KITWE DISTRICT EDUCATION BOARD

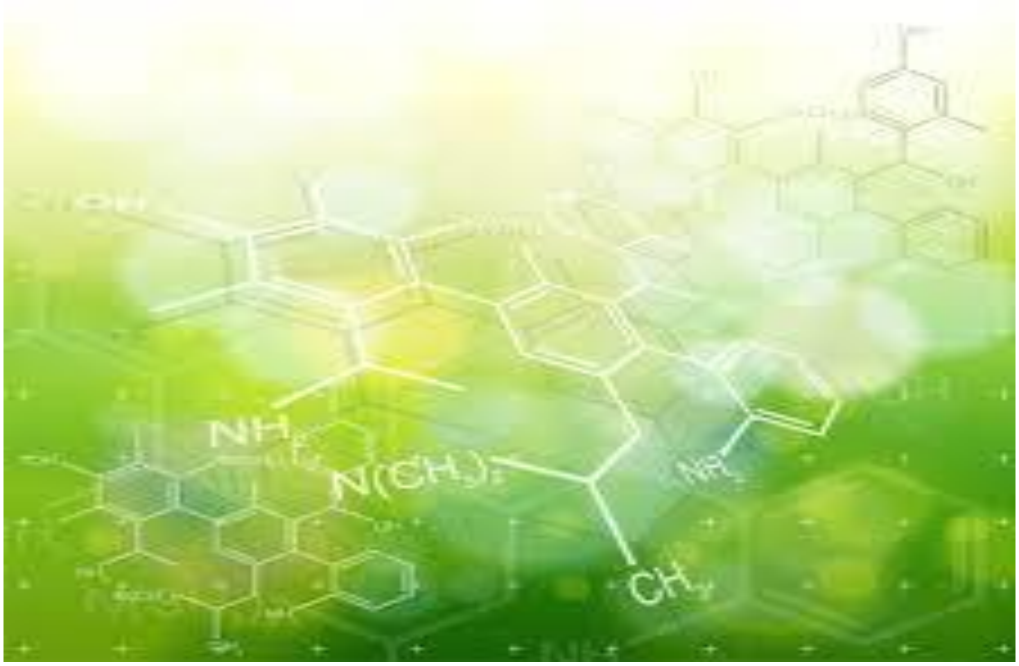
NATURAL SCIENCE

Revision work

Grade 10, 11 & 12

CHEMISTRY

(5070)



VISION; 100% PASS "ICHISWATI"

VOLUME 1

ALLAN KAOMA

TABLE OF CONTENTS

VISION

Our vision is for 100% pass percentage in chemistry (5070).

THE AUTHORS

- **Sakala Chiso, Head of Department natural science, Nkana secondary school (Team leader)**
- **Matendeko Justin, Head of Department natural science, Lima secondary school.**
- **Malama Maniford, Head of Department natural science, Mukuba secondary school**
- **Sifanu Mwaba,Action Head of Department natural science, Kakolo Maposa secondary school**
- **Sakala Isaiah,senior chemistry teacher,Ndeke secondary school.**
- **Mukwinda Gabreiel,chemistry teacher ndeke secondary school.**
- **Kalunga webby, chemistry teacher ndeke secondary school.**
- **Mbene Joseph chemistry teacher wusakile secondary school.**
- **Mambwe Friday, chemistry teacher Mukuba secondary school**
- **Mwanza Misael,chemistry teacher chemwemwe secondary school.**

INTRODUCTION

This book is the volume meant to prepare learners for the Chemistry (5070) school certificate examination in Zambia.

As a matter of information, this is the first book specifically designed to make the Chemistry (5070) syllabus. In an honest way, the idea of developing this, book was started by the, 'District Education Board Secretary" of Kitwe Mr. Allan Kaoma.

Beyond a shadow of a doubt, the book has been written in the simplest possible way while maintaining the language of chemistry. Without a doubt at, all, every pupil will seriously study the information found in this book is going to pass the final examinations with flying colors.

Undoubtedly, it is hoped that the performance of learners in Kitwe will take a turn for the better as far as Chemistry (5070) is concerned.

