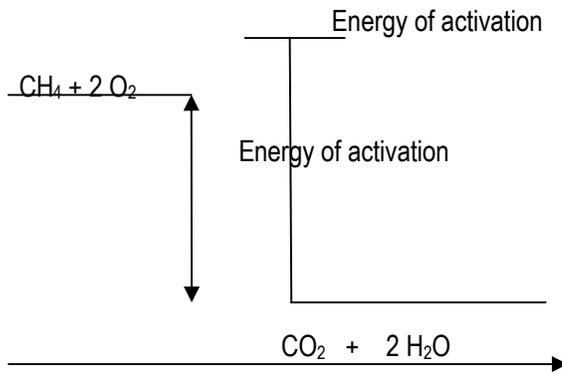


Energy level diagram:

The energy changes that take place during a chemical reaction can be shown by using an energy level diagram. The diagram shows the energy content of the reactants and products. Below are two energy level diagram illustrating an endothermic and exothermic reaction.

1. C = H 435kJ
O = O 497kJ
C = O 803kJ
H = O 464kJ



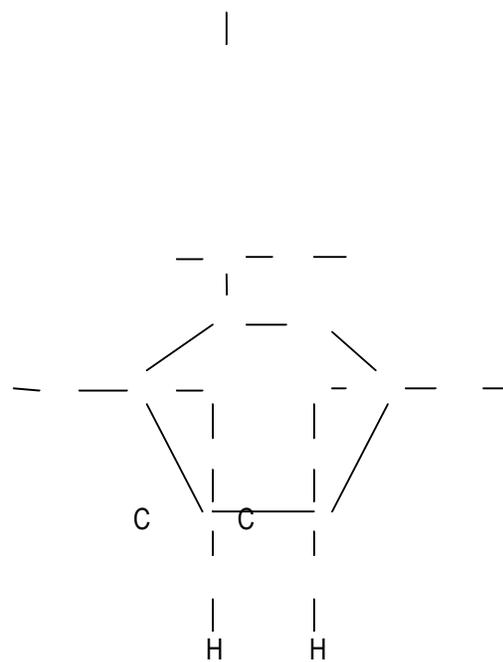
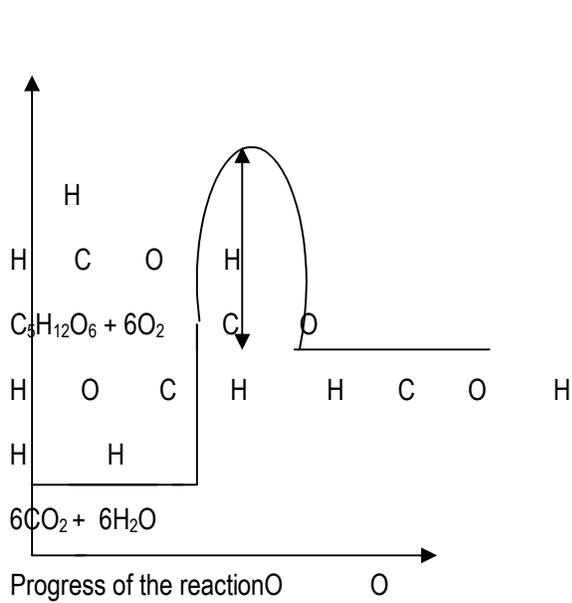


Energy produced - Energy given out

$$2734 - 3462$$

$$= -728\text{kJ}$$

The negative sign shows that the chemical reaction was losing heat to the surrounding. That is, the reaction was exothermic.



Structure of Glucose

A positive sign represents an endothermic reaction this means that the amount of energy need to break the bonds was more than the amount of energy released when new bonds were being formed in the products. In the above energy level diagram, the raw materials CO₂ and H₂O require more energy to form high energy content sugars.

From the energy level diagrams heat change can also be calculated by subtracting the value for reactants from the value of the products.

ENERGY SOURCES

An energy source is a place or a place or a plant designed to produce energy for use by people in industries and homes. Energy sources include water, wind, soil, nuclear, thermal (fuel), Geothermal and tidal wave.

ENERGY FROM WATER.

It is usually called hydroelectricity. This form of energy is widely used in countries where there are rivers and gorges.

Advantages of H.E.P

- No pollution
- It is cheap to run
- It is clean
- It is safe to operate

Disadvantages of H.E.P

- Expensive to construct
- Not all rivers have gorges
- Amount of water in the dam may affect the out put

Nuclearenergy

Nuclear energy production depends on the fusion of Uranium 235. Nuclear fusion occurs when the unstable nucleus of radio-active Isotopes splits up forming smaller particles and in the process producing a large amount of heat energy. The heat energy is used to heat water in boiler to produce steam. The steam is directed into the turbines which rotate and drive the dynamo to produce electricity.

Advantages of nuclear energy

- It can be set up anywhere or close to industries.

Disadvantages:

- The most expensive form of energy plant to construct because it needs careful setting, more elaborate equipment and structure.
- Nuclear energy plants are very expensive to ensure safety of workers and the environment.
- Methods of depositing waste products are very expensive.
- Radioactive leakages pollute the environment.

- Radioactive leakages cause mutations in living organism.

Solarenergy:

This is the energy from the sun. The harnessing of solar energy involves the use of photo voltaic cells. The photo voltaic cells convert light and heat energy from the sun into electrical energy.

Advantages

- Non pollutant
- Can be set up anywhere
- Cheap to run

Disadvantages:

- Affected by weather
- The initial cost of all is very high.

Geothermal:

Water is pumped into hot rocks in the earth crust. The enternal heat of the rocks convert the water to steam. The steam is used to drive turbines and hence generate electricity.

Advantages

- Non pollutant
- It is cheap to install as it is natural.
- Its clean and cheap to run

Disadvantages:

Very few places in the world have sites for this.

Thermalenergy:

Under this form of energy fuels such as coal, diesel, methane etc are burned in power station to heat water to produce steam. The steam is used to drive the turbines.

Advantages

- It can be set anywhere.
- It is cheap to install

Disadvantages:

- Causes pollution as fossil fuel production Carbon monoxide and sulphur dioxide
- Very dirty

WIND:

This method uses the force of the wind to rotate the turbines to produce electricity. Wind machines 24m high are a capable of generating 200kw.

Advantages

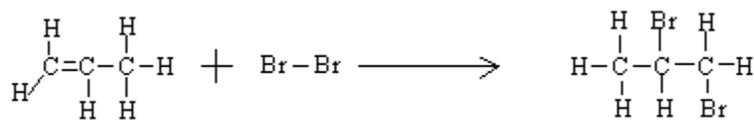
- Non pollutant
- It is cheap to install as it is natural.
- Its clean and cheap to run

Disadvantages:

Affected by weather i.e. no wind no power

Exercises

1) The reaction between propene and bromine is given below.



Given the bond energies in the table below as

Bond	Energy /KJ
C - C	347
C - H	435
Br - Br	193
C - Br	293
C - C	598

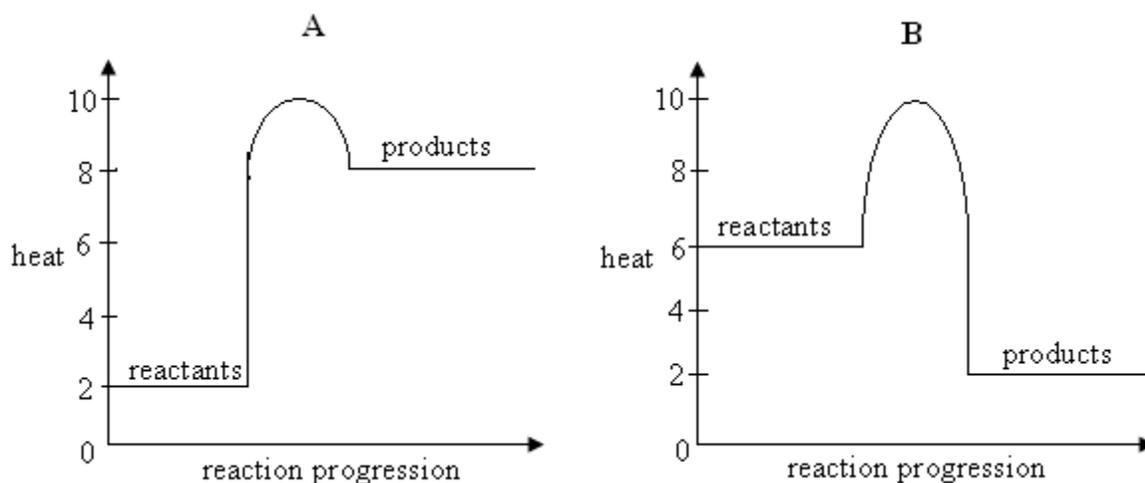
(a) Determine the enthalpy(heat) change of a reaction [3]

(b) What type of a reaction is this? Give a reason for your choice [2]

(c) Give an everyday example of this reaction.[1]

(d) Draw an energy profile diagram to on this reaction. On your diagram include reactants, products and activation energy. [2]

2) Study the energy level diagrams of the two reactions represented by A and B.



a) Which of the energy level diagram above represents?

i) an endothermic reaction? [1]

ii) an exothermic reaction? [1]

b) Give an everyday example of a reaction represented by energy level diagram A.[1]

c) Define activation energy. [1]

d) What is the value of activation energy represented by energy level diagram;

i) A.[1]

ii) B.[1]

e) Calculate the energy change of a reaction represented by energy level diagram B.[2]

3) Define;

a) an exothermic reaction. [1]

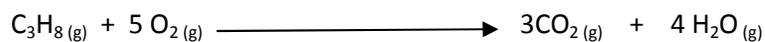
b) an endothermic reaction [1]

c) Give an example of an exothermic reaction and endothermic reaction.

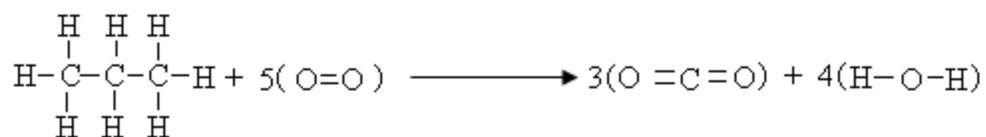
Exothermic [1]

endothermic [1]

4) 'Propagas' is used in some central heating systems where natural gas is not available. It burns according to the following equation.



The structural chemical equation is shown below.



The table below shows the bond energy data for the series of covalent bonds.

Bond	Bond energy/kJmol ⁻¹
C-H	435
O=O	497
C=O	803
H-O	464
C-C	347
C-O	358

a) Using the bond energies above, calculate the enthalpy change of a reaction [4]

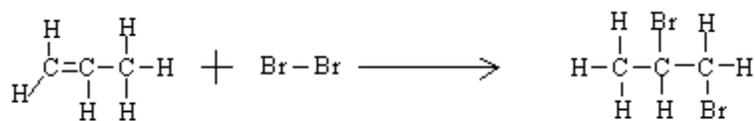
b) Is this reaction exothermic or endothermic? Give a reason for your choice

Choice [1]

Reason [1]

c) Draw an energy level diagram of the reaction above.

5) The reaction between propene and bromine is given below.



Given the bond energies in the table below as

Bond	Energy /KJ
C-C	347
C-H	435
Br-Br	193
C-Br	293
C=C	598

(a) Determine the enthalpy(heat) change of a reaction [3]

(b) What type of a reaction is this? Give a reason for your choice [2]

(c) Give an everyday example of this reaction.

(d) Draw an energy profile diagram to on this reaction. On your diagram include reactants, products and activation

(d) On the same graph, sketch another curve to.

e) Explain the danger associated with fine particles of flour in milling industry

Grade 11 term 3

METALS AND THE ELECTROCHEMICAL SERIES

Metal occupy the greater part of the periodic table making a stair case from aluminium to polonium.



Metals have the following physical properties and chemical properties:

Physical properties

- (i) Metals have high melting and boiling points because of strong forces of attraction between the positive metal ions and the mobile electrons (metallic bonding)
- (ii) They are good conductor of heat and electricity due to the mobile electrons.
- (iii) Metals are malleable i.e. can be hammered into different sheets because the ions can be forced into a particular shape of a sheet due to spaces that exist in the metal lattice.
- (iv) Metals are ductile i.e. they can be drawn into thin wires because rows of ions can slide over each.
- (v) Metals have high densities because the volume occupied by electrons is relatively smaller while their mass numbers are big.
- (vi) They are shiny when polished because when they absorb heat their electrons get excited and move from the ground state to a higher state, and when falling back they release the energy they had absorbed.
- (vii) Metals are sonorous i.e. they make noise when struck. This is because their ions vibrate more when struck.
- (viii) Metals can bend because metallic ions can be shifted into particular shape and direction without snapping (breaking)

Chemical Properties:

- (i) They have less than 4 electrons in their valence shells
- (ii) Metals form positive ions
- (iii) Many metals react with oxygen to form metallic oxides.
- (iv) Many metals react with acids to liberate hydrogen gas and form a metallic salt.

METALS' REACTIVITY SERIES

Metals react with different substances to form compounds. However, not all metals react in the same way when subjected to a particular stimulus or reagent. Some metals react with cold water; for example sodium metal reacts with cold water to form hydrogen gas and sodium hydroxide solution. Some metals react with hot water e.g. Magnesium reacts with hot water to form magnesium hydroxide and hydrogen gas. Other metals will react with steam e.g. iron metal only reacts when it is heated in steam. Other reagents that metals react with include oxygen and acids. The following three tables show the reactivity of metals with water, oxygen and dilute hydrochloric acid

Reactions between metal and water

Metal	Reaction with water	Products
Potassium	Reacts very violently with cold water and catches fire.	$\text{KOH}_{(aq)} + \text{H}_{2(g)}$
Sodium	Reacts very vigorously with cold water producing a hissing sound.	$\text{NaOH}_{(aq)} + \text{H}_{2(g)}$
Calcium	Readily reacts with cold water and sinks to the bottom.	$\text{Ca(OH)}_2 + \text{H}_2$
Magnesium	Reacts slowly with cold water, but fast with hot water and steam.	$\text{Mg(OH)}_2 + \text{H}_2$
Zinc	Reacts with steam slowly.	$\text{ZnO} + \text{H}_2$
Iron	Only reacts when heated in steam	$\text{Fe}_3\text{O}_4 + \text{H}_2$
Copper	No reaction	-
Gold	No reaction	-

Reactions between Metals with Oxygen when heated.

Metal	When heated oxygen	Products
Sodium	Catches fire even with little heating.	Na_2O_2
Magnesium	Catches fire and burns with a white flame forming a white powder.	MgO
Iron	Hot metal glows in Oxygen forming a black substance.	Fe_3O_2
Copper	Hot metals become coated with a black substance.	CuO
Gold	No reaction no matter how much heat is applied.	

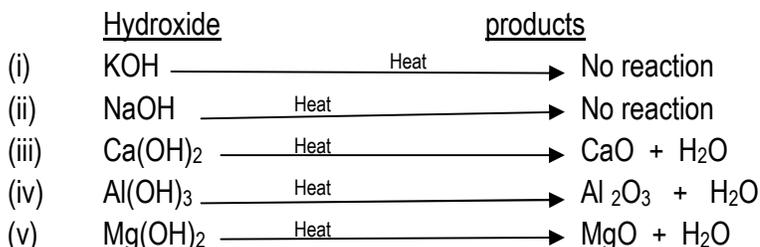
Reactions between metals with acids e.g. HCl

K, Na, Ca	Very violent and not advisable to do experiments with these metals.	H_2 and a metal salt of the acid.
Mg	Vigorous reactions	As above
Zn	The reaction is slow	As above
Fe	The reaction is very slow	As above
Lead	The reaction is very, very slow and only with the concentrated acid.	As above

stable compounds while metals lower down in the series form less stable compounds. In addition, the radicals that form compounds with metals may also determine their stability.

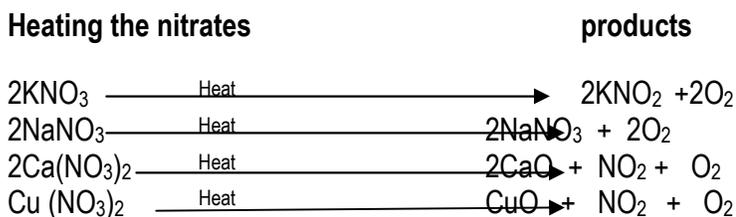
(a) Thermal decomposition of hydroxides.

The equations below summaries the thermal decomposition of metal hydroxides:

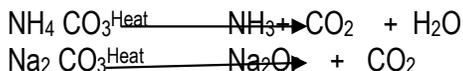


Thermal stability of nitrates;

The table below summaries thermal stability of nitrates:



Carbonates of most metals are less stable hence they easily decompose on heating to give the metal oxide, carbon dioxide and H₂O in some instances.



Carbonates are less stable than nitrates, nitrates less stable than hydroxides

EXTRACTION OF METALS

There two main methods of extracting metals from their ores namely:

- (i) Reduction of the metal from its compound by using Carbon



- (ii) Electrolysis of the molten or fused compound.

The method of extraction of a metal from its compound is determined by its position in the metals' reactivity series. The more reactive the metal is the harder it is to extract the metal from ionic compounds. This is because, very reactive metals form ionic compounds with strong ionic bonds. Very reactive metals such as potassium, sodium and aluminium are extracted by using electrolysis. Less reactive metals such as Zinc, iron and Copper are extracted by using reduction method mainly using Carbon. Metals towards the end of the series are very unreactive. Metals such as silver and Gold occur in nature in native form or as free metal elements.

The chart below shows relative positions of metals in the series and the method of extraction.

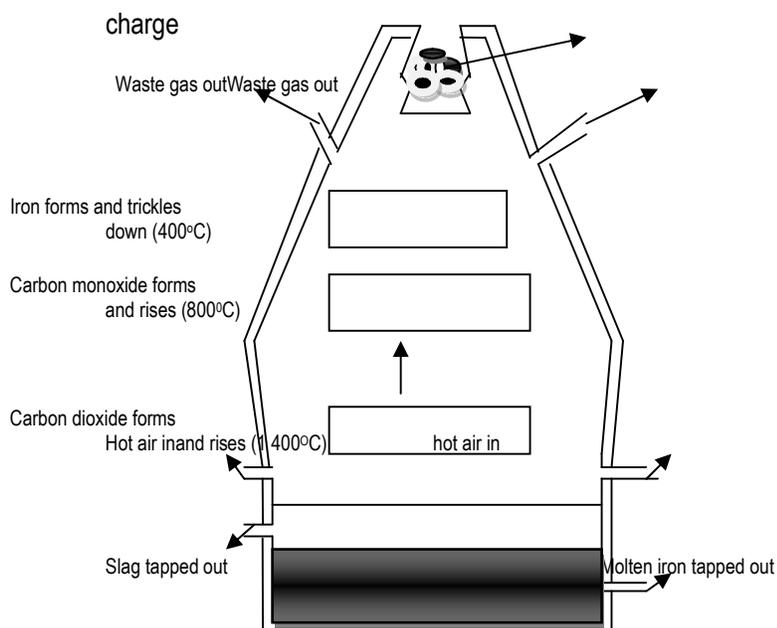
K	}	Extracted by electrolysis of their molten (fused) compound.
Na		
Ca		
Mg		
Al		
Zn	}	Extracted by reduction of the metal oxide using carbon
Fe		
Pb		
Cu		
Ag	}	may occur as free metal in nature
Pt		
Au		

Electrolysis involves the use of large quantities of electricity and it is a very expensive process. It is only used to extract very reactive metals because their compounds are too stable to be reduced by carbon. On the other hand carbon is used to extract a less reactive metal from its metal oxide and a redox reaction takes place. Carbon is said to be a reducing agent because it reduces the metal oxide to the metal by removing oxygen or a sulphide from it.

EXTRACTION OF IRON

Iron is extracted mainly from its oxides; Haematite (Fe_2O_3) and Magnetite (Fe_3O_4). The other raw materials are Coke (Carbon) and limestone (Calcium Carbonate)

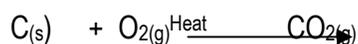
The extraction takes place in a blast furnace. The blast furnace is a steel tower of about 50m high



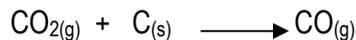
The three raw materials; haematite, coke and limestone are mixed to form a mixture called a charge. The charge is fed into the blast furnace from the top by a hopper.

Hot air is blasted or forced through the bottom of the blast furnace. This hot air makes the charge glow and become white-hot. Several reactions take place in the blast furnace and finally liquid iron is produced and trickles to the bottom. The essential reactions are outlined below:-

- (i) The Coke reacts with oxygen in the hot air to produce CO₂. This reaction is an exothermic reaction.



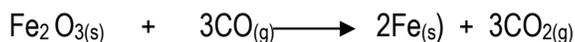
- (ii) The Carbon dioxide reacts with more coke as it moves up wards to produce Carbon monoxide. This is an endothermic reaction as seen from the drop in temperature from 1400°C to 800°C.



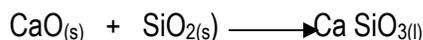
- (iii) The limestone decomposes due to heat to give carbon dioxide and Calcium Oxide.



- (iv) The Carbon Monoxide produced reacts with the Fe₂O₃ (iron III Oxide) to produce a liquid iron which trickles to the bottom of the blast furnace:



- (v) The Calcium Oxide formed when the limestone decomposes reacts with silica in the sand of the iron ore to form a slag called Calcium Silicate:-



The limestone has two functions:

- To form Calcium Oxide, which is usually referred to as flux.
- To form Carbon dioxide which is later reduced to Carbon monoxide, hence increasing the amount of Carbon monoxide in the blast furnace.

The slag runs down the furnace and floats on the iron. It is drained from the bottom of the furnace and it solidifies. The slag is used to build roads.

The waste gases mainly nitrogen and oxide of carbon are very hot and are used to heat the incoming air and helps to reduce the energy cost of the extraction process.

The iron from the blast furnace is called 'Pig' iron or cast iron. Cast iron is hard and brittle hence it has limited use. Much of the pig iron is further purified and mixed with other elements to form

alloys such as stainless steel, mild steel and hard steel. Mild steel is used to make car-bodies while hard steel is used to make blades of cutting tools.

Extraction of Aluminium

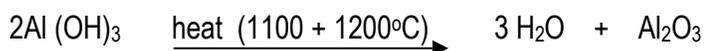
Aluminium is the most abundant metal in the earths' crust. Its main ore is bauxite. Bauxite is aluminium oxide mixed with sand and some iron oxide. Aluminium is extracted by using the electrolysis method. The bauxite is first treated with Sodium hydroxide in the digester to obtain pure aluminium oxide. In the digester impurities like iron (III) oxide and sand are removed by filtering. The equation of the reaction is:-



The aluminium hydroxide is later precipitated to get rid of the sodium hydroxide and obtain aluminium hydroxide according to the equation

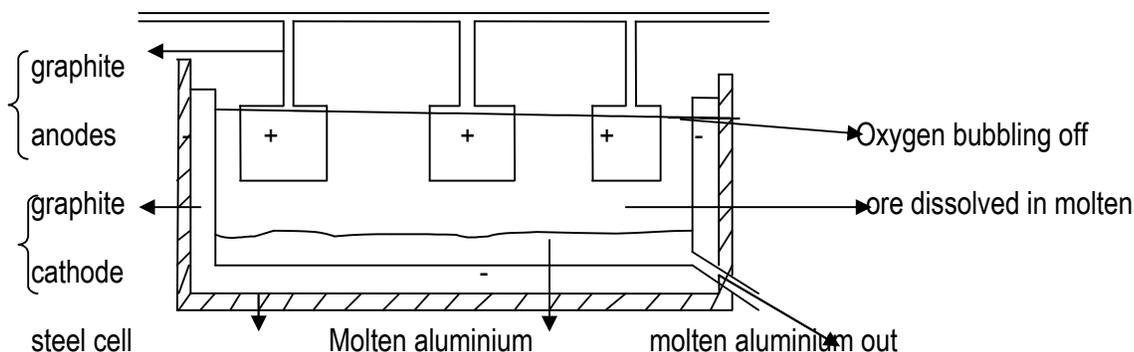


The aluminium hydroxide formed is heated to a very high temperature to produce aluminium hydroxide and water according to the equation;



The pure aluminium oxide is now dissolved in a solution of molten cryolite (sodium fluoro aluminium) cryolite is used to reduce the melting point of pure aluminium oxide from 2017°C to 800°C

The solution is now put in an electrolytic cell. The cell is shown in the diagram below;



In the electrolytic cell oxygen ions migrate to the positive Anode where they lose the electrons and become oxygen atoms according to the equation.



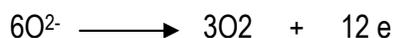
Two oxygen atoms combine to form one oxygen molecule and the oxygen gas bubbles off at the Anode. At the Cathode; Aluminium ions migrate to the cathode where they gain electrons to become aluminium atoms (Aluminium metal). The aluminium ions are therefore, reduced to aluminium metal. The aluminium atoms

collect together and drop to the bottom of the cell as molten aluminium metal. This is run off at intervals by opening the tap at the bottom.