The book take a learner centered approach. Experiment have been included in the book and as much as possible, these should be done by every learner to make learning through enquiry and discovery most effective. The content of this book is all base on answering the Chemistry (5070) syllabus. By this it implies that the stated objective were answered with the view of helping a learner to spend less time on research work when answering past examination papers.

Science is there to meet the needs of the society, industrial applications of the concepts discussed in this book have been given whatever possible.

It is sincerely hope that the content in this book will go a long way to help the learner learn problem solving skills which will be of great value to them in later life.

GRADE 10 TOPICS

10.1 INTRODUCTION TO CHEMISTRY

10.1.1.1 Introduction to Chemistry

10.2THE PARTICULATE NATURE OF MATTER

10.2.1 Matter and the Kinetic theory

10.2.2 Diffusion

10.3EXPERIMENTAL TECHNIQUES

10.3.1 Measuring of quantities

10.3.2 Criteria of purity

10.3.3 Separating mixtures

10.4 ATOMS, ELEMENTS, COMPOUNDS AND MOLECULES

10.4.1 Atomic structure and Periodic Table

10.4.2 Bonding

10.4.3 Macromolecules

10.4.4 Chemical formulae and equations

10.1 INTRODUCTION TO CHEMISTRY

10.1.1 INTRODUCTION TO CHEMISTRY

10.1.1.1 Describe Chemistry.

- The study of the particles that make up matter and the changes the particles can undergo.

10.1.1.2 Classify the branches of chemistry.

Branches of chemistry are;

- **Analytical;** this branch of chemistry investigates what substances are made of. It helps chemists to identify chemicals that are present is the sample and measure their quantity
- **Biochemistry**; this is the field of chemistry that deals with the study of the chemical nature of the living matter. It investigates the chemical compounds and energy changes in living system.
- **Inorganic;** is the study of compounds that do not contain carbon except carbonates.
- **Physical Chemistry;** is the study of physical characteristics of materials and how they react to each other.
- **Organic chemistry;** is the study of substances that contains carbon.

10.1.1.3 Explain the importance of chemistry.

Importance of chemistry

- **Agriculture;** Farmers use fertilizers made from chemical elements such as nitrogen, phosphorous to improve soil quality.
- **Industry** e.g.in mining, chemistry is applied in the extraction of metals such as copper from copper ores.
- **Plastics;** are made from crude oil or coal.
- **Medicine;** in the manufacture of drugs from plants, chemical processes are used.
- **Home use** e.g. detergents and household soaps are designed to do the best job possible from the knowledge.

10.1.1.4 Describe the challenges of chemical industrial activities

- Production of undesired harmful by-products such as sulfur dioxide, carbon monoxide harm environment.

10.1.1.5 Demonstrate an appreciation of safety in the laboratory.

Safety rules in the laboratory

The following are the safety rules of the laboratory;

- You should always wear shoes in the laboratory.

- A learner should not carry out an experiment, without permission from the teacher. When carrying out an experiment you should always follow instructions carefully.
- Any accident or breakage of glassware must be reported to the teacher or to the laboratory technician.
- Eye goggles should be used for eye protection where necessary.
- Only specified and small amounts of substances should be used in reactions to avoid waste and reactions which cannot be controlled.
- Make sure that you know the substance being used unless you are told to use the substance as unknown.
- Do not use broken glassware.
- After working out in the laboratory, clean your apparatus, the bench and your hands.
- Always check that the label on the bottle is that of the chemical you need.
- Never hold bottles by the neck. Always hold the bottom.
- Never remove chemicals or equipment out of the laboratory.