The overall chemical equation is as shown below.



The actual number of electrons gained and lost is summarized as shown below.



The Oxygen liberated at the anode reacts with the graphite (carbon) to produce CO_2 , this reaction burns away the anode and as a result the graphite anode has to be replaced regularly.

Properties of Aluminium

- (i) It is a silvery shiny metal
- (ii) It has a very low density (light)
- (iii) It is a good conductor of heat and electricity.
- (iv) It is ductile and malleable
- (v) It is not toxic
- (vi) It resists corrosion from different solvents (acids and bases)

Uses of Aluminium

- (i) Used to make sauce pans because it is a good conductor of heat and doesn't react with cooking solvents.
- (ii) It is used to wrap food stuffs because aluminium foil is non-toxic and malleable.
- (iii) It is used to make aeroplane parts because it is light and durable and has a high melting point.
- (iv) It is used to make alloys such as duraluminium which is used to make rockets parts.

Extraction of Copper

The main ores of copper are:-

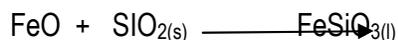
- (a) Copper pyrites (CuFeS_2)
- (b) Cuprites (Cu_2S)
- (c) Copper I Sulphide (Cu_2S)

Copper pyrite is mostly used to extract Copper. This ore contains a mixture of Copper II Sulphite and Iron II Sulphide.

After mining the Copper ore is crushed and put in a froth floatation tank to concentrate the ore. The concentrated Copper ore is roasted in controlled air to produce copper sulphide, according to the equation.



The Iron (II) Oxide is removed by adding silicon dioxide and heating. The reaction between silicon dioxide and Iron (II) Oxide forms Iron (II) Silicate.

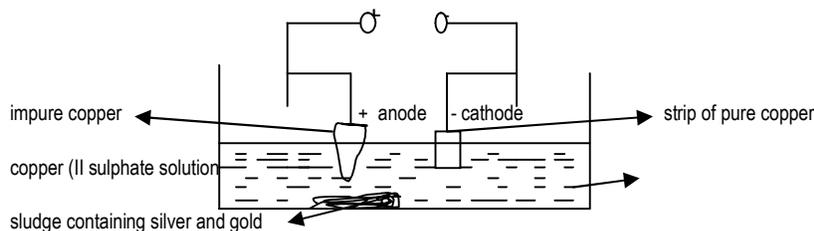


The silicate is poured away. The Copper sulphide is reduced to the metal by heating it in a controlled supply of air.



Purification of copper;

The Copper produced at this stage is too impure to be used in industry for many purposes. To make it pure it is refined in the electrolytic cell. In the electrolytic cell the impure Copper is made into the anode and stripes of pure Copper are made into the Cathode. The electrolyte is Copper (II) Sulphate solution. The diagram below shows the purification of copper.

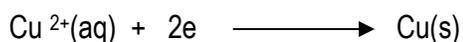


During electrolysis, Copper is transferred from the impure Copper anode to the pure Copper Cathode and a layer of copper builds up on the Cathode. While the anode dissolves away.

The anode dissolves because its atoms give up electrons to form Copper ions according to the equation.



The Cathode becomes big because the Copper ions that have come from the impure anode grab the electrons and become Copper atoms. These atoms stick onto the Cathode hence increasing in size



The mass gained by the cathode is equal to the mass lost by the anode.

Properties of Copper

- (i) It is a reddish brown metal
- (iii) It has a very high melting point and boiling point.
- (iv) It has a very high density
- (v) It is a good conductor of heat and electricity
- (vi) It is ductile and malleable

- (vii) It is not easily attacked by air and water.

Uses of Copper:

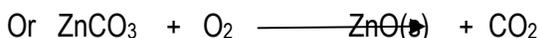
- (i) It is used in electrical connecting because it is a good conductor of electricity.
- (ii) It is used to make alloys e.g. brass which is an alloy of Copper and Zinc.
- (iii) It is used to make Copper Coins
- (iv) It is used in making ornaments e.g. Copper plates.
- (v) It is used to make car radiators.

Extraction of Zinc

The main ores zinc are:

- (a) Zinc Carbonate also known as Calamine (ZnCO_3)
- (b) Zinc Sulphide which is also known as zinc blende. (ZnS)

The ores of zinc are crushed and roasted in air and zinc oxide is formed according to the equations below:-



Zinc oxide is mixed with coke and heated in a fire clay. Zinc oxide is reduced to Zinc metal according to the equation below:



Further purification is done by electrolysis.

Uses of Zinc

- (i) It is used to galvanize iron objects zinc resists rust.
- (ii) It is used for making alloy called brass.

EFFECTS OF MINING.

- (1) Huge amount of rock are dug up and this leaves huge scars on the landscape in case of open pits and huge holes underground in case of underground mining.
- (2) Huge holes underground causes land to sink and the result is severe damage to buildings.
- (3) Unwanted rock materials heap on the surface of the land; this cause land pollution.
- (4) During smelting, poisonous fumes such as sulphur dioxide are emitted into the atmosphere causing air pollution.

- (5) Too much dust may result from mining and this may cause air pollution. Further more the dust may close stomata on leaves leading to death of plants.
- (6) Poisonous compounds like lead, Cadmium and Arsenic brought to the surface together with the ore of the mineral, get washed into the soil and rivers. These kill fish and plant life and can end up in air and food through food chain.

METAL RECYCLING

Minerals are diminishing resources hence it is essential to recycle such minerals as aluminium and iron.

Advantages of Recycling Metals:

- (i) It is a way of conserving minerals so that they last longer.
- (ii) With recycling there will be less mining, less air and water pollution due to mining processes.
- (iii) Demand for metal ores will decrease once scrap metal is identified as a viable source of raw material.
- (iv) Less disposal sites will be needed to dispose scrap metal, this solves the problem of litter accumulation.
- (v) Recycling of metals such as aluminium uses less energy than the full process of extraction of aluminium from bauxite.

Constraints to Recycling:

- (i) Recycling is sometimes not possible because of the cost involved in terms of:-
 - (a) Transporting materials to be recycled to recycling plants.
 - (b) Sorting through the waste to get metal can be time wasting.
 - (c) Cleaning the scrap metal before the recycling process could begin can be expensive.
 - (d) Recycling is at times not possible because most metals are made in form of alloys. Hence difficult to separate the mixture into its components.

Alloys:

Pure metals are soft and hence not suitable for heavy duty engineering works such as construction of bridges and buildings. Pure metals are soft because of the spaces that exist between the atoms, these spaces, make metal atoms slide over each other. To improve the strength of the pure metal, a carefully calculated amount of atoms of another element are added to the atoms of the pure metal. These atoms act as wedges between atoms of the pure metal and hold the atoms in their fixed position. In this state, the metal acquires a higher mechanical and tensile strength suitable for heavy duty engineering.

An alloy therefore, is a mixture of a metal and another metal or a metal and a non-metal. The composition of an alloy depends on the function that the alloy will perform and this has become an area of intense research in the world today and huge amounts of money is spent on designing alloys for various engineering purposes. Usually, the compositions of elements in the alloy are expressed in percentages.

The table shows some of the alloys;

Alloy	Composition	Property	Use
Mild steel	99.5% Iron 0.5%	- Not brittle - Easily worked	- Making Car bodies
Stainless Steel	74% Iron 18% Chromium 8% Nickel	- Very tough - Doesn't corrode	- Making cutlery - Build chemical plants - Making surgical instruments
Brass	65% Copper 35% Zinc	- Harder than Copper - Doesn't Corrode	- Musical instruments - Electrical connection
Cupronickel	75% Copper 25% Nickel	- Attractive - Silver colour	- Silver coins
Duraluminium	95% Aluminium 4% Copper 1% Manganese	- High Melting Point - Light - Does not Corrode - Durable	- Making Aircraft

Rust:

Rust is a loose, reddish brown layer of hydrated Iron (III) Oxide found on the surfaces of Iron and steel objects. The rusting of Iron or steel is a serious problem and wastes huge amounts of money in terms of replacing the rusted iron objects. Both water and oxygen are essential for iron to rust and if one of these two substances is not present, then rusting does not take place.



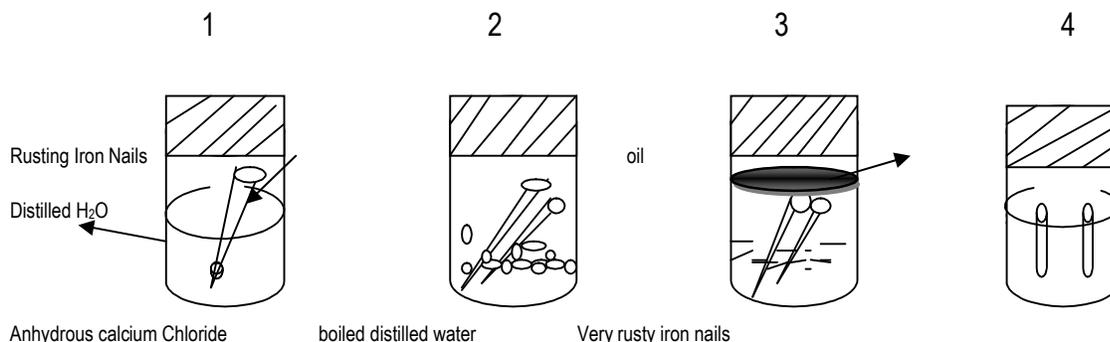
Experiment

Aim: To investigate conditions necessary for rusting to occur.

App: 4 test tubes

Materials: water, Calcium chloride, nail, oil.

Method: The experiment is set up as shown below it is then left for several days.



Observation and explanation:

- In test tube 1: The iron nails will rust because oxygen in air will dissolve in distilled water; hence both oxygen and moisture are available which are necessary conditions for rusting to take place.
- In test tube 2: The iron nails will not rust because the anhydrous calcium chloride absorbs moisture from the air hence one condition necessary for rusting will be absent.
- In test tube 3: Rusting will not occur because oxygen was driven out from water through boiling and oxygen air will not dissolve in distilled water because it is blocked by a layer of oil.
- In test tube 4: The iron nails rusted because oxygen and water were present. There is more rusting in test tube 4 than in test tube 3 because of high concentration of oxygen.

Ways of preventing rust;

To prevent iron from rusting, it is necessary to stop oxygen and water from coming into contact with the iron or steel object. There are several ways of doing this:

(a) Painting:-

Painting is coating the iron or steel object with a layer of paint to stop oxygen and water from getting in touch with the iron metal. However, if the paint is scratched the iron could rust beneath and corrosion could spread to other parts.

(b) Oiling or greasing:

The iron or steel in the moving parts of machinery are coated with oil to prevent rust this is because oil and water are immiscible.

(c) Galvanizing:

Galvanizing is the coating of a steel or iron object with a layer of Zinc. Objects like iron dust bins, wires used to fence areas are dipped in molten zinc. The thin layer of zinc on the surface of the iron object will prevent air and moisture from getting into contact with the iron object. Zinc is more reactive than iron and by coating iron with zinc rusting of iron is prevented as zinc oxide or hydroxide are insoluble in water hence cutting out the water.

(d) Plastic:

The exteriors of iron or steel object are insulated with plastics such as PVC (Polyvinyl Chloride) to prevent the iron object getting into contact with air and water

(e) Sacrificial Protection:

This is coating the iron or steel object with a more reactive metal. Water and air will be reacting with a more reactive metal on the surface leaving the iron object unreacted. When the more reactive metal finishes it has to be renewed.

(f) Electro-plating:

Cans for food can be made from steel coated with tin. The tin is deposited on the steel used to make food by dipping the steel in molten tin. Tin doesn't corrode because it resists air attack hence the iron metal under the tin is prevented from rusting.

ELECTRO-CHEMISTRY

Electrolysis is the process of splitting up or decomposing a substance by passing an electric current through it. The substance which is being decomposed is called the electrolyte. An electrolyte is a substance that conducts electricity in a molten state or in solution. The electricity is carried through the electrolyte by ions. The electric current enters and leaves the electrolyte through the electrodes. Electrodes are usually made of unreactive (inert) metals such as platinum or made of a nonmetal like carbon.

The names given to the two electrodes are the cathode which is a negative electrode and the anode which is a positive electrode. The cathode attracts cations (positively charged ions). The anode attracts anions (negatively charged ions).

Electricity is the flow of charge i.e. either cations or anions or electrons. Cations in solution or molten state move to the cathode. Electrons move from the anode to the cathode through the external circuit. The diagram below summarizes the movement of cations, anions and electrons during electrolysis.

When a charge flows, an electric current is produced and an electric current is measured in units called Amperes while charge is measured in units called Coulombs. When a current of 1 ampere flows for 1 second, a charge of 1 coulomb is produced. This relationship is summarized in the formula.

$$\text{Current (A)} \times \text{Time (s)} = \text{Charge in coulombs.}$$

$$A \times t(s) = C$$

- (1) Find the size of the current flowing through a wire when 0.8C of charge passes a point in the circuit in 2 seconds.

$$Q = I \times t$$

$$0.8 = I \times 2$$

$$I = 0.4A$$

- (2) A current of 20A flows in a conductor for 10 seconds. What is the size of the charge?

$$Q = It$$

$$Q = 20 \times 10$$

$$Q = 200C$$

Sr. Michael Faraday found out that one mole of electrons carries a charge of 96 500C. From this the charge of one electron can be calculated. As follows

$$1 \text{ mole} = 6.02 \times 10^{23} \text{ electrons.}$$

$$6.02 \times 10^{23} \text{ electrons} = 96 500C$$

1 electron = x

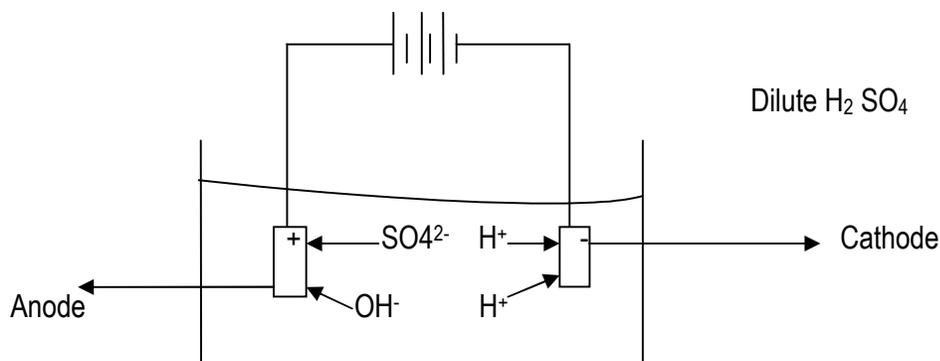
$$6.02 \times 10^{23} \times x = 96\,500^\circ\text{C}$$

$$x = \frac{96500}{6.02 \times 10^{23}}$$

$$x = 1.6 \times 10^{-19} \text{ coulombs}$$

(a) Electrolysis of dilute sulphuric acid:

The diagram below shows the electrolysis cell for the electrolysis of dilute sulphuric acid.

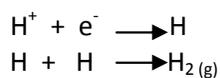


The ions present in the electrolysis of sulphuric acid are:

- (a) OH^- and H^+ from water
- (b) SO_4^{2-} and H^+ from sulphuric acid.

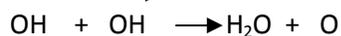
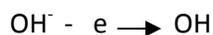
Hydrogen ions migrate to the cathode where they gain electrons and become hydrogen atoms. The hydrogen atoms combine in pairs to form hydrogen molecules. The hydrogen molecules are given off as hydrogen gas. The half cell reaction equation is as follows:-

At the cathode;

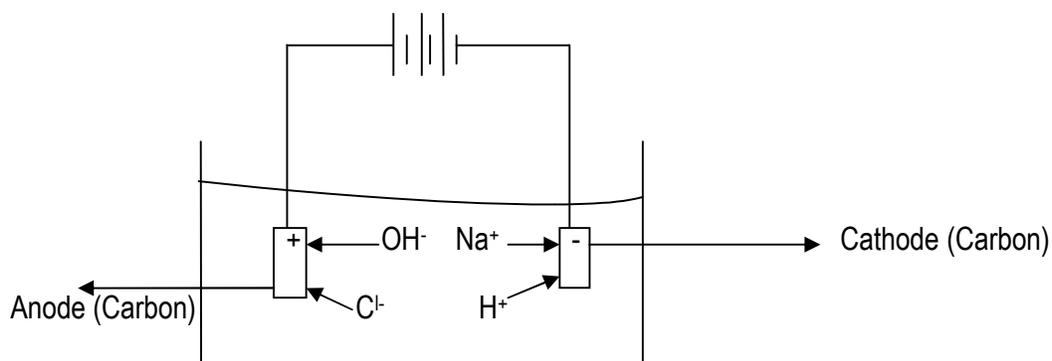


At the Anode:

The sulphate anion and the hydroxide anion migrate to the anode. At the anode OH^- loses electrons and is discharged in preference to the sulphate anion. OH^- is discharged because it is less reactive than SO_4^{2-} . The OH groups chemically combine in pairs to form H_2O and oxygen atoms. The oxygen atoms combine to form oxygen molecules which are given off as oxygen gas. The half cell reaction taking place at the anode is as shown below:-

**.Identity test of gases:**

The gas given off at the cathode burns with a pop sound. The gas given off at the anode relights a glowing splint.

(b)Electrolysis of concentrated sodium Chloride solution (brine):

The solution contains Na^+ and Cl^- from Sodium chloride. It also contains H^+ and OH^- from water. The positive ions go to the cathode while the negative ions go to the Anode.

At the Cathode:

The hydrogen ions will accept electrons and become hydrogen atoms. Hydrogen atoms will combine in pairs to form hydrogen gas which is given out and is lost to the atmosphere. Hydrogen is discharged at the cathode in preference to sodium ions. This is because sodium is more reactive than hydrogen in the metals reactivity series.

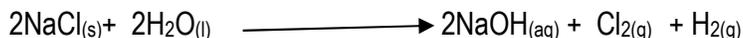


At the anode:

At the anode both the chloride ions and the hydroxide ions migrate there; the chloride ions give up electrons instead of the hydroxide ions (OH⁻) this is because the chloride ions are in much greater concentration than the hydroxide ions. The chlorine atoms chemically combine to form chlorine which is given off as chlorine gas.

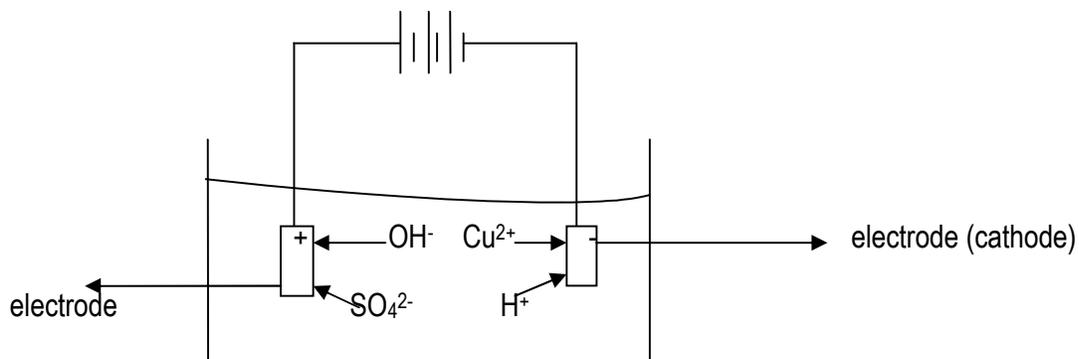


The overall reaction is;



When chlorine gas and hydrogen gas bubble off, sodium ions and hydroxide ions remain. A solution of sodium hydroxide is formed. The evaporation of water leads to the formation of solid sodium hydroxide

(c) Electrolysis of Aqueous Copper (II) Sulphate using inert electrodes (platinum)

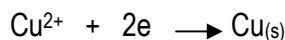


The ions present in the solution are;

1. Copper (II) ions and sulphate ions from copper (II)
2. Hydroxide ions and hydrogen ions from water

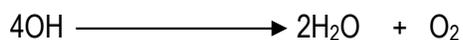
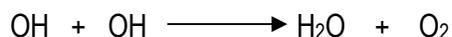
At the cathode;

The half cell reaction taking place at the cathode involves the Copper II ions and hydrogen ions that migrate to the cathode. At the cathode copper two ions gain electrons, become atoms and are discharged in preference to the hydrogen ions because Copper is less reactive than hydrogen in the metals reactivity series. Copper is deposited as a brown layer on the cathode. The half cell equation reaction is:



At the anode;

The hydroxide and sulphate ions migrate to the anode. OH^- being less reactive is discharged in preference to the sulphate ions. The OH^- combines in pairs to form water and oxygen. The oxygen gas is given out to the atmosphere. The half cell reaction taking place at the anode is:



NB: The blue colour of Copper II sulphate ions is caused by the Copper ions and since the Copper ions are discharged the solution loses the blue colour and becomes colourless.

(d) Electrolysis of molten (fused) Lead II Bromide;

In a molten or fused salt, water is not involved and therefore, the ions presents are only those from Lead Bromide (Pb^{2+} and Br^-)

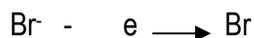
At the cathode:

Lead ions migrated to the cathode and gains two electrons to become Lead atoms, according to the equation;



At the anode

Bromide ions migrate to the Anode where they lose one electron each to become bromine atoms. Bromine atoms chemically combine to become bromine molecules.



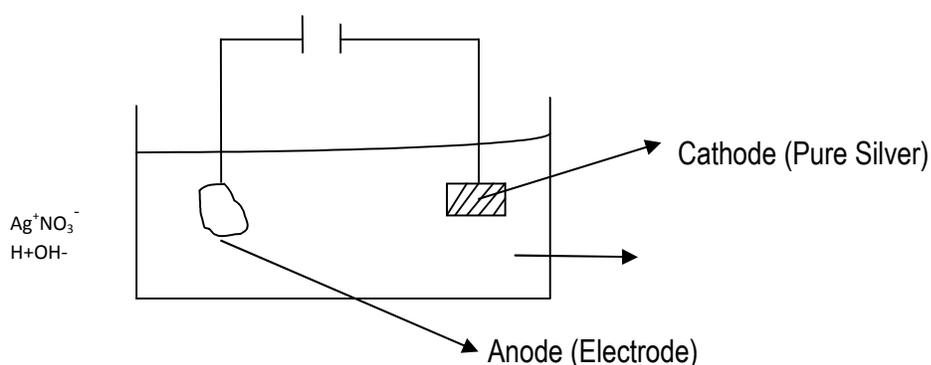
NB: Electrolysis of substances involving the halogens must be done in a fumed hood cupboard so that the poisonous gases can be released into the atmosphere through the chimney.

INDUSTRIAL APPLICATION OF ELECTROLYSIS

- (i) Electrolysis in industries is used to extract very reactive metals such as aluminium, sodium and calcium from their chlorides.
- (ii) Electrolysis in industries is used to purify metals such as the refinery of copper.
- (iii) Electrolysis is also used in formation of hydroxide such as sodium, potassium hydroxide.
- (iv) In industries electrolysis is used in electro-plating.

Electro-plating:

Electro-plating is the coating of one metal onto another by using electrolysis. Electro-plating is done to protect the metal from rust and also to improve its appearance. Objects or metals to be electro-plated are made to be the cathode. For example in silver electro-plating, the electrolyte is silver salt such as silver nitrate.



At the anode:



At the cathode:



In correct conditions, the silver layer deposits on to the cathode article and later the article can be highly polished.

Chromium plating is much used in to improve the appearance of steel parts and protect them from rusting. The steel is usually plated first with nickel or copper because chromium does not adhere well to steel objects.

POSITION OF IONS IN THE ELECTRO-CHEMICAL SERIES

CATIONS | **ANIONS**



K ⁺	Decreasing difficult of discharge	SO ₄ ²⁻ aq	Increasing difficult of discharge
Na ⁺ aq		NO ₃ ⁻ aq	
Ca ²⁺ aq		Cl ⁻ aq	
Mg ²⁺ aq		Br ⁻ aq	
Zn ²⁺		I ⁻ aq	
Fe ³⁺		OH ⁻ aq	
Pb ²⁺ aq			
H ⁺ aq			
Cu ²⁺ aq			
Ag ⁺ aq			

The exception to the above arrangement is that the concentration of a particular ion in the solution may determine whether it is to be discharged or not for example if the ion is in greater concentration it is discharged first despite its position in the series.

FARADY LAWS:

The laws expressing quantitative results were first stated by Sr. Michael Faraday. These laws assert that the amount of substance in moles of an element discharged during electrolysis depends on;

- The time of passing the steady current.
- The magnitude of the steady current passed.
- The charge of the ion of the element.

Factors **A** and **B** can combine in the same experiment because the product of the current measured in amperes and the time measured in seconds gives a measure of electricity known as the charge measured in coulombs. The symbol for charge is **Q**. The above description is summarized in the following formula.

$$\text{Quantity of electricity (Coulombs)} = \text{Current (amps)} \times \text{Time (seconds)}$$

- What is the quantity of electricity when 2 amperes flow in 5 minutes?
- Find the quantity of electricity passing when a steady current of 0.45A flows for 25 minutes.

Solution:

- $$Q = A \times t$$

$$Q = 2 \times 5 \times 60$$

$$Q = \underline{600 \text{ coulombs}}$$
- $$Q = A \times t$$

$$Q = 0.45 \times 25 \times 60$$

$$Q = \underline{675 \text{ coulombs}}$$

MASS PRODUCED BY ELECTROLYSIS

Faraday's first law:

This law states that the mass produced at the electrode is directly proportional to the quantity of electricity passing through the electrolyte and the time taken i.e if we double the coulomb the mass also doubles. The table below shows the results of the mass of copper deposited when a steady current of 0.21Amps is passed through an electrolyte of Copper (II) Sulphate at different length of time in minutes; thus 15 minutes, 30 minutes, 45 minutes and 60 minutes.

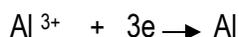
CURRENT	TIME IN SECONDS	QTY OF ELECTRICITY IN COULOMBS	MASS OF COPPER DEPOSITED IN GRAMS
0.21	15 x 60 = 900	189	0.063
0.21	30 x 60 = 1 800	378	0.129
0.21	45 x 60 = 2 700	567	0.187
0.21	60 x 60 = 3 600	756	0.250

Exercise

- Using the correct scale plot the mass of copper deposited against the quantity of electricity passed and state the shape of the graph.

FARADAY'S CONSTANT;

The least quantity of electricity required to liberate 1 mole of an element is 96 500 coulomb and this applies only to elements that have ions with a single charge e.g. Ag⁺, Cl⁻ and Na⁺. This quantity of electricity is given a special name and is called Faraday's constant. Thus one Faraday will liberate one mole of hydrogen ions, 1 mole of sodium ions, and 1mole of silver ions. For elements that have ions with a charge of more than one; the Faraday is multiplied by the magnitude of the charge. The magnitude of a charge is equal to the number of moles of electrons required e.g. to liberate 1mole of aluminium atoms, 3moles of electrons are needed



2 moles of electrons are needed to liberate one mole of Copper or 64 grams of Copper,

TO DETERMINE THE MASS DEPOSITED AT ELECTRODE:

To determine the mass deposited at the electrode, the following procedure is employed.

- Write down the half cell reaction at the cathode.
- Calculate the quantity of electricity in coulombs
- From the relationship of the equation of Faraday by one mole of electrons equals to one mole of a substance or molar mass or dm³ in case of gases produced in the equation.
- From the relationship calculate the proportionality of substances.

Example

Calculate the Mass of silver deposited at cathode from Silver Nitrate solution when 0.45Amps flows for 25 minutes.



$$\begin{aligned}
 Q &= I \times t \\
 Q &= 0.45 \times 25 \times 60 \\
 Q &= 675C \\
 &= \frac{96\,500 \times 675}{675} \text{ g of Ag} \\
 &= \frac{108 \times 675}{96\,500} \\
 X &= \underline{\underline{0.755g}}
 \end{aligned}$$

Exercise

1. Determine the mass of copper deposited at the cathode when 0.1A flows in an electrolyte of copper (II) sulphate for 10minutes
2. What time in hours is required to deposit 135g of aluminium, if a current of 5A was allowed to pass through the circuit?
3. Determine the volume of oxygen produced at the anode measured at room temperature and pressure when a current of 0.1A flows for 10 minutes in a solution of copper(II) sulphate.

Effect of charge of ions on the amount produced at the electrode.

The higher the charge on the ion the less the number of moles of the elements produced or discharged at the electrode. This is because to produce one mole of the element, more than 1 mole of electrons are needed for example, to produce 1 mole of copper, 2 moles of electrons are needed. On the other hand to produce 1 mole of silver, only 1 more of electrons are needed.

Example:

Compare the number of moles produced at the cathode if a steady current of 0.45Amps flows for 25 minutes in the Copper sulphate solution and Silver Sulphate solution.

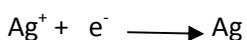
Solution:

$$\begin{aligned}
 Q &= A \times T \\
 Q &= 0.45 \times 25 \times 60 \\
 Q &= 675 \text{ Coulombs}
 \end{aligned}$$

For Copper;

$$\begin{aligned}
 \text{Cu}^{2+} + 2 \text{e}^- &\rightarrow \text{Cu} \\
 2 \times 96\,500 &: 1 \\
 675 &= x \\
 X &= \frac{675}{193\,000} \\
 X &= \underline{\underline{0.0035 \text{ moles of copper}}}
 \end{aligned}$$

For Silver;



$$\begin{aligned} 96\,500 &= 1 \\ 675 &= x \end{aligned}$$

$$X = \frac{675}{96\,500}$$

$$X = \underline{0.007 \text{ moles of Silver}}$$

From the calculations above, the number of moles of silver is double the number of moles of copper

Faraday's Second Law:

This law states that when the same quantity of electricity is passed through solutions of different electrolyte, moles of elements produced at the electrodes are inversely proportion to the charge on their ions. Using the same quantity of the electricity in solutions involving calcium ions, aluminium ions and sodium ions would mean more mass or more moles of sodium would be produced followed by calcium and the least amount would be produced from aluminium. This is because sodium has the least number of charge (1⁺) while calcium has a charge of 2⁺ and aluminium has a highest charge of 3⁺.

Exercises

QUESTION1: Extraction of metals

The main ore of iron, haematite, contains the compound iron (III) oxide, Fe₂O₃. The iron (III) oxide is reduced in a blast furnace to form iron.

- Explain how the furnace is heated to the high temperature of about 1990°C needed for the reduction of the iron (III) oxide.
- Give two ways by which carbon dioxide is produced in the blast furnace.
- Carbon dioxide reacts with carbon to form carbon monoxide. Write a balanced chemical equation for this reaction.
- In the blast furnace, calcium oxide and silicon dioxide react together to form slag according to the equation below:



- State the sources of calcium oxide and silicon dioxide found in the furnace
 - Explain why these two oxides react together.
- Work out the percentage by mass of iron in iron (III) oxide.

QUESTION2

Explain each of the following

- Metals are lustrous. [1]
- Non metals are oxidizing agent [1]
- Ionic compounds have high melting point. [1]
- Solid sodium chloride will not conduct electricity but aqueous sodium chloride will do. [1]
- Difference in conduction of electricity by aluminium and molten aluminium oxide. [1]

QUESTION3: Electro-

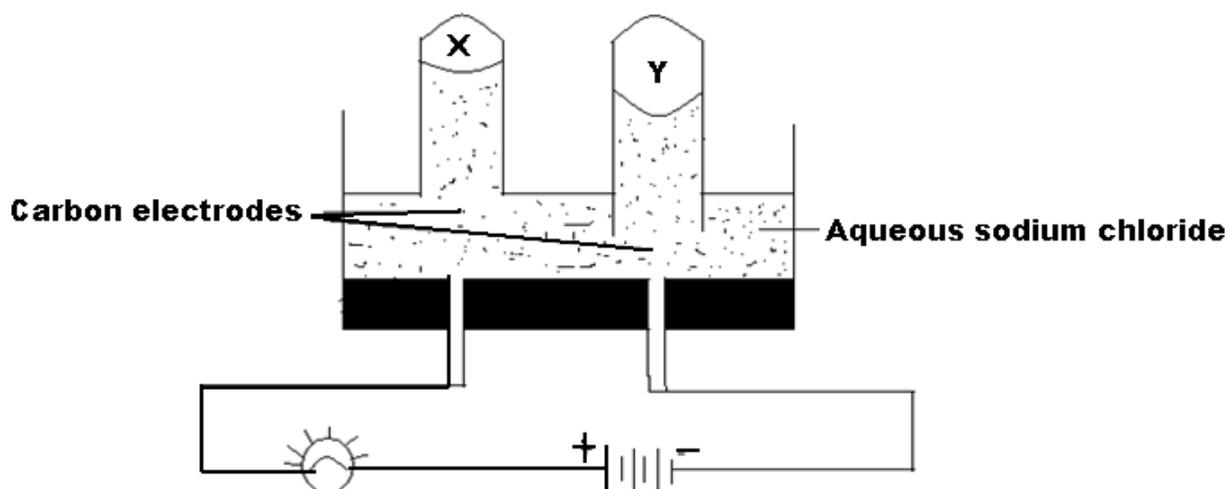
This question is concerned with the following list of substances;

Potassium, lead (II) bromide, oxygen, zinc oxide, hydrogen and carbon. Each substance can be used once, more than once, or not at all. Name a substance from the list above which...

- (a) reacts violently with water [1]
- (b) Conducts electricity when molten but not when solid [1]
- (c) is amphoteric
- (d) has a formula of the type XY_2 [1]
- (e) has an allotrope with a structure similar to that of silica. [1]
- (f) has a low boiling point [1]
- (g) is produced at the cathode during electrolysis of dilute sulphuric acid. [1]
- (h) is used as an electrode in manufacture of Aluminium. [1]
- (i) is the main component in photographic film development. [1]
- (j) is used to assist respiration in hospitals in incubators. [1]

QUESTION 4: Electro-

Aqueous sodium chloride was electrolysed in the apparatus shown below.



- (a) (i) Name the gas X. [1]
- (ii) How can gas X be identified? [1]
- Test: [1]
- Result: [1]
- (b) (i) Name the gas Y [1]
- (ii) Write the equation for the half-reaction in which Y is produced. [2]
- (c) Why is the volume of X smaller than that of Y? [1]
- (d) The electrolysis of aqueous sodium chloride is used to manufacture important chemicals in industry. Name the major product of this electrolysis. [1]

QUESTION5: Electro-

(a) The table below shows the electrical conductivities of some metals relative to iron (iron = 1 unit) and their densities.

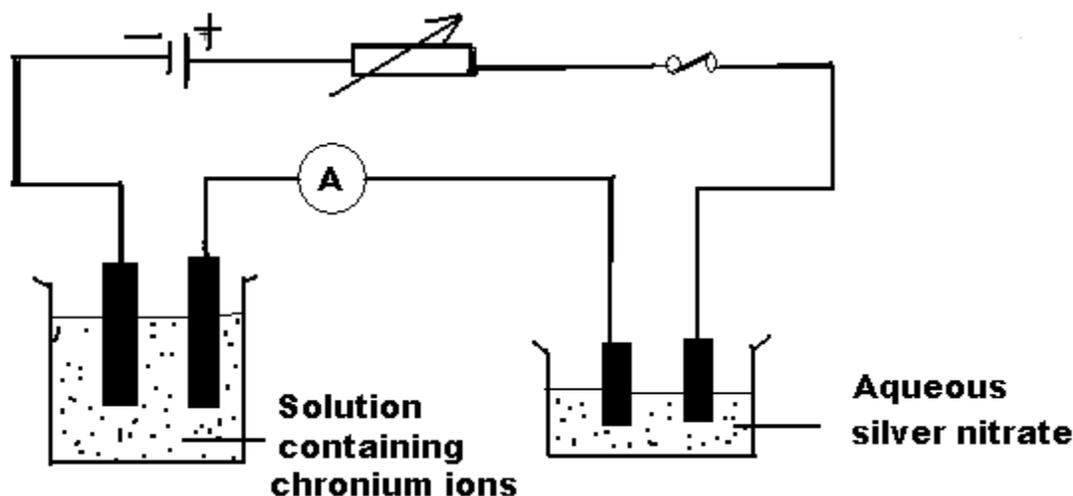
Metal	Electrical conductivity relative to iron	Density in g/cm^3
Aluminium	3.8	2.7
Copper	5.9	8.9
Gold	4.4	19.3
Iron	1.0	7.9
Lead	0.5	11.2

Silver	6.2	10.5
Zinc	1.7	7.1

- (i) Which metal is the best conductor of electricity? Suggest a reason why this metal is not used in electrical cables. [2]
- (ii) Aluminium does not conduct electricity as well as copper. Why is aluminium used in overhead electricity cables rather than copper? [1]
- (iii) Which of the metals in the table has the least electrical conductivity? [1]
- (b) A small sample of copper may be prepared in the laboratory by heating copper (II) oxide with carbon. The gas produced in the reaction turns lime water milky.
- (i) Write a balanced chemical equation for the reaction of copper (II) oxide with carbon. [2]
- (ii) The carbon is said to be a reducing agent. Explain the term reducing agent.
- (iii) If 20g of copper (II) oxide are reduced with carbon, 16g of copper are obtained. What is the mass of copper you would expect to obtain from 8g of copper (II) oxide? [2]
- (iv) Write an equation for the reaction of carbon dioxide with lime water forming a milky solution. [2]
- (v) What is the common name of the compound that causes the effect in (IV)
- (vi) Write the formula of compounds below.
- (A) lime
- (B) Slaked lime

QUESTION6: Electro-

In order to determine the charge on chromium ion, two electrolytic cells were connected in series, one containing an aqueous solution of chromium ions and the other an aqueous solution of silver nitrate as shown in the diagram below.

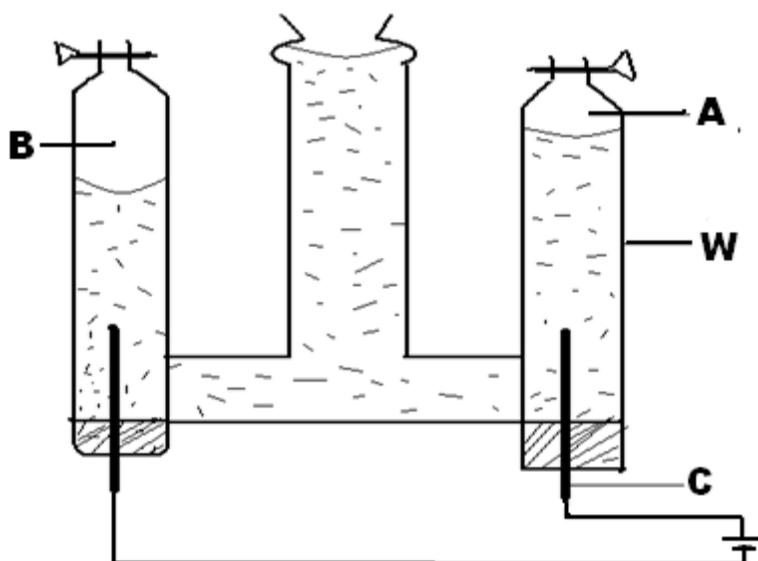


- (a) What name is given to the electrode where chromium and silver are deposited?
- (b) Write down the half reaction showing the discharge of silver ions. [1]
- (c) The two aqueous solutions in the diagram above conduct electricity. What name is given to compounds which conduct electricity and are decomposed? [1]
- (d) At the end of the experiment, it was found that 2.02g of silver and 0.324g of chromium had been deposited. Calculate the charge on a chromium ion. Show

- your working. [4]
- (e) Suggest the formula of the compound, chromium chloride, in the solution. [1]
- (f) Chromium is a metal. Explain how the structure of metals enables them to conduct electricity and be beaten into thin sheets. [2]
- (g) Aluminium is produced industrially by electrolysis of aluminium oxide.
- (i) Name one common ore of aluminium [1]
- (ii) Cryolite is usually added to aluminium oxide during the process. Give a reason why this is done. [1]
- (iii) In the process carbon anodes are regularly replaced. Explain why. [1]
- (iv) Construct an ionic equation to show the production of aluminium. [2]

QUESTION7: Electrolysis

The diagram below shows an experiment carried out to investigate the conductivity of solutions. Pure water was used followed by acidified water.



- (a) (i) What is the name of apparatus W?
(ii) Explain why no reaction is observed when the liquid in apparatus W is pure water.
- (b) A few drops of concentrated sulphuric acid are added to the pure water. Two gases **A** and **B** are collected after 5 minutes.
- (i) Name the gases **A** and **B** collected after 5 minutes.
(ii) Write down the ionic equations to show the formation of the gases **A** and **B**
(iii) What would be the best material for **C**? Give a reason for your choice.
- (c) A current of 0.5A is passed through the acidified water in the 5 minutes of the experiment. (1 Faraday = 96500C).
- (i) Calculate the number of moles of electrons that passes through the acidified water during this time.
(ii) Calculate the volumes of the gases **A** and **B** formed in the 5 minutes interval.
- (b) The diagram below shows some of the properties and reactions of aqueous ammonia and some other substances.

GRADE 12 TERM 1

ORGANIC CHEMISTRY

Organic chemistry is the chemistry of carbon compounds except carbon compounds of carbon dioxide, carbonates, bi-carbonates and carbides.

Carbon occupies a unique position in the periodic table. It is in group four (4) and contains 4 electrons in the valency shell. Carbon exhibits a property of catenations i.e. Carbon atoms can bond to each other through single covalent bonds, double covalent bonds and triple covalent bonds. E.g. carbon can form C-Clong straight chains, long branched C = C chains and rings. Because of the catenations C = C properties Carbon forms an endless variety of Carbon compounds.

Organic compounds are Carbon compounds found in living organisms or are derived from living organisms.

The bodies of mammals contain quantities of muscle tissue which is largely a protein compound containing large amount of atoms. Plants on the other hand contain large amount of cellulose which consists of long chains of Carbon atoms linked together.

HOMOLOGOUS SERIES

Since organic Chemistry consists of uncountable Carbon compounds, Chemists have put these compounds into groups or families called Homologous Series. Homologous series is a family of compounds in which a previous member differs from the next member by CH_2 (ethelylene).

A homologous series is made up of compounds of similar properties. Members in the series have the following characteristics:

- They have a common functional group; a function group is a chemical property of compounds.
- All members have general formula, $\text{C}_n \text{H}_{2n + 2}$.
- All the members show similar chemical characteristics though with varying vigor.
- Each member in the homologous series differs from the next member by CH_2 (ethelylene)
- Physical properties of members change gradually in the group.
- Members in the series can be prepared using similar methods.

ALKANES (Paraffin)

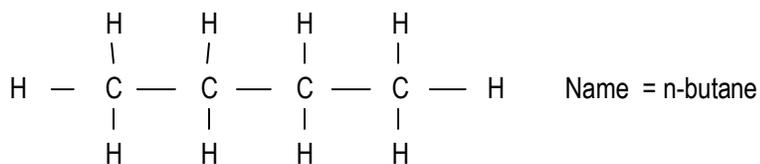
Alkanes are hydrocarbons i.e. their molecules are made only of Carbon and Hydrogen. Alkanes are saturated hydrocarbons because they contain single bonds between Carbon atoms. All alkanes conform to the general formula $\text{C}_n \text{H}_{2n + 2}$: where n is a number of carbons.

The table below shows members of the alkane homologous series.

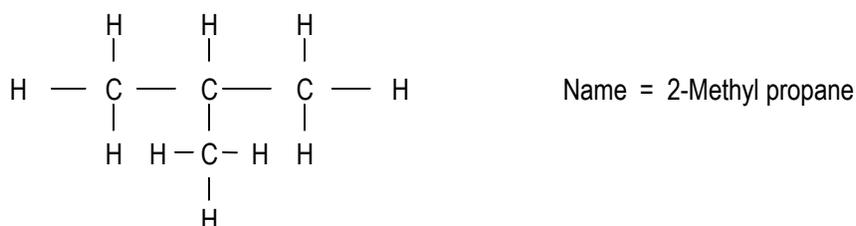
ISOMERISM

Isomerism is the occurrence of two or more compounds with the same molecular formula but different structural formula. Since they have different structural formula they also have different physical and chemical properties. They are caused by the ways in which the Carbon atoms are bonded together. In alkanes isomerism starts from Butane and the number of Isomers increases sharply with the increasing number of Carbon atoms e.g.

Isomers of butane

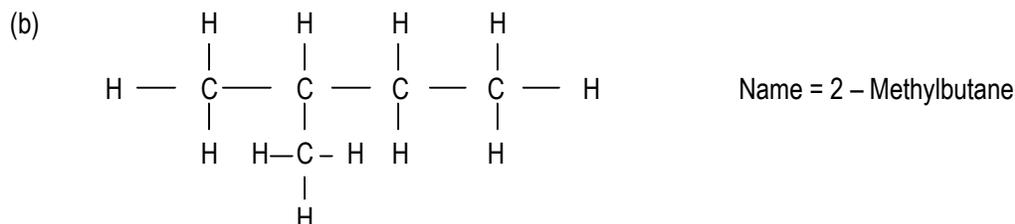
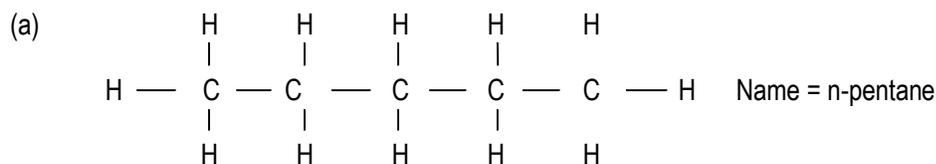


Molecular formula (C_4H_{10})

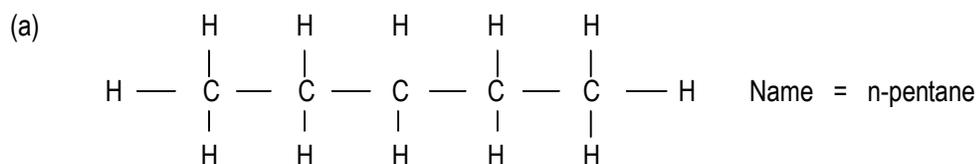


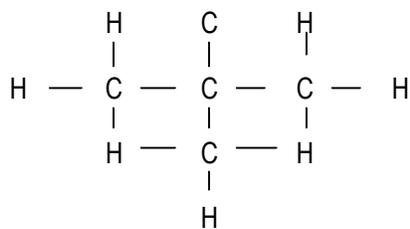
Molecular formula (C_4H_{10})

2. Pentane C_5H_{12}



Isomers of Pentane C_5H_{12}





Name = **2, 2 Methylpropane**

PHYSICAL PROPERTIES

- The melting point and boiling point increase with increase in number of carbon atoms contained.
- They are insoluble in water.
- Their densities increase with increasing molecular mass.
- The table below summarizes the melting, boiling state at r.t.p.

Name	Molecular formula	Mp ($^{\circ}\text{C}$)	State at r.t.p	State at r.t. p
Methane	CH_4	-182°C	-162°C	Gas
Ethane	C_2H_6	-183°C	-89°C	Gas
Propane	C_3H_{10}	-188°C	-42°C	Gas
Butane	C_4H_{10}	-138°C	0°C	Gas
Pentane	C_5H_{12}	-130°C	36°C	Liquid
Hexane	C_6H_{14}	-95°C	69°C	Liquid

CHEMICAL PROPERTIES:

Alkanes are less reactive compounds they do not react with acids and alkalis for example, they are unreactive due to the fact that in their molecules there are no sites or areas where electrons for a chemical reaction can easily be found. All electrons in the valency shell of Carbon are bonded to hydrogen and for this reason alkanes are described as saturated hydrocarbons.

Although they are less reactive they can be made to undergo certain useful reactions such as combustion and substitution.

Combustion of Alkanes:

- (a) Alkanes are a good source of fuel or energy when heated to very high temperatures alkanes burn to produce heat or light. Heat from alkanes is used in:
- Internal engine combustion and the force created is used to drive machines e.g. Petrol, diesel, paraffin etc.
 - Heat from combustion of alkanes is used in gas stoves, Bunsen burners, etc.

The following equations show the combustion of alkanes in different amounts of air.

- In plenty supply of air.**

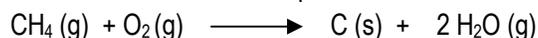
$$\text{CH}_4 + \text{O}_2 \longrightarrow \text{CO}_2 + 2 \text{H}_2\text{O} + \text{Energy}$$
- In low supply of air.**

$$2\text{CH}_4 + 3\text{O}_2 \longrightarrow 2\text{CO} + 4 \text{H}_2\text{O} + \text{Energy}$$

The Carbon Monoxide produce is deadly pollutant as it poisons red blood cells in human beings, there by reducing

the amount of oxygen in blood.

(iii) In limited amounts of air to produce Carbon, water vapour and energy according to the equation.



SUBSTITUTION REACTION:

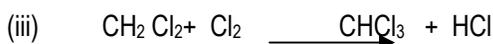
Other useful reaction is substitution reaction and it involves reaction between alkanes and halogens under fused light or the presence of U-ultraviolet The first step in the reaction involves the splitting up of halogen by u-ultraviolet light to create reactive radicals of the halogen. The reactive radical can then substitute the halogen atoms in the alkane molecules. The first step is called the termination step. The following equations show how the reaction proceeds.



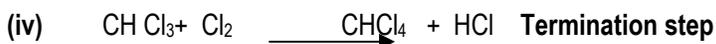
Chloromethane



Dichloromethane



Trichloromethane (Chloroform)



Tetrochloromethane

C.F.C AND OZONE LAYER

Substitution reaction involving halogens and alkanes may result into the formation of chloroform carbon **A.K.A C.F.C** the substances are fairly unreactive in the past they found a lot of uses in fridges as coolant and as propellant aerosol can.