Safety warning in the laboratory;

- Never fight or play from the laboratory.
- Never smell gases directly. Do so only cautiously by wafting fumes of gas toward your nose.
- To avoid accidental fire, turn off the gas when you are not using the burner.
- Never get your clothes or head near a Bunsen flame.
- Always work steadily and carefully.
- Never bring flammable substances near a flame. In case fire breaks out, turn off the gas and electricity. Use water if necessary.
- If a chemical is taken into the mouth accidentally or spills onto any part of t the body, it should be washed off immediately with water and the accident to the teacher.
- When you are to smell a substance or vapour, do not hold it near the nose. Hold it away from the nose and wave the vapour with your hand toward the nose and sniff carefully.
- Never hold very hot objects with your bare hands. Always hold them with a test tube holder, tongs or a piece of cloth or paper. Hot objects should be placed on a heat proof mat.
- When performing an experiment, do not point the mouth of the test tube containing a chemical being heated towards yourself or another person.
- Always add concentrated acid to water not water to concentrated acid.

- Never eat, drink or test anything in the laboratory unless you are allowed by the teacher.
- Always report accident however minor to your teacher or laboratory attendant.

10.2 THE PARTICULATE NATURE OF MATTER

10.2.1 MATTER AND THE KINETIC THEORY.

10.2.1.1 Describe matter.

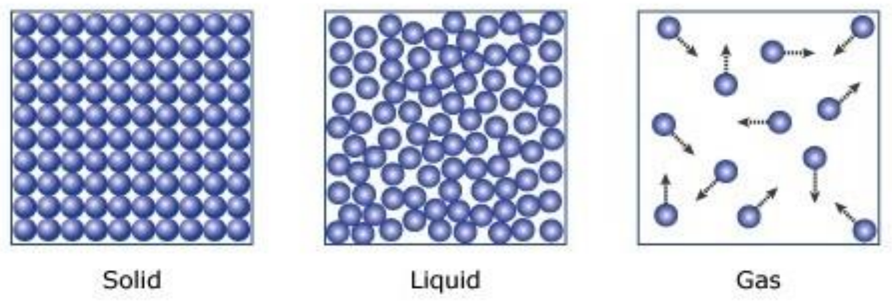
- Matter is anything that occupies space and has mass.

10.2.1.2 Classify the basic units of matter

- The particles found in matter are atoms ,molecules and ions

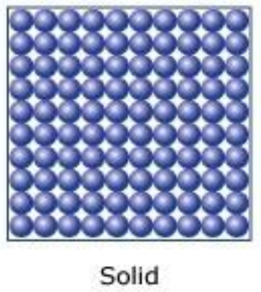
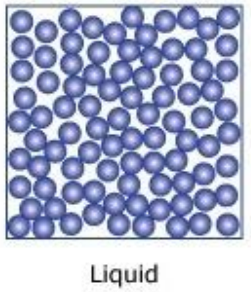
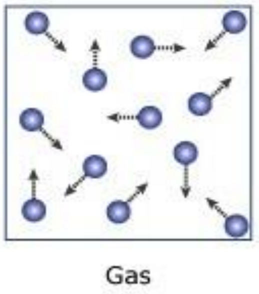
10.2.1.3. Classify the states of matter. (Kinetic theory; in terms of particle arrangement and movement)

- **Solid;** particles of a solid are closely packed and in constant vibration in their fixed position
- **Liquid;** distance between particle of a liquid are far apart from each other as compared to the distance between solid particle and move about randomly
- **Gas;** the particles that make up a gas are very far from each other when compared to the distance of particle in liquids. This is the why gas is easily compressed.



Figure; showing the three state of matter.

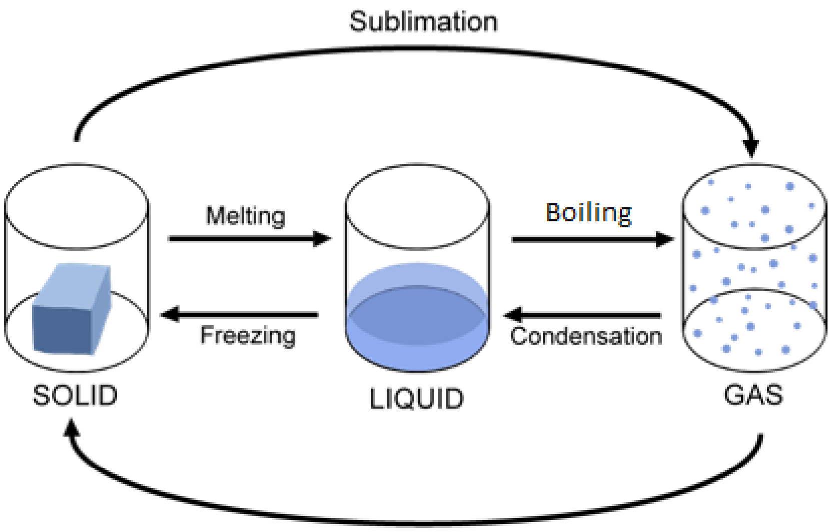
Table; showing the characteristics of the three state of matter.