Scientists now believe that the C.F.Cs realized aerosol cans are destroyed in the Ozone Layer (O_3) trioxide. These happen when the inert C.F.Cs are subjected to v.r. in the atmosphere. Reactive radicals are produced which later react with ozone layer causing holes in the ozone layer. This reaction eventually depletes the ozone layer. Our atmosphere protects us from harmful Ultra-violet rays through the Ozone Layer and if the Ozone Layer is depleted greater amounts of ultra-violet rays will react with the earth and will cause in some individuals. Because of the damaging effects that CFCs have on the Ozone Layer. CFCs are regarded as pollutants.

FRACTIONAL DISTILLATION OF CRUDE OIL

Crude Oil is a complex mixture of compounds known as Hydrocarbons. Hydrocarbons are molecules which contain only the elements Carbon and Hydrogen bonded together covalently. Crude Oil is not very useful unless it is processed by the process known as "Oil Refining"

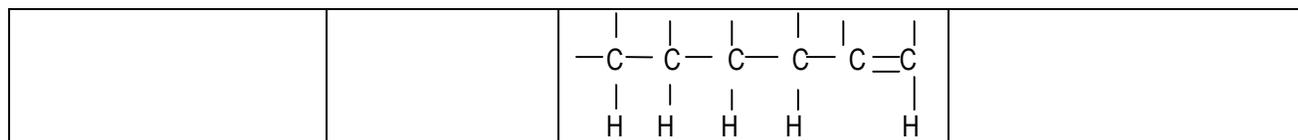
S/NO.	NAME OF FRACTION	NO. OF CARBON ATOMS	BOILING POINTS	USES OF FRACTIONS.
1	Refinery gas	C to C ₄	Below 0°C	For fuel heating and light e.g. Methane
2	Gasoline	C ₅ to C ₆	30°C – 70°C	Petrol fuel in cars
3	Naphtha	C ₇ to C ₁₀	110°C	Making chemicals/plastics.
4	Kerosene	C ₁₀ to C ₁₆	180°C	Fuels in jet engines.
5	Diesel Oil	C ₁₆ to C ₂₀	260°C	Fuels in diesel engines
6	Lubricating	C ₂₀ to C ₃₀	300°C	Oil in cars and other machine for station
7	Fuel Oil	C ₃₀ to C ₄₀		
8	Paraffin wax	C ₄₀ to C ₅₀	310°C	Polish and grease – making candles.
9	Bitumen	C ₅₀ to C ₂₀ And upwards	400°C	For surfacing roads and filling in roofs. Waxed papers and H ₂ O Proofing Martin

ALKENES

Alkenes form another homologous series of hydrogen carbon. Alkene contains at least one carbon to carbon double bond in their molecules. All alkenes conform to the general formula C_n H_{2n} where n is a natural number greater than 1. The simplest member is ethene. Alkenes have names ending in **-ene** this is done by dropping the **a** and replacing it with **e** on the **k** of the parent name of their alkanes.

Because of the double bond in their molecules alkenes are called unsaturated hydrogen carbons. The table below gives the name and formula for the first six members.

Name	Molecular formula	Structural formula	Condensed formula
Ethene	C ₂ H ₄	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{C} = \text{C} \\ \quad \\ \text{H} \quad \text{H} \end{array}$	CH ₂ CH ₂
Propene	C ₃ H ₆	$\begin{array}{c} \text{H} \quad \quad \text{H} \\ \quad \quad \\ \text{H} - \text{C} - \text{C} = \text{C} \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array}$	CH ₃ CH CH ₂
Butene	C ₄ H ₈	$\begin{array}{c} \text{H} \quad \text{H} \quad \quad \text{H} \\ \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} = \text{C} \\ \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$	CH ₃ CH ₂ CH CH ₂
Pentene	C ₅ H ₁₀	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} = \text{C} - \text{H} \\ \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \quad \text{H} \end{array}$	CH ₃ CH ₂ CH ₂ CH CH ₂
Hexene	C ₆ H ₁₂	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \quad \\ \text{H} - \text{C} - \text{C} - \text{C} - \text{C} - \text{C} = \text{C} - \text{H} \\ \quad \quad \quad \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \quad \text{H} \quad \quad \text{H} \end{array}$	CH ₃ CH ₂ CH ₂ CH ₂ CH CH ₂



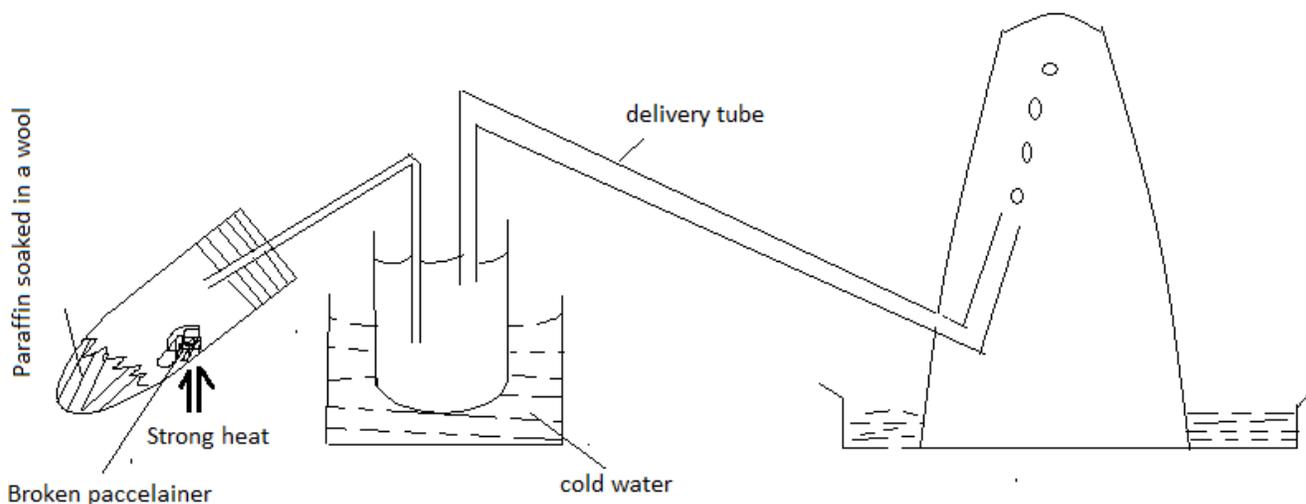
PHYSICAL PROPERTIES OF ALKENES

1. They are insoluble in water.
2. The first four are gases at r.t.p.

PREPARATIONS OF ALKENES

Very few alkenes are found in nature most of the alkenes used by petrol industry are obtained by breaking up larger less useful alkenes. This is usually done by a process called catalyst cracking.

In this process the alkenes over a mixture of aluminum and chromium oxides heated to about 500°C the diagram below shows a simple apparatus to crack dodecane (Paraffin molecules in the laboratory.)



Aluminium Oxide

The reaction in the above reaction will result into a decane and an ethane according to the equation.

1. Dodecane $\xrightarrow[\text{Al oxide \& Chromium oxide}]{\text{heat}}$ decane + ethane
2. $C_{12}H_{26} \xrightarrow[\text{Al oxide \& Chromium oxide}]{\text{heat @ } 500^{\circ}\text{C}}$ $C_{10}H_{22} + C_2H_4$

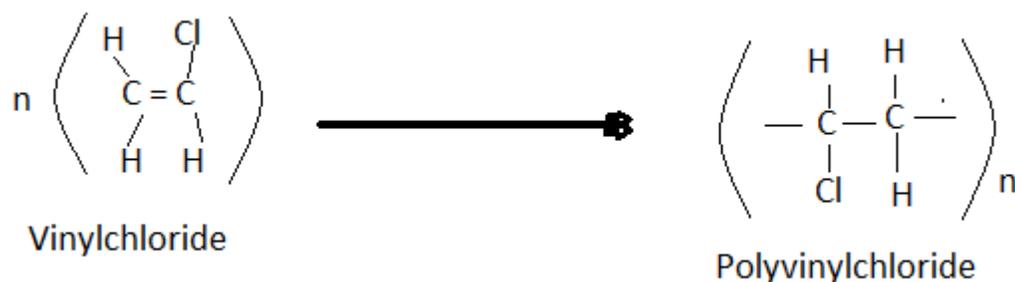
CHEMICAL PROPERTIES

Alkenes are unsaturated hydrocarbons because they have a double bond between two carbon atoms. The double bond is rich in electrons and therefore elements such as chlorine bromine and others can add to the double bond to form saturated compounds as illustrated in the addition reaction.

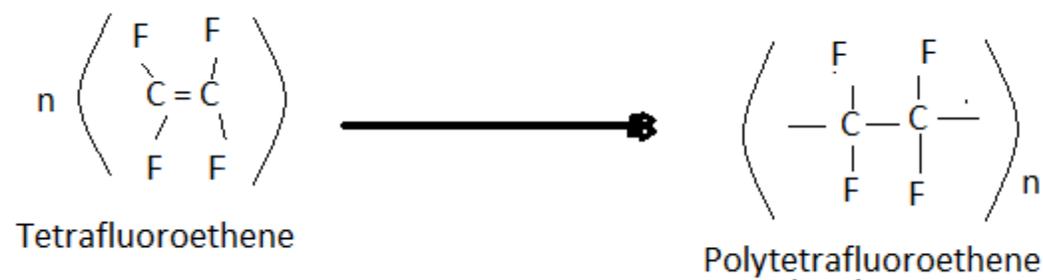
Where $n > 10,000$

OTHER ADDITION REACTIONS

Many other addition reactions that produce useful polymers have been designed and developed. These include:



Monomers of vinylchloride or Chloroethene form Polyvinyl Chloride $n > 1\,000$



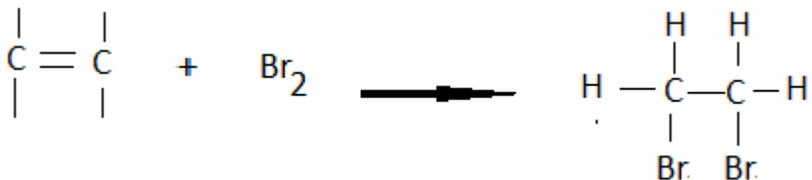
Polyvinylchloride is strong, hard and less flexible polymer. It is used to make p.v.c pipes, rain coats, plastic buckets, drinking bottles etc.

Poly Tetrafluoroethane is non stick and withstand high temperature used to make handles of frying pans.

TEST FOR ALKENES

Alkenes are tested by using either brown bromine water or Potassium Permanganate (VII). Bromine water is added to the double bond to form 1,2 dibromo alkane.

Thus:



The bromine water gets decolourised or the bromine H₂O become colourless. Ethene reacts rapidly with acidified Potassium Permanganate (VII). The green colour of the Potassium Permanganate disappears and the ethene is converted to an ethanol by water and oxygen. The decolourisation of ethene form gaseous alkanes.

ALKANOLS OR ALCOHOL

Alcohol from another homologous series with a General formula C₂H_{2n} + 1 OH where n is a natural number. All alcohol passes an OH functional group. The functional group is a group of atoms of a number of double bonds that determine the chemical characteristics of a compound. The table below gives some of the Alcohols.

Name	Molecular formular	Structural formular
Methanol	CH ₃ OH	$ \begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{OH} \\ \\ \text{H} \end{array} $
Ethanol	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} $
Propanol	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} $
Butanol	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} $
Pentanol	C ₅ H ₁₁ 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} $
Hexanol	C ₆ H ₁₃ 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} $

		$ \begin{array}{cccccc} \text{H} & - & \text{C} & - & \text{OH} \\ & & & & & & & & & & & & \\ & & \text{H} & & \end{array} $
Heptanol	$\text{C}_7 \text{H}_{15} \text{OH}$	$ \begin{array}{ccccccc} \text{H} & & \text{H} \\ & & & & & & & & & & & & \\ \text{H} & - & \text{C} & - & \text{C} & - & \text{C} & & \text{C} & - & \text{C} & - & \text{C} & - & \text{OH} \\ & & & & & & & & & & & & & & \\ \text{H} & & \end{array} $
Octanol	$\text{C}_8 \text{H}_{17} \text{OH}$	$ \begin{array}{ccc} \text{H} & & \text{H} \\ & & \\ \text{H} & - & \text{C} & - & \text{C} & - & \text{OH} \\ & & \\ \text{H} & & \text{H} \end{array} $
Nonanol	$\text{C}_9 \text{H}_{19} \text{OH}$	$ \begin{array}{cccccccc} \text{H} & \text{H} \\ & & & & & & & & \\ \text{H} & - & \text{C} & - & \text{OH} \\ & & & & & & & & & & & & & & & & \\ \text{H} & & \end{array} $
Decanol	$\text{C}_{10} \text{H}_{21} \text{OH}$	$ \begin{array}{cccccccccc} \text{H} & \text{H} \\ \text{H} & \text{C} & \text{OH} \\ \text{H} & \text{H} \end{array} $

PHYSICAL PROPERTIES OF ALCOHOLS

1. The boiling points increases with increasing molecular mass e.g methanol and ethanol are liquids at room temperature and pressure.
2. They dissolve in water (soluble)
3. They are colourless substances.
4. They have a characteristic smell e.g. Beer.

CHEMICAL PROPERTIES:

Alcohol contains the hydroxyl functional group OH which is responsible for the chemical behavior. Alcohol such as ethanol can burn to produce heat, CO_2 and H_2O as the case is with methyl rated spirit

(a) Combustion:



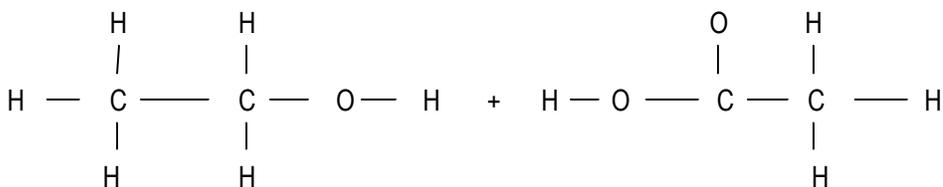
(b) Reaction with Alkali Metals

Alcohols such as ethanol react with alkali metals such as Na (Sodium) to produce hydrogen gas and Sodium ethoxide.



(c) Esterification:

Alcohol reacts with organic acids to form esters e.g. ethanol reacts with ethanoic acid to form an ethyl ethanoate.



ETHANOL

Ethanol is the main ingredient in beers and wine. The strength of beers and wines depends on the concentration of ethanol in that particular beer or wine which is mostly expressed in percentages.

• USES OF ETHANOL:

- Used as a solvent for fats and other sources such as dyes and paints.
- Used as a beverage in beers and wines.
- Used as a disinfectant e.g. methylated spirit.
- Used as a fuel for heating and lighting.
- Used as a solvent in liquids.

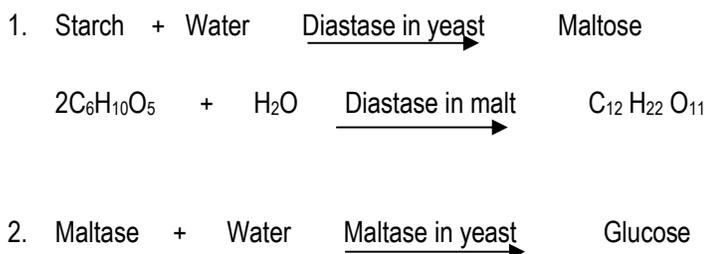
MANUFACTURE OF ETHANOL:

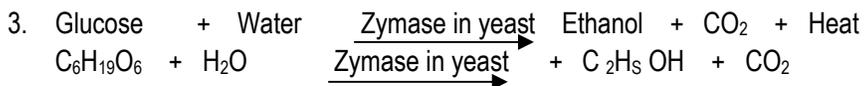
Ethanol is by far the most important of all the alcohols and it is often just called alcohol. Ethanol can be produced by fermentation and by hydration of ethene.

(a) FERMENTATION

Any substance that contains starch, cellulose, sucrose, maltose and any of the mono saccharides can be fermented to produce ethanol. Industrially starch from maize, potatoes, cassava, sorghum and Barley etc. is under pressure cooked to release starch granules sprouted barley for 1 hour at 60°C. Malt supplies an enzyme called diastase the diastase hydrolyses the starch to maltose.

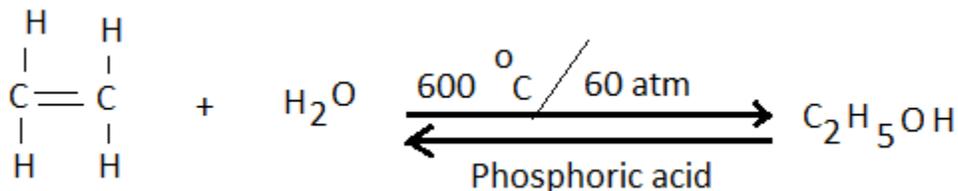
Yeast is added to maltose at r.t.p. The enzyme called Maltase in yeast hydrolyses maltose to glucose. Another enzyme of yeast called Zymase hydrolyses glucose to ethanol and carbon dioxide. The hydrolysis of glucose to ethanol is an exothermic reaction. The three main chemical equations for fermentation are:





(b) Hydration of Ethene:

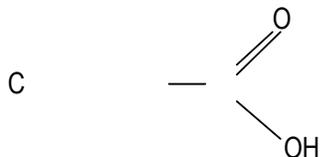
Ethene is obtained by cracking long chains of alkanes from crude oil. Ethene and steam are passed through phosphoric acid at 600°C and 60atm. Steam is added to the double bond in the ethene molecule to produce ethanol.



The hydration of ethane is a reversible reaction hence it is a continuous reaction. Ethene is red in one end of the reaction container and ethanol comes out at the other end. Unreacted ethane is recycled.

ORGANIC ACIDS OR ALKANOIC ACIDS OR CARBOXYLIC ACIDS.

Alkanol acids form a homologous series of the general formula $\text{C}_n\text{H}_{2n-1}\text{COOH}$. The characteristic functional group is



Alkanol acids are named as if they are derived from alkanes by replacing one of the hydrogen atoms by the COOH

The table below shows the first five members of the series:-

Name	Condensed	Structured formular	Molecular formular
Methanoic Acid	Molecular HCOOH	$ \begin{array}{c} \text{O} \\ \\ \text{H} - \text{C} - \text{OH} \end{array} $	HCOOH
Ethanoic Acid	CH ₃ COOH	$ \begin{array}{c} \text{H} & \text{O} \\ & \\ \text{H} - \text{H} - & \text{C} - \text{OH} \\ \\ \text{H} \end{array} $	C ₂ H ₃ COOH
Propanoic Acid	CH ₃ CH ₂ COOH	$ \begin{array}{c} \text{H} & \text{H} & \text{H} & \text{O} \\ & & & \\ \text{H} - \text{C} - & \text{C} - & \text{C} - & \text{C} - \text{OH} \\ & & \end{array} $	C ₂ H ₅ COOH

		H H H	
Butanoic Acid	CH ₃ CH ₂ CH ₂ COOH	$ \begin{array}{cccccc} & \text{H} & \text{H} & \text{H} & \text{H} & \text{O} \\ & & & & & \\ \text{H} & - \text{C} - \\ & & & & & \\ \text{OH} & & & & & \\ & \text{H} & \text{H} & \text{H} & \text{H} & \\ \end{array} $	C ₃ H ₇ COOH
Pentanoic Acid	CH ₃ CH ₂ CH ₂ CH ₂ COOH		C ₄ H ₁₁ COOH

Physical Properties:

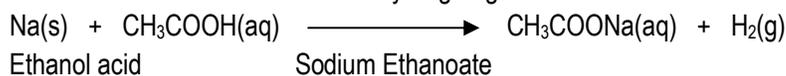
Like inorganic acids they are:

- Colourless
- Soluble in water
- Have a pH of less than seven
- Turn blue litmus paper red.
- Has a sour taste e.g. vinegar.

Chemical Properties:

Like the mineral acids Carboxylic acids react with;

- (a) Reactive metals to form a salt and hydrogen gas.



- (b) React with Carbonates to liberate Carbon dioxide and to form water and a salt.



- (c) React with alcohols to form esters

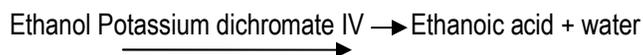


Formation of Ethanoic Acid

Ethanol can be oxidized to ethanoic acid by powerful oxidizing agents such as warm acidified Potassium dichromate VI or by atmospheric oxygen.

1. Oxidation by warm acidified Potassium Dichromate

Ethanol can be oxidized to ethanoic acid more rapidly by warming it with acidified Potassium dichromate solution, during the reaction the orange colour of Potassium dichromate changes to a dark green colour as the ethanol is oxidized to ethanoic acid and water. The equation below summarizes the reaction.



Or



2. Oxidation of Ethanol to Ethanoic acid:

When wines or beers are exposed to atmospheric oxygen they eventually turn to vinegar which is a dilute solution of ethanoic acid. This happens because bacteria in air use oxygen to oxidize the ethanol in beers and wines to ethanoic acid.

Ethanol $\xrightarrow{\text{atmospheric oxygen}}$ ethanoic acid + water.



ESTERS:

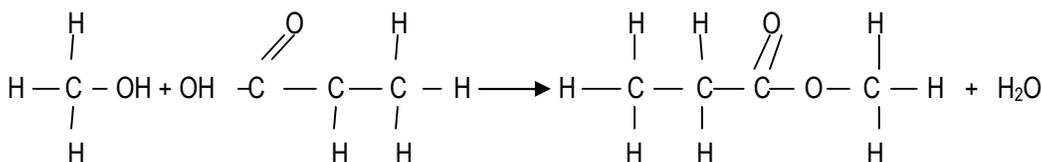
Esters are organic compounds that occur naturally in plants and animals but they can also be synthesized in the laboratory in reacting alcohols and organic acids usually esters have a pleasant fruity smell or pleasant aroma. The smell in fruits is due to esters. Each fruit has its ester that gives its characteristic smell. Short chain esters are gases and mostly found in fruits. Long chain esters are found in fats and oils.

USES OF ESTERS:

- (i) Used in perfumes
- (ii) Used in food flavouring
- (iii) Used in disinfectant.

NAMING OF ESTERS

The naming of esters formed follows the order: Alcohol and then organic acids. For example if methanol reacts with propanoic acid, the ester formed will be called methyl propanoate.



Other esters can be named in a similar way. The alky part is from the alcohol while the other part from the carboxylic acid. For example propyl butanoate is formed from the reaction between propanol and butanoic acid.

MACRO-MOLECULES (POLYMERS)

Macro-molecules or Polymers are synthesized from smaller units called Monomers. Monomers chemically combine together to form polymers. There are artificial (synthetic) polymers and natural polymers.

ARTIFICIAL POLYMERS:

Include fats, starch, cellulose, rubber, protein silk.

Macro-molecules are formed either by addition reactions; the process by which monomers chemically combine together to form a polymer without eliminating a molecule. On the other hand condensation polymerization is the process by which monomers chemically combine together to form a polymer with the elimination of a simpler molecule usually (H₂O) water.

(a) ADDITION POLYMERISATION:

Usually occurs between monomers that have double bonds between carbon atoms. The double bonds break and the monomer molecules with Thousands of Carbon atoms resulting into a macro-molecule. A variety of plastics are made through addition polymerization.

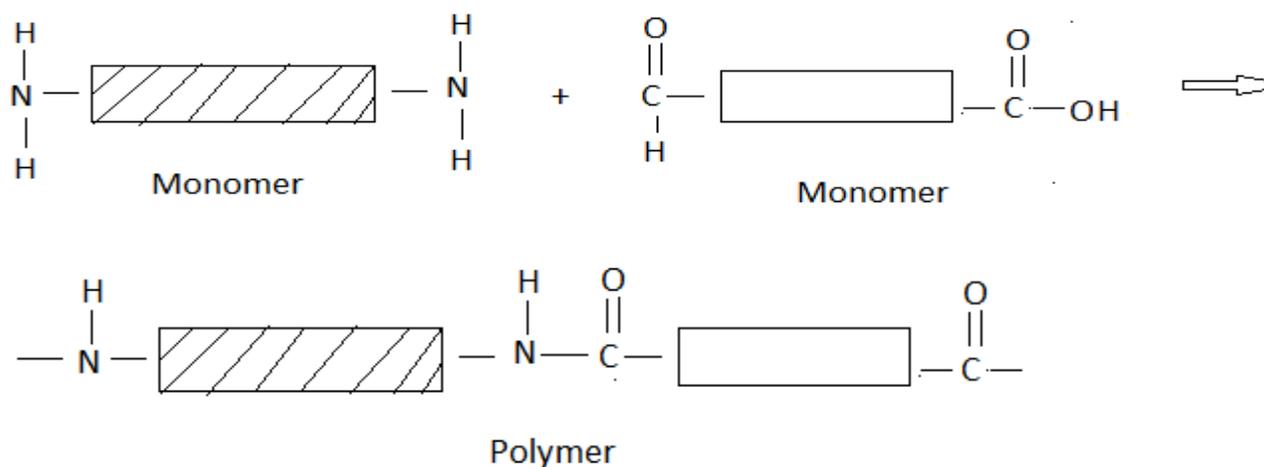
Monomer	Part of the polymer	Equation for reaction	Uses
$\begin{array}{c} \text{H} & \text{H} \\ & \backslash \\ \text{C} = & \text{C} \\ & \\ \text{H} & \text{H} \end{array}$	$\begin{array}{cccccc} \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\ & & / & & & \\ -\text{C} - & \text{C} \\ & & \backslash & & & \backslash \\ \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \end{array}$	$\begin{array}{c} \text{H} & \text{H} \\ & \\ \text{C} = & \text{C} \\ & \\ \text{H} & \text{H} \end{array} \quad n \longrightarrow \begin{array}{cccc} & & & / \\ \text{C} - & \text{C} - & \text{C} - & \text{C} \\ & & & \end{array}$	Carrier bag
$\begin{array}{c} \text{H} & \text{H} \\ & \backslash \\ \text{C} = & \text{C} \\ & \\ \text{H} & \text{H} \end{array}$	$\begin{array}{cccccc} \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\ & & & & & & \\ \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} & \text{C} \\ & & & & & & \\ \text{H} & \text{Cl} & \text{H} & \text{Cl} & \text{H} & \text{Cl} & \text{H} \end{array}$	$\begin{array}{c} \text{H} & \text{H} \\ & \\ \text{C} = & \text{C} \\ & \\ \text{H} & \text{H} \end{array} \quad n \longrightarrow \begin{array}{cccc} \text{H} & \text{H} & \text{H} & \text{H} \\ & & & \\ \text{C} & \text{C} & \text{C} & \text{C} \\ & & & \\ \text{Cl} & \text{H} & \text{H} & \text{Cl} \end{array}$	Making PVC pipes Making rain coats

CONDENSATION POLYMERISATION:

Condensation polymerization occurs between monomers that have reactive sites. These sites usually are atoms of hydroxides, hydrogen or halogens. In the process a simpler molecule usually or hydrogen chloride is eliminated. After eliminating a simpler molecule, the remaining molecules then bond together to form long chain molecules. Some artificial or synthetic polymers that are made through condensation polymerization include nylon and terylene.

NYLON:

Nylon 66 is made by reacting two different monomers together. The two different monomers are 6,1 diamino hexane, and 1,6 dioylchloride hexane. The product of this reaction is a molecule of a nylon. The reaction is shown in the following equation below:



The two monomers in the molecule of nylon are joined together through the amide link.

The molecule of nylon can link to other nylon molecules to form a long chain nylon fibre as shown in the following:-

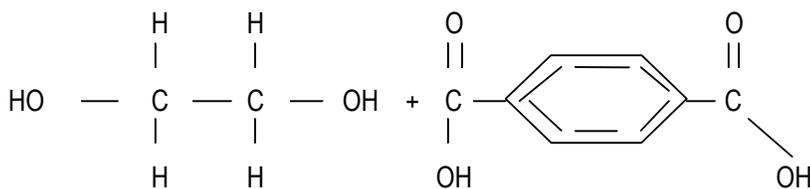
Where $n > 1000$

USES OF NYLON:

- (a) Used for clothing (make nylon clothes)
- (b) Used for making curtains
- (c) Used for making ropes
- (d) Used for making fishing nets
- (e) Used for making convert belts
- (f) Used for making lawn tennis and badminton request

(b) TERYLENE:

Terylene is manufactured by condensation polymerization of $\frac{1}{4}$ benzene dicarboxylic acid and 1,2 diethanol according to the equation.

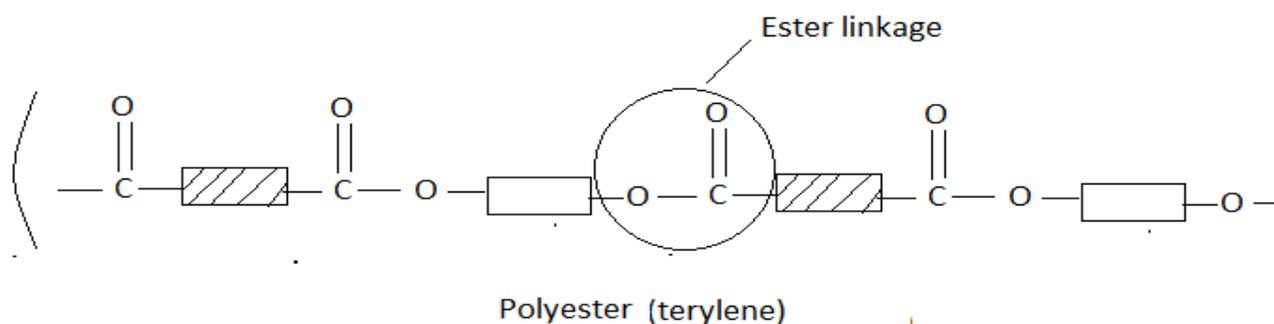


Ethane 1,2 diol

1,4 Benzene dicorboxylic acid



The molecule of Terylene repeats itself over and over through ester linkage and produce a long chain molecule called poly ester (Terylene).



USES OF TERYLENE:

1. Used to make climplins suits.
2. Makes curtains/fishing lines, convey belts.

LIPIDS:

Fats are high molecular weight naturally occurring in esters. Fats are formed from glycerol acting as alkanol (alcohol) and a wide range of long carboxylic. Some of the carboxylic acid involved are:

- | | |
|--|---------------|
| (a) $\text{CH}_3(\text{CH}_2)_{10}\text{COOH}$ | Lauric acid |
| (b) $\text{CH}_3(\text{CH}_2)_{14}\text{COOH}$ | Palmitic acid |
| (c) $\text{CH}_3(\text{CH}_2)_{12}\text{COOH}$ | Myristic acid |
| (d) $\text{CH}_3(\text{CH}_2)_{16}\text{COOH}$ | Stearic acid |

The long chain Carboxylic acid reacts with an alcohol such as 1,2,3 tripropanol to form a complex ester called fats e.g.



USES OF LIPIDS:

1. Fats and oils are sources of energy in plants and animals.
2. In industries fats and oils are used to manufacture soap (Soapification).
3. In plants and animals fats and oils are used to make phospholipid bilayer of cell membrane.
4. Fats are solvents for soluble vitamins (A, D, E and K).

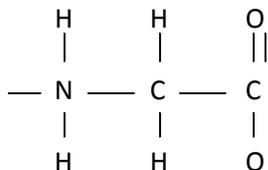
HYDROLYSIS OF LIPIDS:

Million of tonnes of soap are manufactured worldwide every year. Soap is manufactured by heating natural fats and oil of either plants or animals with a strong alkali hydrolysis of fats. The esters are broken down in the presence of water (hydrolysis). The esters break down into fatty acids and glycerol. Fatty acids react with the alkali like sodium hydroxide. This is called alkali hydrolysis of fats. Fatty acids react to form soap and glycerol remains unreacted. The glycerol is removed and used to make glycerine. The equation of the reaction is:

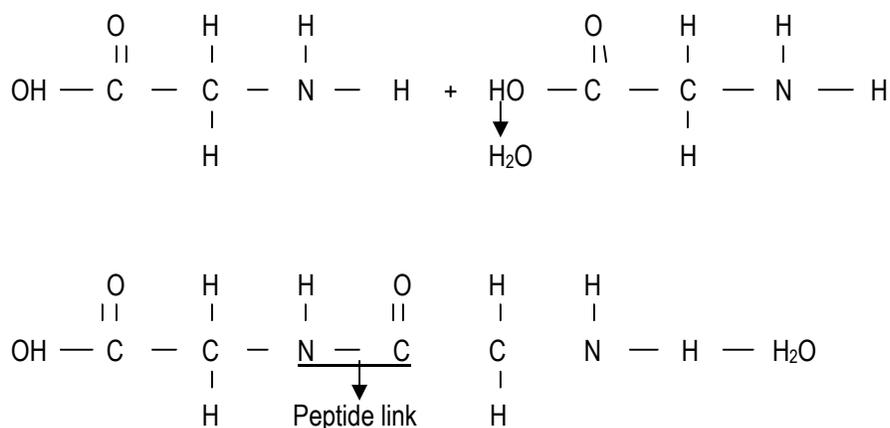
About 27 different Amino acids have been identified by scientists. Two simple Amino acids are:-

- (a) Alanine
- (b) Glycine

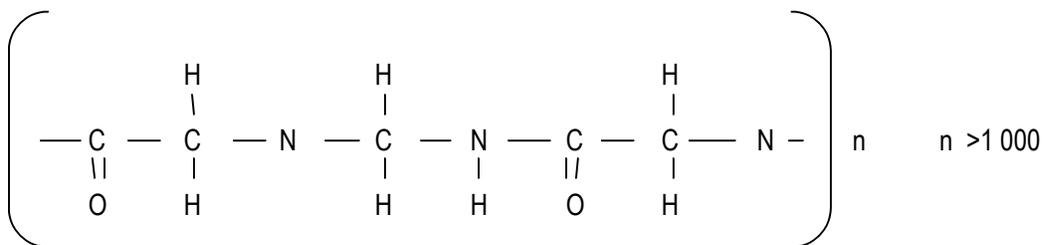
The molecule of glycine is: $\text{NH}_2\text{CH}_2\text{COOH}$ and its structural formula is:



Molecules of the glycine can combine together to form a protein's molecules through Amide links, but in case of proteins the links are called Peptides links, hence the name polypeptide.



The structure above can repeat itself several times to form a molecules of proteins.



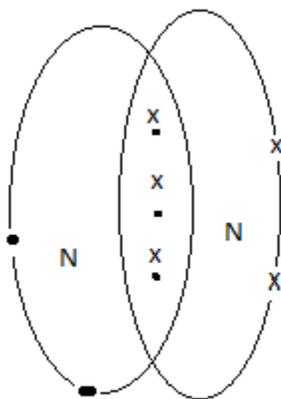
Proteins such as hair, names etc. can be hydrolyzed the Amino acid and the different Amino acids in the particular protein can be identified by using chromatography.

Grade 12 term 2

NON- METALS

NITROGEN:

Nitrogen as an element is found in group five in the Periodic Table. At room temperature and pressure it exists as a diatomic gas molecule. Since three pairs of electrons are shared, the bonds in nitrogen gas are triple single covalent bonds. These bonds are very strong requiring a large amount of energy to break and as a result nitrogen gas is very inert.



IMPORTANCE OF NITROGEN

Nitrogen is an essential element, necessary for the synthesis of proteins both in plants and animals. As a gas it occupies 78% of the air by volume and it is very inert. Being inert it prevents the atmosphere from catching fire. Nitrogen fixing is the process of changing nitrogen from the inert form to the reactive form. Animals obtain the nitrogen they need by feeding on plants and other animals. Most plants obtain the nitrogen they need from the soil as Nitrates.

Nitrogen, Potassium and Phosphorus from the three macro elements needed for proper plant growth.

Element	Use	Deficiency sign
Nitrogen	Synthesis of plant protein	Leaves turn yellow stunted growth
Phosphorous	For making ATP	Leaves turn purple
Potassium	Inactivates enzymes that regulates photosynthesis.	Leaves become scotched (Chlorosis)

Properties of Nitrogen as a gas

1. It is a colourless gas
2. It has no smell
3. It is only slightly soluble in water
4. It is very un reactive compound to oxygen

Chemical Properties

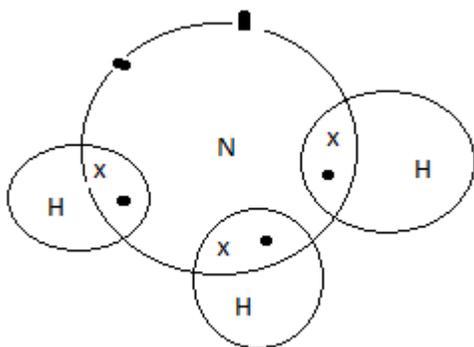
1. It is very un reactive. However, it can react with hydrogen to form ammonia at high temperatures, and high pressure in the presence of a catalyst.
2. It also combines with oxygen at high temperature to form oxides of nitrogen in internal combustion engines and during lighting.

Uses of Nitrogen

1. Used in the manufacture of ammonia
2. Used in the manufacture of fertilizers. E.g. Ammonium Nitrate and super phosphate of ammonia.
3. Used to quick – freeze food and keep it fresh.
4. Used in electric bulbs to provide an inert atmosphere.

AMMONIA:

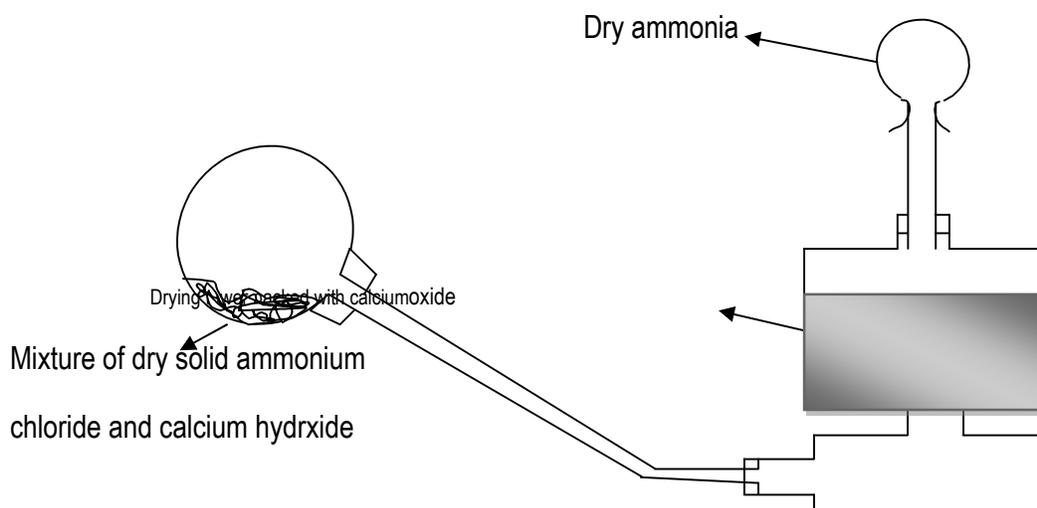
Ammonia is a gas with the formula NH_3 . The bonding in Ammonia is shown in the diagram below:



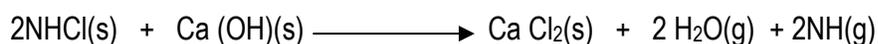
Laboratory preparation of Ammonia.

Ammonia is prepared in the laboratory by heating any Ammonium compound with a strong base such as Sodium hydroxide and Calcium hydroxide using an aluminium foil as a catalyst. The base drives out or displaces ammonia gas from the compound

The apparatus for the preparation of ammonia is shown below.



In reaction above the chemical equation is;



Physical Properties of Ammonia

1. It is colourless
2. It is less dense than air.
3. It has sharp pungent smell
4. It is very soluble in water
5. It is easily liquefied by cooling to -33°C air by compressing it.

Chemical Properties

1. It reacts with hydrogen chloride gas to form a white smoke of tiny particles of ammonium chloride;

$$\text{NH}_3 + \text{HCl(g)} \longrightarrow \text{NH}_4\text{(Cl)}$$
2. It turns damp red litmus paper blue; the ammonia reacts with water on the litmus paper to form ammonium ions and hydroxide ions;

$$\text{NH}_3\text{(g)} + \text{H}_2\text{O(l)} \longrightarrow \text{NH}_4^+\text{(aq)} + \text{OH}^-\text{(aq)}$$
3. Since ammonia solution is an alkali; it neutralizes acids to form salts and water only

$$\text{NH}_3\text{(aq)} + \text{HNO}_3\text{(aq)} \longrightarrow \text{NH}_4\text{NO}_3\text{(aq)} + \text{H}_2\text{O}$$

The above reaction is important because it forms a fertilizer ammonium nitrate.

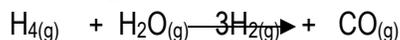
4. Ammonium solution reacts with cations of Al^{3+} , Cu^{2+} , Fe^{3+} , Ca^{2+} e.t.c. to form precipitates that are used to identify cations.

The ammonia gas turns damp red litmus blue.

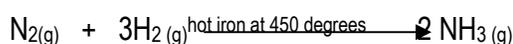
Industrial Production of ammonia.

In industry, ammonia is made from hydrogen gas and nitrogen gas in a process called the Haber process.

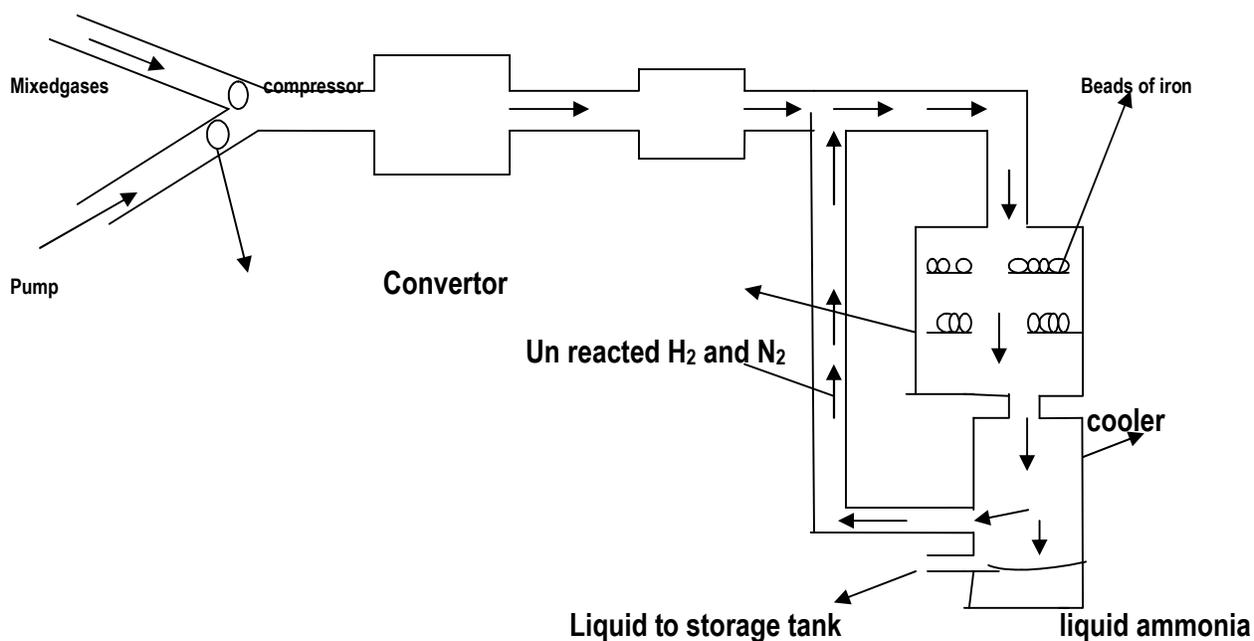
- Nitrogen is obtained from the atmosphere by burning hydrogen in air to remove oxygen and remain with nitrogen.
- Hydrogen is obtained from the reaction between methane and steam in a process called steam reforming.



The two gases; hydrogen and nitrogen are mixed and the mixture cleaned to remove impurities. The mixture is then compressed to a pressure of 200 atmospheres. From the compressor the mixture is pumped into a converter. The converter is a round tank containing beads of hot iron at a temperature of 450°C. At this temperature and in the presence of hot iron, hydrogen reacts with nitrogen to form ammonia according to the equation below:-

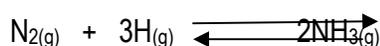


A mixture of ammonia and the un reacted gases of nitrogen and hydrogen leave the converter into the cooling tank. At negative 33°C ammonia liquefies and the un reacted gases are pumped back into the converter. The liquid ammonia is drained into containers and stored. The diagram below shows the process in ammonia production by **Haber process**.



Le-chatelier's principle:

The reaction in the Haber process is exothermic and the mixture leaving the reaction contains 15% ammonia. The 15% of the ammonia produced is not enough and the reason for this is the reversible nature of the reaction that is once the ammonia is made from nitrogen and hydrogen, it decomposes back into hydrogen and nitrogen; according to the equation.



Soon there comes a point when the rate of the forward reaction to produce ammonia becomes equal to the rate of the reverse reaction i.e. the rate at which ammonia decomposes to produce nitrogen and hydrogen. This point or situation is called Chemical equilibrium and because the process continues to happen the situation is called dynamic equilibrium.

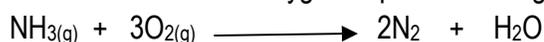
To continue producing ammonia the equilibrium must be shifted towards the ammonia. This is done by:-

- Increasing pressure
- Increasing the concentration of the reactants.
- Removing the product i.e. ammonia as soon as it is formed.
- An optimum temperature of 450°C is selected
- A catalyst of iron is used to lower the activation energy

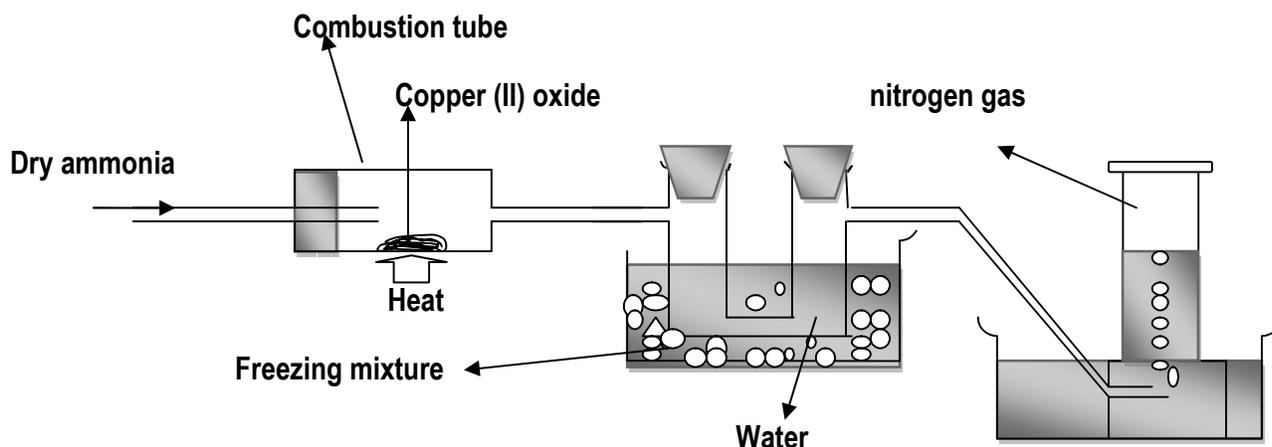
Whenever one or two of the factors stated above are imposed on any chemical system in equilibrium, the chemical system reacts in such a way as to oppose the stress imposed on it. This is the Le-chatelier's principle.

Other chemical properties of ammonia.

- a) Ammonia reacts with oxygen to produce nitrogen.



- b) Ammonia is used to reduce copper (II) Oxide to copper metal. The diagram below shows the reduction of copper oxide. At the end of the reaction copper metal remains in the combustion tube



Uses of Ammonia

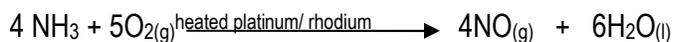
- (1) Ammonia solution is used in laundry work. It removes temporary hardness of water by precipitating the Calcium hydrogen carbonate as Calcium Carbonate.
- (2) Ammonia solution dissolves out oxides left by evaporation of precipitation from underclothing.
- (3) Ammonia is used to make nitric acid
- (4) Ammonia is used to make fertilizers such as ammonia sulphate, ammonium nitrate and super phosphate of ammonia.

MANUFACTURE OF NITRIC ACIDS

One of the largest uses of ammonia is in the production of nitric acid. The Raw materials for manufacturing nitric acid are; ammonia, air and water. The process has 3 stages.

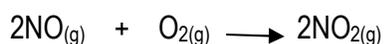
Stage 1:

A mixture of air and ammonia is heated to 230°C and passed through a metal gauze made of a mixture of platinum and rhodium. The reaction produces nitrogen monoxide and water.



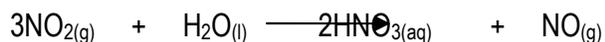
Stage 2:

The colourless nitrogen monoxide gas produced in the first stage is then reacted with more oxygen to produce nitrogen dioxide



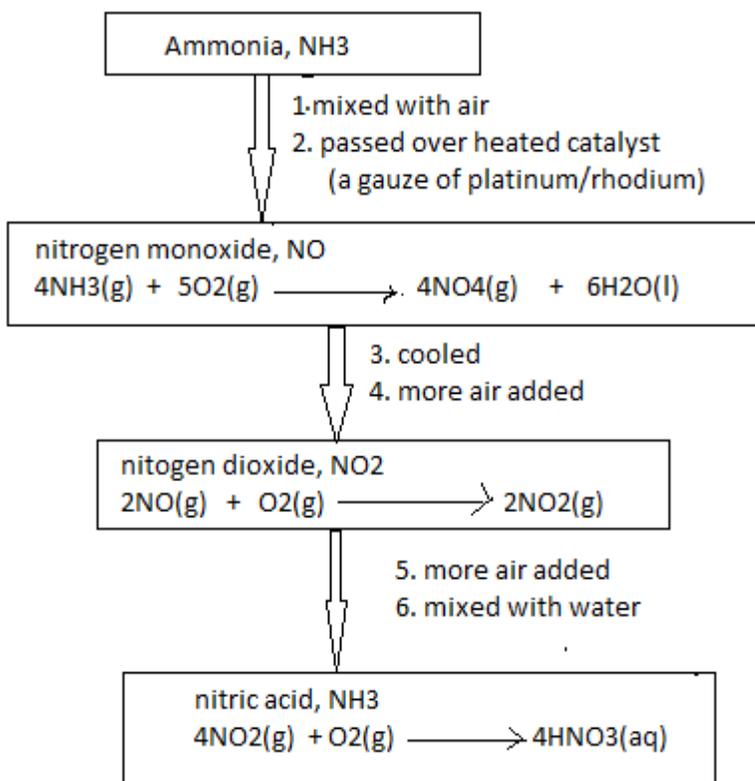
Stage 3:

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



The overall result is that ammonia is oxidized to nitric acid.

The diagram below shows the stages



Precaution:

Chemical engineers should take necessary precautions that:

- (1) Nitrogen monoxide and nitrogen dioxide does not escape from the plant into the atmosphere because they can cause acid rain hence pollution.
- (2) Nitric acid should not go down the drain as it may end up into rivers killing fish and other aquatic life.

Physical Properties of nitric acid:

Just like any other inorganic or mineral acid; nitric acid has the following properties.

- (i) It is soluble in water with a ph of less than 7.
- (ii) It has sour sharp taste.
- (iii) It is colourless
- (iv) It turns blue litmus paper red
- (v) It is very corrosive

Chemical properties of nitric acid:

Just like any other inorganic acid it has the following chemical properties.

- (i) It reacts with metals to liberate hydrogen gas.

- $$\text{Mg}_{(s)} + \text{HNO}_{3(aq)} \longrightarrow \text{Mg}(\text{NO}_3)_2 + \text{H}_2_{(g)}$$
 (ii) It reacts with carbonates to liberate carbon dioxide;

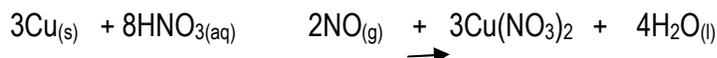
$$\text{CaCO}_3 + 2\text{HNO}_3 \longrightarrow \text{Ca}(\text{NO}_3)_2 + \text{CO}_2 + \text{H}_2\text{O}$$
 (iii) It neutralizes bases and alkalis;
 (a)
$$\text{NaOH}_{(aq)} + \text{HNO}_{3(aq)} \longrightarrow \text{H}_2\text{O} + \text{NaNO}_3(aq)$$

 (b)
$$\text{HNO}_3(aq) + \text{CaO} \longrightarrow \text{H}_2\text{O}(l) + \text{Ca}(\text{NO}_3)_2$$

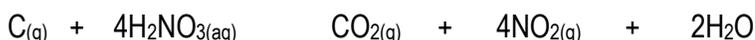
Oxidising properties of nitric acid:

In addition to the above chemical properties nitric acid is a powerful oxidizing agent:

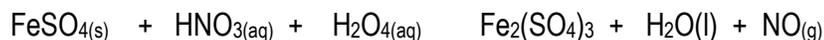
- i. It oxidizes Copper to produce nitrogen monoxide Copper (II) Nitrate and Water.



- ii. It oxidizes Carbon to Carbon dioxide



- iii. It oxidizes salts iron (II) to iron in the presence of sulphuric acid.



Uses of Nitric Acid

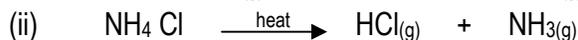
- (i) It is used to make fertilizer e.g. ammonium Nitrate.
- (ii) Nitric acid is used to make explosives such as Nitrotoluene.
- (iii). It is used to make drugs
- (iv). used in the manufacturing of nylon and terylene

Undesirable effect of Nitrogenous fertilizers and acidic effluents

- (1) If too much fertilizer is applied to the land, rain water washes it into rivers and streams where it causes Eutrophication.
- (2) If too much fertilizer is applied to the land, the land becomes acidic and this kills micro-organisms in the soil.
- (3) The overgrowth of aquatic plants due to the presence of nitrate may also cause the growth of algae. Some algae are poisonous to fish and human beings, people swimming in these waters or drinking the water may get eye irritation, skin rashes, vomiting and diarrhea.
- (4) Excess nitrate in the river may get into drinking water increasing the risk of blue – baby syndrome in pregnant women.
- (5) Acidic liquids from factories and manufacturing industries as well as domestic water wastes are known as effluents. If these effluents are allowed to run into the rivers and streams they can also cause Eutrophication and skin diseases in fish

thermal dissociation of ammonium compounds:

Thermal dissociation or decomposition reaction is brought about by the application of heat on the compounds. When ammonium compounds are heated they normally break down into smaller units:

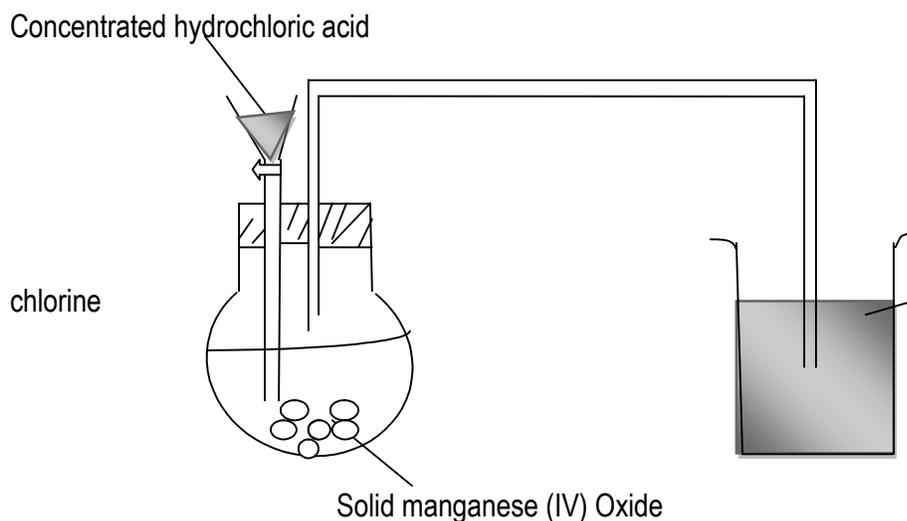


CHLORINE

Chlorine as an element is found in group 7 on the Periodic Table. It is a member of the halogen family. At room temperature and pressure it exists as a diatomic gas molecule. Chlorine is a very reactive element and it is never found as a free element. It occurs mainly in nature as sodium chloride or rock salt

Laboratory preparation of chlorine;

Chlorine is made in the laboratory by oxidizing concentrated hydrochloric acid with manganese iv oxide as shown in the diagram below:-

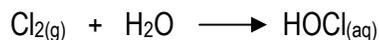


PHYSICAL PROPERTIES OF CHLORINE

- (i) It is denser than air.
- (ii) It has a greenish yellow colour
- (iii) It is soluble in water forming an acidic solution
- (iv) It bleaches litmus paper
- (v) It has evil smell and very poisonous.

CHEMICAL PROPERTIES OF CHLORINE

- (i) Chlorine gas reacts with water to form hypochlorous acid.



(ii) chlorine burns in hydrogen to hydrogen chloride;



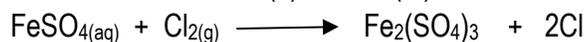
(iii) Chlorine reacts with many metals to form metal chlorides e.g.



(iv) Chlorine displaces any halogen below it in the group from its aqueous solution e.g.



(v) Chlorine oxidizes iron (II) to iron (III) in the iron salt.



USES OF CHLORINE

- (1) Used to make plastics e.g. (PVC) polyvinylchloride.
- (2) Used to make hydrochloric acid since it is soluble in water.
- (3) It is used to make disinfectants and insecticides because it is toxic and poisonous hence it kills insects and germs.
- (4) It is used to make bleaching solvents such as Jik.
- (5) Chlorine is used to sterilize drinking water or domestic water and swimming pool water.

TEST FOR CHLORINE

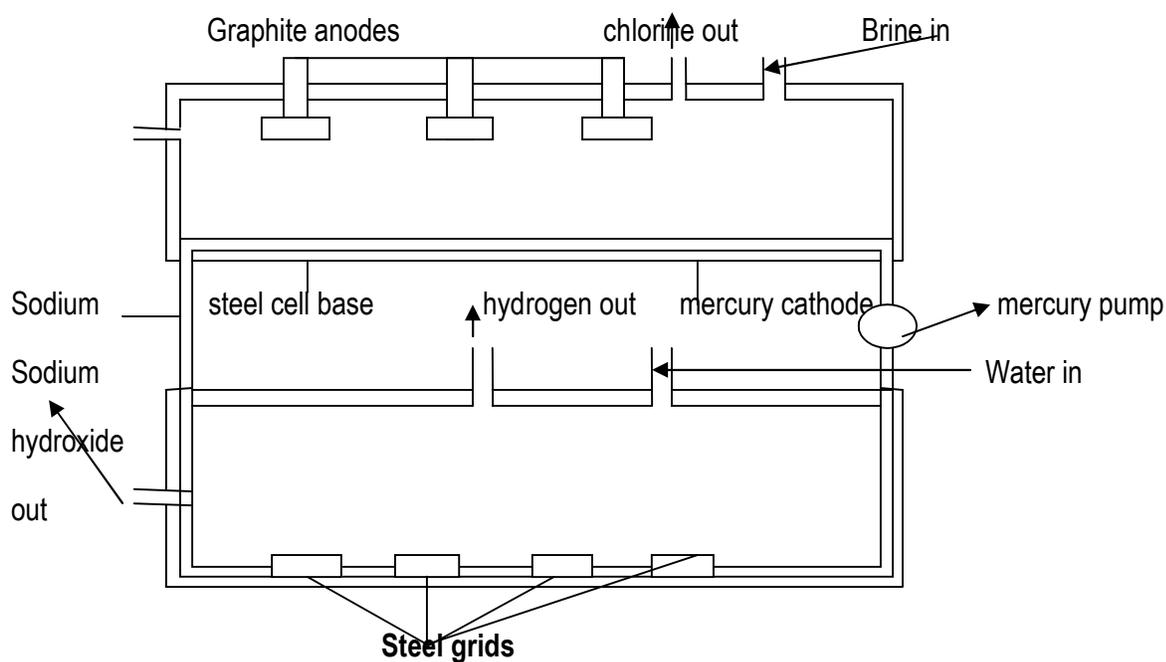
Chlorine bleaches litmus paper. If the litmus paper is blue it first turns red then get bleached. It first turns red because chlorine first reacts with water on the litmus paper to form hypochlorous acid and hydrochloric acid (HCl)



The hydrochloric acid turns the blue litmus paper red first, after which the hypochlorous acid (HOCl) reacts with the dye on the litmus paper to bleach it (to turn whitish).

INDUSTRIAL PREPARATION/PRODUCTION OF CHLORINE.

Chlorine is produced commercially by the electrolysis of concentrated sodium chloride solution (NaCl) (Brine). Chlorine is released at the anode while the other product which is (NaOH) sodium hydroxide is released at the cathode. The two products must be kept apart because they react.



Reaction of chlorine with nonmetals;

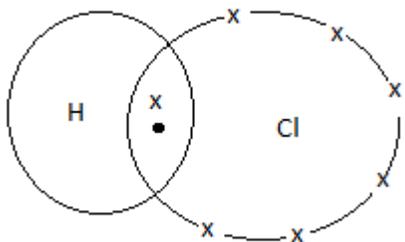
1. Chlorine reacts with hydrogen to form hydrogen chloride gas.
2. Chlorine reacts with Oxygen to form dichloroxide

$$2\text{Cl}_2 + \text{O}_2 \longrightarrow 2\text{Cl}_2\text{O}$$
3. Chlorine reacts with phosphorous to form phosphochloride

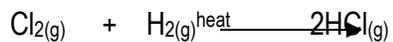
$$2\text{P} + 3\text{Cl}_2 \longrightarrow 2\text{PCl}_3$$
4. Chlorine reacts with sulphur dioxide to form sulfuryl chloride
5. $\text{SO}_2 + \text{Cl}_2 \longrightarrow \text{SO}_2\text{Cl}_2$

Hydrogen chloride gas;

Hydrogen chloride is a covalent compound



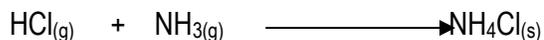
Hydrogen chloride is made in industry by bringing hydrogen in chlorine gas as shown in the chemical equation below:-



The gas has the following properties:

1. It is denser than air.

2. Very soluble in water to form hydrochloric acid
3. It has a choking smell and it irritates the eyes and the respiratory system.
4. As a dry gas; it has no effect on the dry blue litmus paper.
5. Hydrogen chloride gas reacts with ammonia to form white precipitate of ammonium chloride.



Hydrochloric acid;

Hydrochloric acid is made in industry by dissolving hydrogen chloride gas in water. Hydrochloric acid is a typical acid and has all the properties of an acid.

SULPHUR

Sulphur is a non-metallic element. As an element it is found in group 6 in the Periodic Table. At room temperature and pressure it exists as a yellow solid of the chemical formula S_8 .

In nature sulphur is found in many forms through out the world as a yellow substance usually 200m below the surface of the earth. Sulphur is also found in metal ores as copper pyrites (CuFeS_2), zinc blende (ZnS) and galena (PbS)

Allotropes of sulphur;

Allotropy is the existence of an element in many forms in the same state. Sulphur is one of the few nonmetals that exist as allotropes. The main allotropes of sulphur are rhombic and monoclinic. Allotropes differ from each other by the way in which molecules pack together i.e. in one allotrope the molecules may pack more tightly together and in the other they may pack together loosely.

Physical properties of sulphur

1. It is a yellow crystalline solid at room temperature and pressure
2. It has a faint smell when not pure
3. It has no taste
4. It boils at 444.75°C

Chemical properties

1. It exhibits oxidation numbers; -2, 0, +2, +4, and +6. Oxidation number -2 is seen in sulphides e.g. H_2S
2. It burns in air to form sulphur dioxide
3. It combines with many elements to form sulphides

Uses of sulphur

1. 90% of sulphur is used to make sulphuric acid
2. Sulphur is used to toughen rubber for making tyres; this is called vulcanization.
3. Sulphur is used to make drugs such as sulphonamides which are used to manage bacterial infection

- Sulphur is used to manufacture pesticides and fungicides
- Sulphur is used to make match sticks head
- Sulphur is added to cement to make a sulphur concrete which is not attacked by acids.

Sulphur dioxide;

Sulphur dioxide is formed when fossil fuels such as coal burn in oxygen or when substances containing sulphur such as metal ores are heated in oxygen

Physical properties of sulphur dioxide:

- It is a colourless gas.
- It has a strong choking smell
- It is denser than air.
- It is an acidic oxide.
- It is soluble in water forming sulphurous acid (H_2SO_3)
- It turns damp blue litmus paper red.

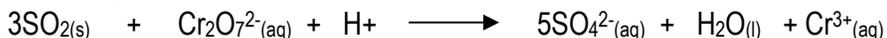
Chemical Properties of Sulphur dioxide

Sulphur dioxide is one of the major pollutant gases and it is the gas which is responsible for much of the acid rain. It has the following chemical properties.

- It dissolves in water to form sulphurous acid
- $\text{SO}_{2(g)} + \text{H}_2\text{O}_{(l)} \longrightarrow \text{H}_2\text{SO}_{3(aq)}$
- It reacts with hydrogen sulphide to form sulphur and water
 $\text{SO}_{2(g)} + \text{H}_2\text{S} \xrightarrow[\text{Calcious/iron oxide}]{300\text{degrees}} \text{H}_2\text{O}_{(l)} + 3\text{S}_{(s)}$
- Sulphur dioxide reacts with sodium hydroxide to produce sodium sulphate and water
 $2 \text{NaOH}_{(aq)} + \text{SO}_{2(g)} \longrightarrow \text{Na}_2\text{SO}_{3(s)} + \text{H}_2\text{O}$

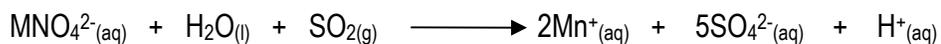
Sulphur Dioxide as a reducing agent.

- Sulphur dioxide acts as a bleach when it is damp or in solution. Some coloured materials lose colour when they lose oxygen. Remember that loss of oxygen by a substance is reduction. When they lose oxygen they get bleached.
- Sulphur dioxide reduces acidified potassium dichromate solution ($\text{Cr}_2\text{O}_7^{2-}$). The solution contains the orange dichromate ion ($\text{Cr}_2\text{O}_7^{2-}$). Sulphur dioxide reduces it to the green chromium ion (Cr^{3+}) according to the equation.



The above reaction is used as the identity test for sulphur dioxide gas.

- Sulphur dioxide reduces purple potassium permanganate to a colourless substance. When potassium permanganate reacts with sulphur dioxide it gets decolourised. This is because potassium permanganate gets reduced by sulphur dioxide.



Industrial Preparation of Sulphuric Acid:

The major use of sulphur is in the production of sulphuric acid. Sulphuric acid is probably the most important industrial chemical. The process by which sulphuric acid is produced is known as the contact process.

The raw materials needed for the production of sulphuric acid are: Sulphur dioxide, air and water. Sulphur dioxide is got from the extraction of metals such as copper, zinc and lead.

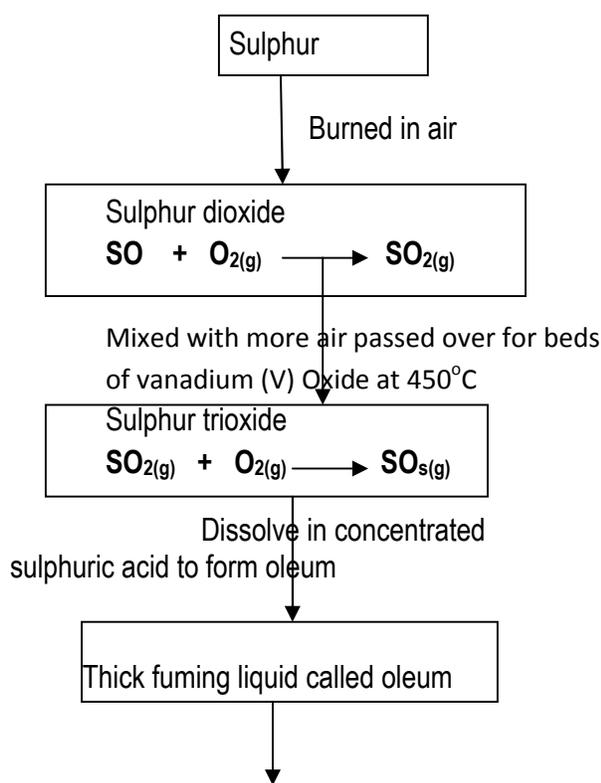
Sulphur is burned in air to form sulphur dioxide. Sulphur dioxide is mixed with more air and heated to 450°C in the beds of a catalyst called vanadium oxide to form sulphur trioxide. This reaction is reversible hence the mixture is passed on the vanadium catalyst to let the gases react again and to increase the yield. The sulphur trioxide is removed according to Le-chatelier's principle.

Below 400°C the catalyst fails to work and above 450°C the reaction becomes exothermic and reversible this affects the yield. (the yield of sulphur trioxide drops)

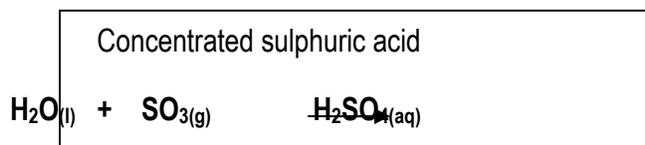
To keep the temperature below 450°C heat must be removed from the catalyst by passing cold water pipes over the hot catalyst. The water is converted to steam and used to generate electricity or sold to nearby factories for heating.

The sulphur trioxide is dissolved in concentrated sulphuric acid rather than water. If sulphur trioxide is dissolved in water, a thick mist of sulphuric acid is formed. The thick mist is a serious pollution hazards.

The contact process



Mixed with water carefully

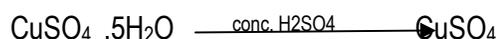


Chemical Properties of Sulphuric acid;

- (i) It turns blue litmus paper red.
- (ii) It reacts with metals to form hydrogen gas and a metal sulphate.
- (iii) It reacts with metal oxide to form water and metal only.
- (iv) It reacts with carbonates to give water, carbon dioxide and metal sulphate.

Properties of Concentrated Sulphuric acid.

- (i) It is colorless, oily liquid.
- (ii) It is non-volatile
- (iii) It is a dehydrating agent i.e. it removes water from sugar, wood and paper because there are made of hydrogen, carbon and oxygen.
The acid removes hydrogen and oxygen as water.
$$\text{C}_6\text{H}_{12}\text{O}_6 \longrightarrow 6\text{C}_{(s)} + 6\text{H}_2\text{O}_{(l)}$$
- (iv) It will also dehydrate flesh by removing hydrogen and oxygen as water.
- (v) It is a hygroscopic substance i.e. it removes water from a substance and because of this it is used as a drying agent. E.g. sulphuric acid removes water from hydrated Copper (II) Sulphate to give the anhydrous Copper (II) sulphate.



- (vi) Sulphuric acid is an oxidizing agent as it oxidizes both metals and non-metals. Examples below illustrates the oxidizing properties of sulphuric acid;

- a. $\text{Fe}_{(s)} + \text{H}_2\text{SO}_4 \longrightarrow \text{FeSO}_4 + \text{H}_2$
- b. $\text{C}_{(s)} + \text{H}_2\text{SO}_4 \longrightarrow \text{CO}_2 + \text{SO}_2 + \text{H}_2\text{O}$
- c. $\text{S}_{(s)} + \text{H}_2\text{SO}_4 \longrightarrow \text{SO}_2 + \text{H}_2\text{O}$

Uses of sulphuric acid:

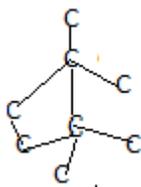
- (i) It is used to make fertilizers e.g. ammonium sulphate
- (ii) It is used for making paints and dyes.
- (iii) It is used in the manufacturing of plastics and fibers.
- (iv) It is used in the manufacturing of nylon as a dehydrating agent.
- (v) It is used as electrolyte in car batteries.
- (vi) It is used to make and detergents.

CARBON AND CARBONATES:

A small amount of carbon occurs as the free element in the earth's crust. It occurs as diamond, graphite, amorphous carbon and fullerenes. These are the allotropes of Carbon.

Diamonds:

Diamond is the purest form of carbon. In diamond each carbon atom is covalently bonded to 4 other carbon atoms. This bonding produces a giant structure called diamond. In diamond, there are no free electrons since all the 4 electrons in the valence shell are used in bonding. Diamond therefore does not conduct electricity. The diagram below shows the simplified diagram for the structure of diamond



Properties of Diamond:

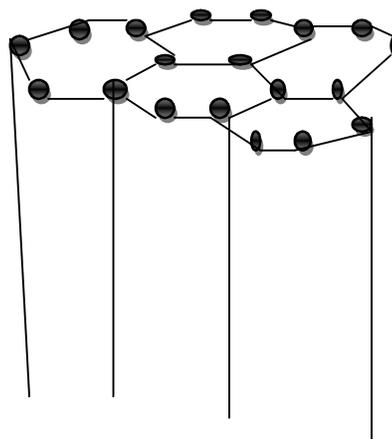
- (i) It is the hardest known substance to date because of the short covalent bond.
- (ii) It has a density of 3.5g/cm^3
- (iii) It is colourless
- (iv) It is transparent and crystal
- (v) It sparks when cut and polished.
- (vi) Non-volatile.

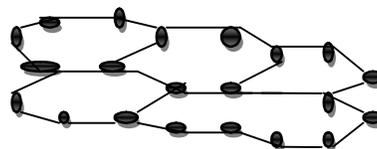
Uses of Diamond:

- (i) Since it sparks it is used as jewellery or ornaments.
- (ii) It is used to make edges of cutting saws because it is very hard.
- (iii) It is used to make the tips of drilling tools because it is hard and has a high melting point.

Graphite:

This is an impure form of Carbon. In graphite each carbon atom is covalently bonded to three other carbon atoms. This bonding leaves one electron in the valency shell unused. This unused electron is mobile and is responsible for the electrical properties of graphite, therefore, graphite conducts electricity. The bonding in graphite produces a hexagonal structure of layers of graphite. These layers can slide over each other. The diagram below shows the structure of graphite





Properties of Graphite

- (i) It conducts electricity because it has a free electron
- (ii) It is dark grey in colour
- (iii) It is greasy and slippery
- (iv) It is soft
- (v) It is a density 2.25g/cm^3

Uses of Graphite

- (i) It is used as a lubricant because it is soft/greasy/soapy.
- (ii) It is used in pencil lead because it is grey and shades some atoms when rubbed against a surface.
- (iii) It is used in electrical connections such as carbon brushes because it conducts electricity

Comparing Graphite, Diamond, Amorphous and Fullerenes

Property	Diamond	Graphite	Amorphous	Fullerenes
Appearance	Colourless, transparency, sparking.	Grey-black, opaque, shiny.	Black, Opaque, dull	Brown – black powder.
Density	3.5g/cm^3	2.3g/cm^3	1.5g/cm^3 in	1.7g/cm^3
Hardness	Hardest substance known	Soft greasy and soapy	Soft and Brittle	Fine 500+ powder
Combustion in air	Does not burn easily at 900°C only above.	Turns readily at 700°C	Burns readily at 500°C	
Product on complete combustion	CO_2	CO_2	CO_2	CO_2
Electro-conductivity	Poor conductor	Good conductor	Fair conductor	Fair conductor
Uses	Making Jewellery, glass cutters, polishers, drill bits and studded saws	Lubricant, making pencil lead and electrical connection	Charcoal for absorbing gases and dyes	For making super conductor in electrical connections.

Carbon dioxide:

Carbon dioxide is formed when carbon containing substances such as fuel burns in plentiful supply of air according to the equation.



Carbon dioxide is also formed in respiration. When sugars are oxidized in the living tissue according to the equation.



In the atmosphere carbon dioxide occupies a very small volume of about 0.03%

Properties of Carbon dioxide:

- (i) It is a colourless gas.
- (ii) It has no smell
- (iii) It is $2\frac{1}{2}$ times denser than air.
- (iv) It is slightly soluble in water
- (v) It doesn't support combustion
- (vi) It sublimates at -78°C .
- (vii) It turns lime water milky.

Uses of Carbon dioxide:

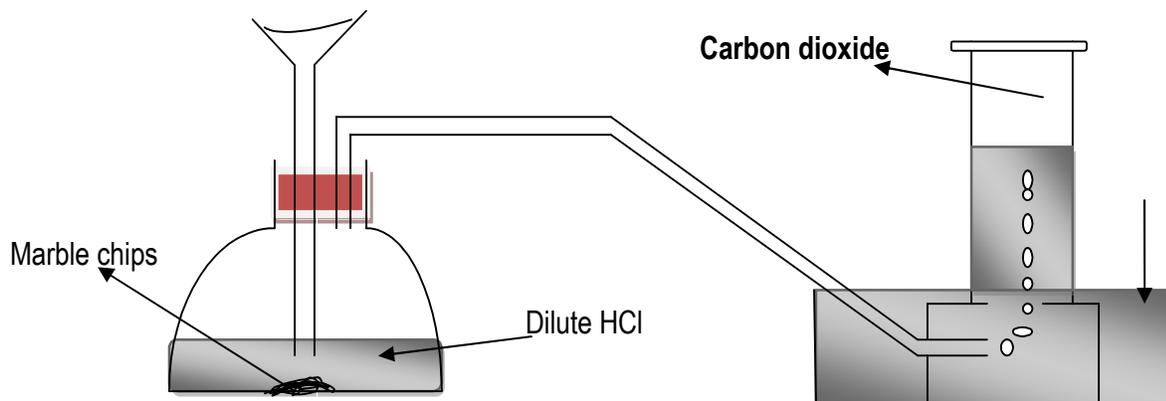
- (i) It is used in fire extinguishers because it does not support burning and is denser than air.
- (ii) It is used as a coolant in refrigerators to keep foods fresh as they being transported by road because it sublimates. It is called dry ice.
- (iii) It is put in drinks such as coca-cola to make them fizzy because it is slightly soluble in water and being acidic it gives a sharp pleasant taste.

Identity test for Carbon dioxide:

When Carbon dioxide is bubbled into clear lime water; the lime water turns milky because of fine white precipitate of Calcium carbonate that forms according to the equation.



Laboratory preparation of Carbon dioxide:

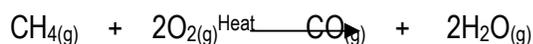


Any carbonate can react with a dilute acid to liberate carbon dioxide. Carbon dioxide is collected by either the downward displacement of water because it is slightly soluble in water or by downward delivery because it is denser than air. Downward delivery is also called upward displacement of air.

Formation of Carbon monoxide:

Carbon monoxide is formed when there is incomplete combustion of fuels; such as wood, charcoal and any hydrocarbon.

Incomplete combustion occurs when there is insufficient air supplied to the burning fuel for example. If methane is burned in insufficient air supply the products are carbon monoxide and water according to the equation.

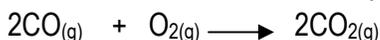


Physical Properties of Carbon monoxide:

- (i) It is a colourless gas
- (ii) It has no smell
- (iii) It is insoluble in water
- (iv) It has the same density as air
- (v) It is very poisonous and toxic.

Chemical properties

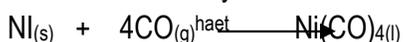
- (i) Carbon monoxide burns with a pale blue flame forming carbon dioxide.



- (ii) Carbon monoxide is a good reducing agent; it reduces metal oxides



- (iii) Carbon monoxide reacts with certain metals to form carbonyls. For example it reacts with nickel to form nickel carbonyl.



Poisonous nature of Carbon monoxide:

Carbon monoxide is a very poisonous gas. It readily combines with haemoglobin in red blood cells in blood, and prevents oxygen from combining with haemoglobin. This causes oxygen starvation in the body tissues and eventually causes death by suffocation. 0.1% of carbon monoxide in the air is enough to cause death

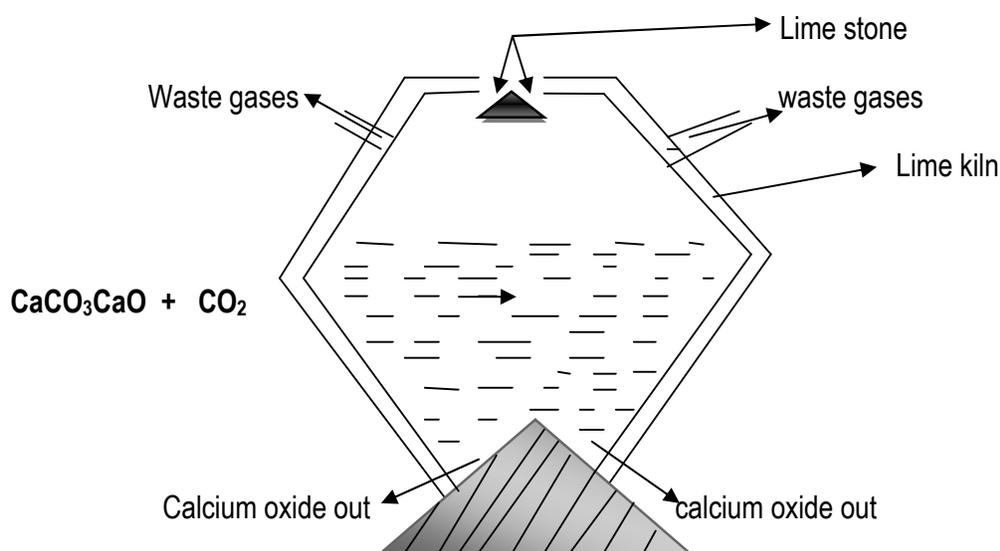
It is therefore, not advisable to keep a brazier in a room when windows are closed because the carbon monoxide from the brazier may accumulate in the room to dangerous levels. Young children should not play in garages as they contain carbon monoxide from cars' exhaust pipes.

Uses of Carbon monoxide

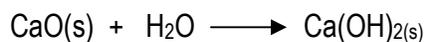
- (i) It is used as a fuel to heat blast furnaces.
- (ii) It is used as a reducing agent to extract iron from its ores

Manufacture of Lime:

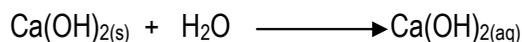
The most common carbonate is calcium carbonate. It is found as Chalk, marble or limestone. All the three have the same chemical composition. They only differ in their hardness. Chalk is very soft, limestone is hard and marble is the hardest. Lime is made by heating limestone. Limestone is mixed with coke and heated in lime-kiln. The product is quick lime. This quick lime is impure as it is mixed with ash. Pure form of quick lime is made by heating the kiln with producer gas. The equation for the reaction is



If a small amount of water, sufficient to moisten the lime is added, the lime glows, swells and forms slaked lime or solid calcium hydroxide according to equation.



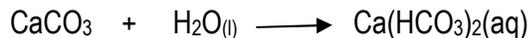
When more water is added and stirred vigorously slaked lime dissolves to form calcium hydroxide solution (lime water) according to the equation;



When carbon dioxide is bubbled into clear lime water, the lime water turns milky due to the formation of solid calcium carbonate particles according to the equation;



When more carbon dioxide is bubbled into lime water the milky solution becomes colourless due to the formation of a soluble substance called calcium hydrogen carbonate; according to the equation.



Uses of Lime and Slaked lime;

- (i) Lime or quick lime is added to acidic soils to neutralize the soils. However, too much addition of lime makes the soil alkaline
- (ii) Slaked lime is used to neutralize acidic spillages on roads and acidic effluents from industries. Slaked lime has the advantage that it is slightly soluble in water hence even when added in excess it does not significantly change the P^{H} of the soil.

Properties of Carbonates;

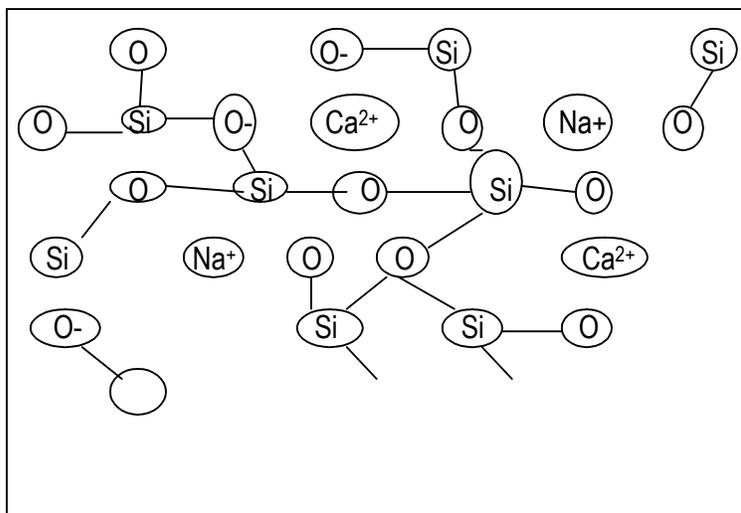
1. They are insoluble in water, except carbonates of potassium, Sodium and ammonium
2. Most carbonates upon heating strongly they decompose into an oxide and carbon dioxide, except the carbonates of potassium, Sodium and ammonium.
3. Carbonates react with acids to form carbon dioxide, water and a salt.

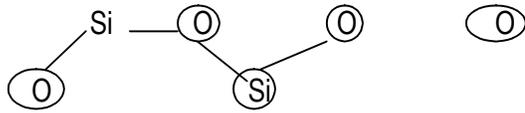
Uses of lime stone;

1. Limestone is used to manufacture lime
2. Limestone is used to manufacture glass
3. Limestone is used to manufacture cement

Manufacture of glass;

Glasses are irregular giant structures held together by strong covalent bonds. Glass is made by heating silicon (iv) oxide with calcium carbonate and other substances until a thick viscous liquid is formed. As this liquid cools, the atoms present can not move freely enough to return to their original arrangement in pure silicon(iv) oxide structure. The diagram below shows the arrangement of the particles in glass





Coloured glasses are made by adding coloured ions of transition metals.

Manufacture of cement;

Cement is a finely ground inorganic material which when mixed with water forms a paste which sets and hardens by means of dehydration reactions and retains its strength and suitability even under water.

PORTLAND CEMENT:

This is a grey mixture of argillaceous (clay-like) calcareous containing limestone mixed with gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$) sintered and pulverized into a fine product. Portland is a name of a place where cement was discovered.

Types of Portland Cement:

- (i) Ordinary Portland Cement
- (ii) Rapid hardening cement – high heat of hydration.
- (iii) Moderate heat-cement-low heat of hydration.

Rapid Hardening Cement:

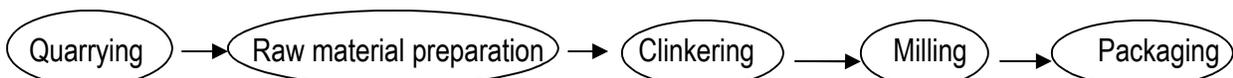
This type of cement is used in Pre-Cast Concrete, pipes and tiles. It is finely ground so that it can hydrate more quickly and it has more gypsum than other cements.

Moderate Heat Cement:

It is used in the construction of dams and bridges. It is preferred for this kind of use as opposed to ordinary cement because of the large heat produced which creates an even expansion leading to cracking when a large volume of concrete is used.

The Manufacture Process

Cement is made by heating raw materials rich in oxides of aluminium, iron, silicon and Calcium to temperatures of around $1\ 200^\circ\text{C}$ - $1\ 400^\circ\text{C}$. Chemical reactions that occur within the partially molten mass result in the formation of three main cement materials i.e. ordinary cement, rapid hardening cement and the moderate heat cement.



Stage 1; Quarrying:

This involves getting raw materials i.e. 80% Calcium Carbonate (CaCO₃) and clay 20%. Clay is a source of Silica, Alumina and Iron (III) Oxide. These are stored separately. The lime and silica provide the main strength of the cement while the iron reduces the reaction temperature and gives the cement its grey colour.

Stage2: Raw Materials Preparation:

The material is prepared based on the process to be utilized during clinkering and the processes are dry process and wet process in clinkering.

Stage 3: Clinkering:

The finely ground material is dried, heated (to allow the sintering reaction to occur and then cool again). The powder is preheated to 70 – 80°C where the moisture evaporates and up to 20% decarbonation occurs.



At the Kiln:

The kiln is heated by injecting pulverized coal dust into the discharge and where it spontaneously ignited due to very high temperatures.

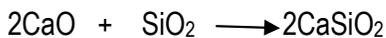
Zone 1: Of the kiln 800 – 1100°C between 0 - 35minutes.

- Here there is decarbonation.



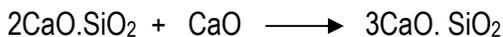
Zone 2: 1100°C to 1300°C (35 – 40 minutes)

- Exothermic reaction and formation of secondary silicate phases occurs as follows:-



Zone 3: 1400 - 1450°C (40 - 50minutes)

- Reactions occur within the melt to form ternary silicates and tetra-calcium alumino ferrate.



Zone 4: 1300 - 1000°C → (50 - 60 minutes)

- Here we have cooling and crystallization of the various mineral species formed in the kiln.

The Cooler:

It is situated immediately after the kiln and cools the clinker from 1 000°C to 150°C. This is achieved by forcing air through a bed of clinker via perforated plates at the base of the cooler.

Step 4: Cement Milling:

The clinker can react with water and harden but then it is composed of 1-3cm particles in diameter which can not be used because they are too coarse to be used. The clinker is mixed with gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$) which is added as a set retarder and ground in large tube mills for about 30 minutes. The particles size is measured by laser diffraction particle analysis and the particle of materials entering the mills are adjusted accordingly. After milling the cement is packed and ready for marketing

Effect on Environment:

I. Dust emission:

The process generates large quantities of dust which must be prevented from escaping into the atmosphere. Dust is emitted from kilns, mills, transport equipment, quarry process. The transport equipment should be enclosed and the air from the kiln and mills should be treated with an electro-static precipitator.

Limits for dust emission vary from 50 to 450 mg/m^3 , based on environmental policies in place.

II. Carbon dioxide:

This is produced during the calcinations phase and burning of fossil fuels. The solution to limiting the amount of CO_2 is to use alternative source of energy.

III. Quarry and Plant water run off:

Water runoff is usually trapped in wetland areas where it can be treated in a controlled manner. Within the factory runoffs can be contaminated by oils and lubricants and should therefore be monitored before being released into the environment.

IV. Chrome bricks:

Kiln bricks used to be made out of hexavalent chrome which is carcinogenic and causes dermatitis, the use of the materials have been replaced with environmentally friendly Magnesium – spined bricks.

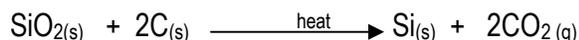
SILICON AND ITS COMPOUNDS

The element silicon belongs to the same family as carbon. It is the most common element in the earth's crust after oxygen. It usually occurs combined with oxygen as:

1. Silicon (iv) oxide, this is also called silicon dioxide and silica. This has a formula SiO_2 and one example of silica is quartz. Sand is an impure form of quartz
2. In complex silicates. Silicates contain ions such as SiO_4^{4-} and $\text{Si}^4\text{O}_{11}^{6-}$. One example of silicates is Mica. Mica is a mineral found in clay and granite.

Extraction of silicon;

Silicon is got by heating sand and coke in an electric furnace as shown in the equation below



In the above equation silicon (iv) oxide is reduced to silicon and coke is the reducing agent.

Physical properties of silicon;

1. At room temperature it is a hard grey solid.
2. It has a high melting point of 1410°C
3. It conducts electricity but not as well as a metal. So it is called a semiconductor.

Chemical properties of silicon;

1. It is a metalloid that means it behaves like a non metal and partly like alkali metals.
2. It combines with organic carbon compounds to form silicones.

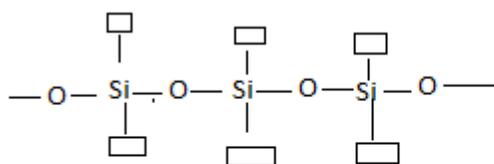
Uses of silicon and silicon;

1. Silicon is used to make computer chips because it is a semiconductor
2. Silicon is used to make solar cells because it is a semiconductor

For the above uses; silicon must be **very** pure

Silicones:

Silicon combines with organic carbon compounds to form a large range of polymers called silicones. In silicones, the backbone is a long chain of silicon and oxygen atoms. Carbon atoms or chains of carbon atoms are bonded to the silicon atoms. The diagram below shows how the bonding is done



where represents carbon atoms or carbon chains

Properties of silicones;

1. They may be oils, waxes or plastics depending on the length of the chains and the carbon groups attached. Plastics may be hard, or flexible, or stretchy like rubber
2. They are unreactive and water proof.
3. They do not break down on heating. They are more stable than carbon based polymers i.e. they do not melt or burn so easily. This is because the Si-O bond is stronger than the C-C bond in carbon based polymers. The product of combustion of silicon is Silicon (iv) oxide which is a solid where as

the combustion product of carbon based polymers is mostly carbon dioxide and water which are gases. The solid covers the silicon and cuts out oxygen hence burning stops.

Uses of silicones;

Because of their properties, silicones have a wide range of uses;

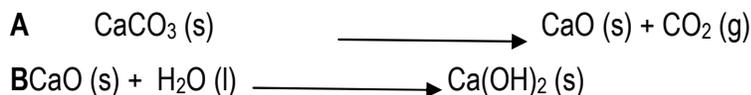
1. Used in vanishes, paints, polishes and adhesives.
2. Used as coating for dishes
3. Used to waterproof material for rain coats
4. Used as a scratch resistant coating, for car windshields
5. Used as fire proof foams and sealants in buildings
6. Used as waterproof sealants around baths and sinks
7. Used as ingredients in cosmetics, shaving creams and suntan lotions.

QUESTION1: non metal

Limestone is one of the forms of calcium carbonate.

- (a) Name two other forms of calcium carbonate. [2]
- (b) Limestone is used to make calcium oxide, which is also used to make calcium

hydroxide according to the equation below:



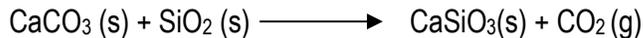
Reaction **A** absorbs heat from the surrounding where as reaction **B** gives out heat to the surrounding.

What name is given to a reaction which.....

- (i) Gives out heat? [1]
- (ii) Absorbs heat? [1]
- (c) What is the common name for
- (i) Calcium oxide? [1]
- (ii) Solid calcium hydroxide? [1]
- (d) When calcium hydroxide is dissolved in water, lime water is produced.
- (i) Describe and explain what happens to limewater when you blow through it for a few seconds using a straw. [2]
- (ii) Write an equation for the reaction that occurs. [1]
- (iii) Describe and explain what happens to limewater when you blow through it for a long period of time. [2]

(iv) Write an equation to show the reaction that occurs. [1]

(e) Limestone reacts with sand (silicon dioxide) to produce calcium silicate:



Sodium carbonate reacts in a similar way with sand.

(i) Write the chemical equation for the reaction of sodium carbonate with sand.

(ii) Name the salt produced in the reaction between sodium carbonate and sand.

QUESTION 2: Non metal

(a) When limestone (calcium carbonate) is strongly heated, it decomposes into calcium oxide and carbon dioxide. The reaction is reversible. Write down the chemical equation, with state symbols, for the dissociation of limestone.[2]

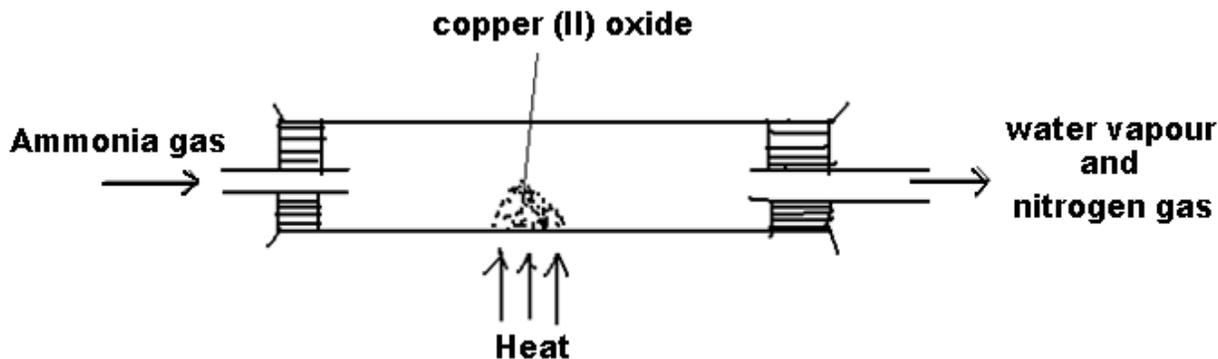
(b) When water is slowly added drop by drop to freshly made calcium oxide, a hissing sound is heard, then the oxide cracks, puffs and finally crumbles to a powder to form slaked lime which dissolves in water to give lime-water.

(i) What is the chemical name and formula of slaked lime?[2]

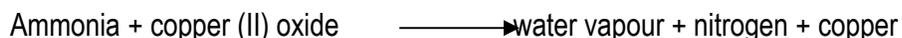
ii) What is the chemical name and formula of lime water?[2]

QUESTION 3

Ammonia gas is passed over heated copper (II) oxide as shown in the diagram below:



The word equation for the reaction which occurs is:



(a) Write the balanced chemical equation including state symbols for the reaction above.

(b) The above reaction is 'redox' reaction.

(i) What is meant by a 'redox' reaction?

(ii) State which substance is oxidized and which one is reduced. Give a reason for each answer.

(iii) State how the oxidation numbers of nitrogen and copper change in this reaction. [2]

Grade 12term 3

OXYGEN

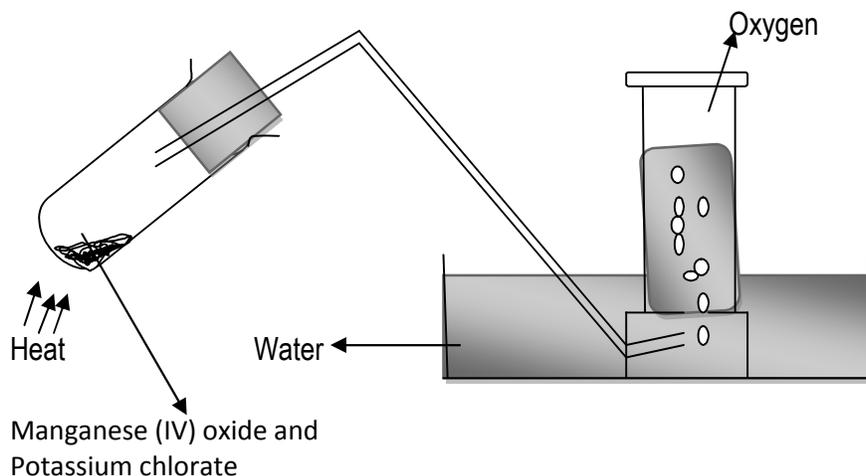
Oxygen as an element is in group six together with sulphur and selenium. At room temperature and pressure it exist as a diatomic gas molecule. The two atoms of oxygen are bonded together a double covalent bond($O=O$). In nature the main source of oxygen is photosynthesis.

In uncombined state oxygen exist in air, forming 21% by volume. In combined state it exist in form of water, silicates, metallic oxides, non metallic oxides and in the form of salts.

LABORATORY PREPARATION OF OXYGEN

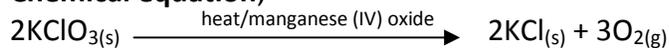
The diagrams below shows how oxygen can be prepared in the laboratory

a. Using potassium permanganate and manganese (IV) oxide.



This method is possible because oxygen gas is slightly soluble in water

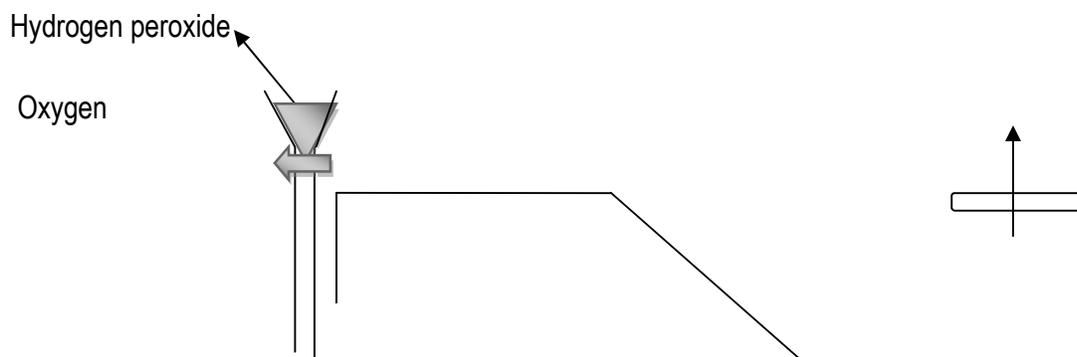
Chemical equation;

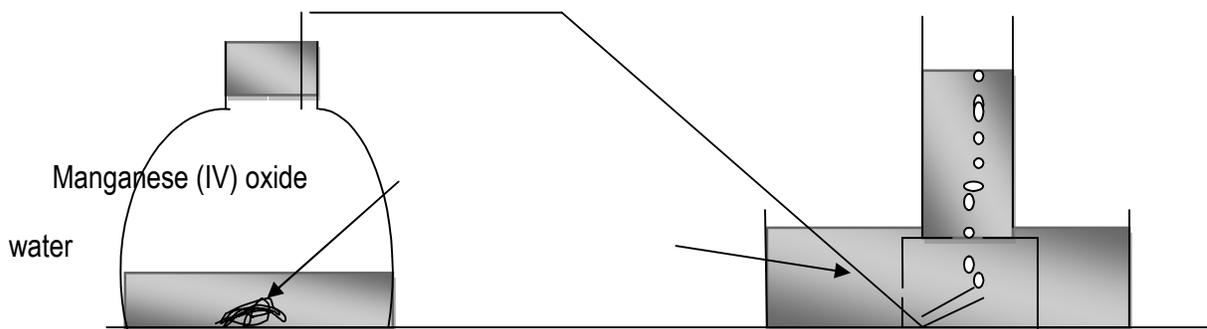


Caution:

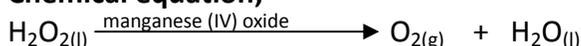
Potassium chlorate may be explosive so care must be taken when using potassium chlorate. The above method can also be used to prepare oxygen by using potassium permanganate

b. Using hydrogen peroxide and manganese (IV) oxide





Chemical equation;



Industrial preparation of oxygen

Air is the main source of oxygen, nitrogen and noble gases. The gases are obtained by fractional distillation of liquid air. The gases are passed through fine filters to remove dust. The air is then cooled to -80°C to remove water vapour and carbon dioxide as solids. If these are not removed then serious blockade of pipes may happen. After removing water and carbon dioxide had been removed the cold air is compressed to 100 atmosphere of pressure and this warms up the air. The warmed air is passed into a heat exchanger to cool it down again. The cold compressed air is allowed to expand rapidly and that cools it still further. The process of compression followed by expansion is repeated until the air reaches a temperature below -200°C . At this temperature the majority of the air gases liquefy i.e. they become liquids. The liquid air is passed into a fractionating column and the gases distilled by fractional distillation. Oxygen is distilled at -183°C . It is stored in tanks and cylinders and it is ready for use

Physical properties of oxygen;

1. It is a clear colourless gas
2. It has no smell
3. It is slightly soluble in water
4. It less dense than air
5. It is neutral to litmus paper

Identity test;

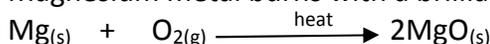
When a glowing splint of wood is introduced into the gas jar containing; the glowing splint re-lights or rekindles or bursts into flame.

Chemical properties;

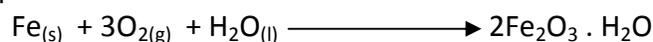
1. It is a very reactive gas combining with many metals to form the oxide of the metal;
- a. It reacts with sodium metal to sodium oxide;



- b. Magnesium metal burns with a brilliant white flame.



- c. In the presence of water it reacts with iron and this causes rusting



Iron + oxygen + water \longrightarrow hydrated iron (III)oxide

2. It is a very reactive gas combining with many nonmetals to form the oxide of the nonmetal;

a. It reacts with hydrogen to form water;



b. It reacts with sulphur to produce sulphur dioxide



3. It supports life because it is used in respiration;



4. It supports combustion of fuels;



Uses of oxygen;

1. Used in steel production, this is done by blowing oxygen into hot cast iron to burn away residues of carbon, sulphur and other nonmetal impurities.
2. It is used for producing very hot flames for welding and cutting; this is done by mixing it with acetylene to form oxy-acetylene.
3. It is used in hospitals to help patients with breathing difficulties especially those admitted to the intensive care
4. It is used to restore life to polluted lakes and rivers and in the treatment of sewage
5. It is use by fire fighters so that they can breathe while in the midst of smoke. Astronauts to carry oxygen so they can respire in space.
6. Space shuttles use oxygen gas in fuel cells which convert chemical energy to electrical energy
7. It is carried in space rockets so that hydrogen and kerosene can burn

Pollution

Some of the oxidation reactions involving oxygen cause pollution of the environment. The common pollutants arising from the reaction of oxygen with substances include; sulphur dioxide, carbon monoxide, lead dioxide, carbon dioxide and oxides of nitrogen

OZONE (O₃)

Ozone is an important gas in the earth's atmosphere because it prevents harmful ultraviolet radiation from reaching the earth. In recent years it has been noticed that there has been reduction in the amount of ozone in our atmosphere and the formation of ozone holes. The reduction or depletion of ozone in our atmosphere will lead to increased risks of diseases such as skin cancer and respiratory diseases.

HYDROGEN

Hydrogen is the lightest of all the elements. It is so light that there is none in the air; it has escaped into the outer atmosphere.

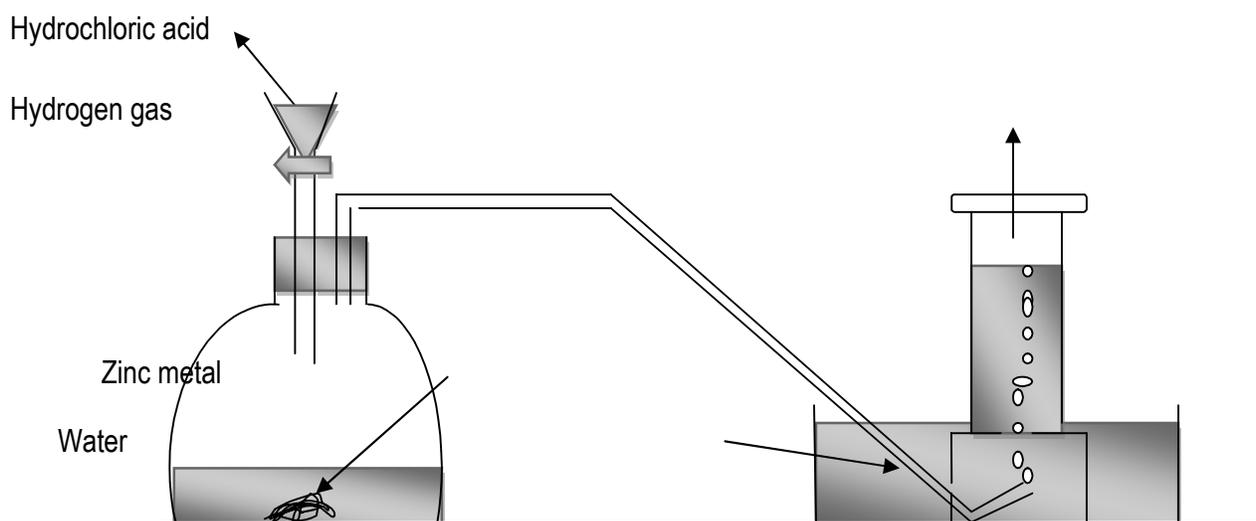
Production of hydrogen

Hydrogen is produced when reactive metals react with water, steam and dilute acids. Thus;

- a. Water + calcium metal \longrightarrow calcium hydroxide + hydrogen gas
 $\text{Ca}_{(s)} + \text{H}_2\text{O}_{(l)} \longrightarrow \text{Ca}(\text{OH})_{2(aq)} + \text{H}_{2(g)}$
- b. Team + magnesium \longrightarrow magnesium oxide + hydrogen gas
 $\text{H}_2\text{O}_{(g)} + \text{Mg}_{(s)} \longrightarrow \text{MgO}_{(s)} + \text{H}_{2(g)}$
- c. Dil. hydrochloric acid + Zinc metal \longrightarrow Zinc chloride + hydrogen gas
 $\text{Zn}_{(s)} + 2\text{HCl}_{(aq)} \longrightarrow \text{ZnCl}_{2(aq)} + \text{H}_{2(g)}$

Laboratory preparation of hydrogen;

Hydrogen gas is prepared by reacting metal with a dilute acid. The diagram below shows how hydrogen is prepared in the laboratory.



In the above diagram hydrogen is prepared and collected by downward displacement of water. This method is possible because hydrogen is slightly soluble in water.

Chemical equation;



Identity test for hydrogen gas;

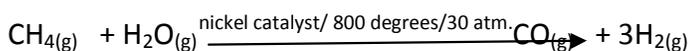
When a burning splint is introduced into a gas jar containing hydrogen gas; the hydrogen gas burns with a 'pop' sound.

Industrial manufacture of hydrogen gas;

There are two chief methods of manufacturing hydrogen. The two methods are steam reforming and electrolysis.

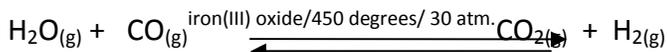
a. Steam reforming;

Hydrogen is made from hydrocarbons in a process called steam reforming. One hydrocarbon that is used to manufacture hydrogen is methane. Methane (natural gas) is mixed with steam and passed over nickel catalyst at 800°C and 30 atmospheres. The equation for the reaction is;





The reaction above is reversible. A high temperature and pressure are needed to give a good yield. The carbon dioxide is converted to carbon dioxide in another reversible reaction using iron (III) oxide.



The carbon dioxide is removed by scrubbing the gases with an alkali

b. Electrolysis;

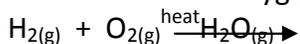
Hydrogen is obtained as a by-product in the electrolytic manufacture of chlorine from sodium chloride

Physical properties of hydrogen gas;

1. It is the lightest gas. It is about 20 times lighter than air.
2. It is colourless
3. It has no smell
4. It is slightly soluble in water
5. It has no taste
6. It has no effect on litmus paper

Chemical properties of hydrogen gas;

1. It combines with oxygen to form water. A mixture of the two gases explode when lit



2. Hydrogen is a reducing agent because it removes oxygen from a substance e.g.



3. Hydrogen is added to an alkene to form an alkane. This is called hydrogenation of alkenes;



4. Hydrogen reacts with chlorine to form hydrogen chloride gas. The hydrogen chloride gas if dissolved in water form hydrochloric acid.



Uses of hydrogen;

1. Hydrogen is used to harden vegetable oils to make margarine
2. Hydrogen is used as a fuel for space rockets because it is light and does not produce pollutants
3. Hydrogen is used in the manufacture of ammonia in the Haber process
4. Hydrogen is used in the synthesis of hydrochloric acid
5. It is used in oxy-hydrogen flame for cutting and welding.
6. Hydrogen is used to convert coal to synthetic 'petrol'

WATER:

Water is an important oxide of hydrogen because of the following properties;

1. It is a solvent for many solutes
2. In a liquid state it has a wide temperature range; from 0°C to 100°C
3. It expands when it changes from a liquid state to a solid state
4. It has no taste when pure.

Uses of water in homes;

In homes water is used for cooking, Laundry, bathing, gardening, drinking and cleaning.

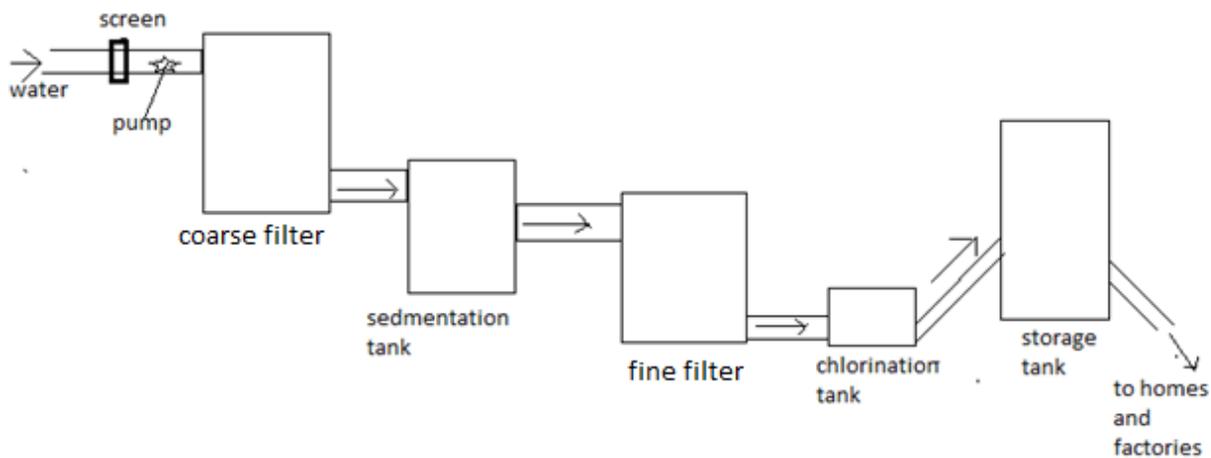
Uses of water in industry;

In industry water is used in beer brewing, washing cars, detergent manufacturing, sugar manufacturing, extraction of raw materials etc.

Clean water supply to homes;

Although there is a lot of water on earth most of it is unfit to drink because it is either salty, or polluted by industrial effluents, or polluted with sewage or simply contains a lot of fecal matter.

Water is made safe to drink by filtration and chlorination. Filtration removes insoluble substances while chlorination kills germs. The diagram below illustrates the process at water works.



1. At water works water is pumped through a screen to remove rags, wood, paper and other debris.
2. From the screen water is pumped through the coarse filter which traps larger solid particles that escaped the screen stage. Coarse filter could be sand or anthracite.
3. From the coarse filter water is pumped into a sedimentation tank where a chemical called alum is added to make smaller particles stick together and drop to the bottom.

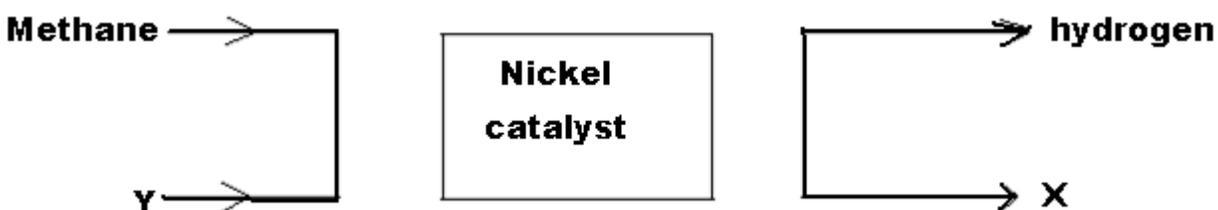
- From the sedimentation tank water is pumped through a fine filter. The fine filter could be fine sand or carbon granules with thousands of tiny pores. The carbon removes tastes and smells.
- From the fine filter water is pumped into chlorination tank where chlorine is added to kill bacteria.

At times fluoride is added before water leaves the storage tank for homes to help fight tooth decay

Exercises

QUESTION 1: non metal

Hydrogen can be manufactured from methane using a nickel catalyst as shown in the diagram.



- Identify the substance X and Y. [2]
- Write a balanced equation for the chemical reaction in the manufacture of hydrogen. [2]
- State one chemical property of hydrogen. [1]
- Mention one industrial use of hydrogen. [1]

IDENTIFICATION OF IONS AND GASES

An ion is a charged particle of matter. A positively charged ion is called a cation while a negatively charged ion is called an anion.

Identification of cations

s/n	Cation	Method/ test	Result	Inference
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1.	Aluminium (Al ³⁺)	Add a bit of sodium hydroxide solution to the portion, then add in excess		
		Add a bit of ammonia solution, then add in excess		
2	Calcium (Ca ²⁺)	Add a bit of sodium hydroxide solution to the portion, then add in excess		
		Add a bit of ammonia solution, then add in excess		
3	Zinc (Zn ²⁺)	Add a few drops of sodium hydroxide solution to the portion, then add in excess	A white precipitate is formed, if more is added to the solution the precipitate dissolves giving a colourless solution	
		Add a bit of ammonia solution, then add in excess		
4	Copper (Cu ²⁺)	Add a bit of sodium hydroxide solution to the portion, then add in excess		
		Add a bit of ammonia solution, then add in excess		
5	Iron (II) (Fe ²⁺)	Add a bit of sodium hydroxide solution to the portion, then add in excess		
		Add a bit of ammonia solution, then add in excess		
6	Iron (III) (Fe ³⁺)	Add a bit of sodium hydroxide solution to the portion, then add in excess		
		Add a bit of ammonia solution, then add in excess		
7	Ammonium (NH ₄ ⁺)	To the portion add sodium hydroxide, aluminium foil and heat gently		

Identification of anions

s/n	Anion	Method/ test	Result	Inference
	Sulphate	To the solution add a bit of	A white precipitate of	

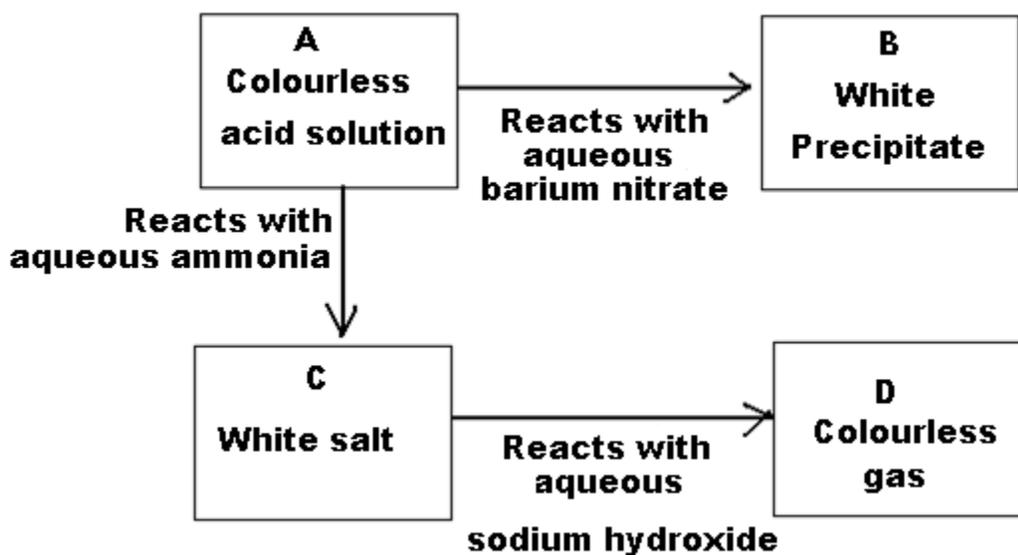
1	(SO ₄ ²⁻)	nitric acid, then add barium nitrate	barium sulphate is formed	Sulphate ions present
2	Carbonate (CO ₃ ²⁻)	Add dil. acid to the portion of the sample	Effervescence and a colourless gas given out. The gas turns clear lime water milky	Carbonate ions present
3	Nitrate (NO ₃ ⁻)	To the solution add a few drops of aqueous sodium hydroxide, add aluminium foil and heat gently.	A colourless gas is given out; the gas turns damp red litmus paper blue.	Nitrate ions present
4	Chloride (Cl ⁻)	To the solution add few drops of nitric acid. Add aqueous silver nitrate	White precipitate of silver chloride formed	Chloride ions present
5	Iodide (I ⁻)	To the solution add few drops of nitric acid. Add aqueous lead(II) nitrate	A yellow precipitate of lead iodide formed	Iodide ions present

Identification of gases/ identity test for gases

s/n	Gas	Method	Result	Inference
1	Oxygen (O ₂)	Introduce a glowing splint into the gas jar	The glowing splint bursts into flames	Oxygen gas present
2	Hydrogen (H ₂)	Introduce a burning splint into the gas jar	The gas burns with a 'pop' sound	Hydrogen gas present
3	Chlorine (Cl ₂)	subject a damp blue litmus paper to the gas	The damp blue litmus paper is bleached	Chlorine gas present
4	Ammonia (NH ₃)	introduce a damp red litmus paper to the gas	The damp red litmus paper turns blue	Ammonia gas present
5	Carbon dioxide (CO ₂)	Bubble the gas into clear lime water	The clear lime water turns milky	Carbon dioxide present
6.	Sulphur dioxide (SO ₂)	Bubble the gas into aqueous potassium dichromate	Aqueous potassium dichromate turns to green	Sulphur dioxide gas present

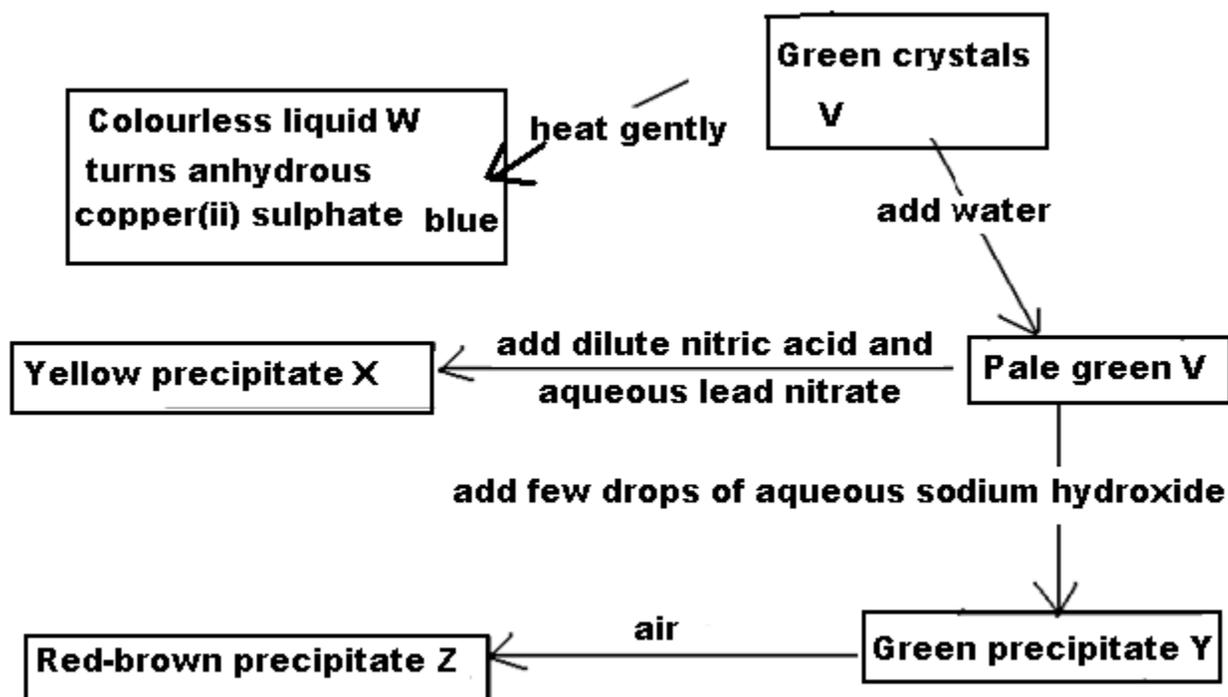
Exercise

1. The diagram below shows some of the properties and reactions of aqueous ammonia and some other substances.



- (i) Identify, by name, the substances A, B, C and D. [4]
- (ii) Write an ionic equation including state symbols for the reaction of
- (a) A with aqueous barium nitrate to form B [2]
 - (b) A with aqueous ammonia to form C [2]
 - (c) C with aqueous sodium hydroxide to form D [2]
- (iii) Name the method you would use to collect gas D [1]

3. Study the reaction scheme below and attempt the questions that follow



- (a) Identify by name;
- (i) The green precipitate Y [1]
 - (ii) The yellow precipitate X [1]

- (iii) the red-brown precipitate Z [1]
- (iv) the colourless liquid W [1]
- (v) the green crystals V [1]
- (b)(i) Name the only anion identified by aqueous sodium hydroxide and aluminium foil. [1]
- (ii) What reagent is used to distinguish Zn^{2+} and Al^{3+} ? Describe how? [2]
- (iii) Why are the compounds of zinc and aluminium soluble in aqueous hydroxide?