STATE	CHARACTERISTICS	HOW ARE THE PARTICLES ARRANGED	EXAMPLES
SOLID	Solids have fixed volume and shape. To change the shape of a solid requires force, such as when it is broken or cut with an instrument.	 Solid	Stone ,ice block, table salt, wood
LIQUID	Liquids have fixed volume but not fixed shape. A liquid can change shape its shape and take on the shape of the container.	 Liquid	Water, paraffin, cooking oil
GAS	Gas have no fixed shape and volume. A gas will expend to take the full volume of the container it is in.	 Gas	Oxygen ,carbon dioxide, chlorine gas

10.2.1.4 Illustrate changes of states of matter.

- **Melting is** the change of state from solid to liquid as a result of increase in temperature. The distance between particles increase
- **Freezing point**, is the change of state from liquid to solid. The distance between particles of the liquid reduce until solid state is achieved.

- **Boiling poin**t, is the temperature at which liquid changes to gas. The temperature remains constant at this point until all the liquid has changed to gas. The energy supplied to the liquid is used to increase the distance between the particles of the liquid.
- **Condensation**, is the change of state from gas to liquid.
- **Sublimation, is the** change of state from solid to gas or gas to solid.
- **Evaporation r vaporization;**



10.2.1.5 Describe the absorption of heat and release of heat during changes of states of matter

- Changing states of matter occurs by;
- Exothermic reaction, which is the release of heat during a reaction and endothermic reaction, absorption of heat during a reaction.

10.2.1 Diffusion

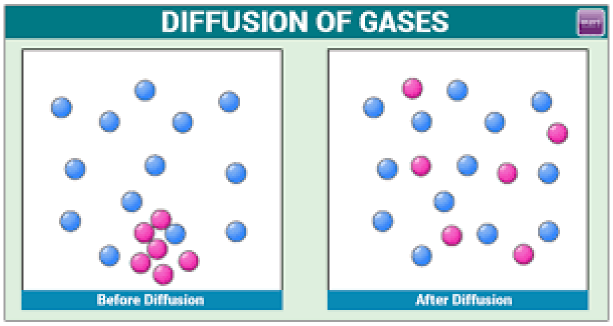
10.2.2.1 Define diffusion

- Is the movement of particles from region of higher concentration to region of lower concentration?

10.2.2.2 Demonstrate diffusion in fluids. (Liquid and gases; Brownian motion)

Brownian motion is the random movement of microscopic particles suspended in a liquid or gas caused by collisions with molecules of the surrounding medium.

- Gases; Perfume sprayed in one corner of the classroom eventually is detected in all parts of the room.



Before diffusion After diffusion

Figure; showing diffusion in gases.

- Liquids; A coloured juice poured in water eventually distribute itself uniformly in water until all parts have the colour.

10.2.2.3 Describe the factors that affect the rate of diffusion.

- **Molecular mass**; the bigger the molecular mass the slower the rate of diffusion. The smaller the mass the faster the rate of diffusion.
- **Temperature**; at higher temperature particles move faster hence faster rate of diffusion. At low temperature particles move slower hence the slower rate of diffusion. Therefore we can state that the rate of diffusion is directly proportional to temperature.
- **Concentration**; The higher the difference in concentration the faster the rate of diffusions. The lower the difference in concentration of the two substances the lower the rate of diffusion.

10.3 EXPERIMENTAL TECHNIQUES

10.3.1 MEASURING OF QUANTITIES

10.3.1 Demonstrate how different quantities are measured.

- **Time** is measured using a stop watch by determining the period elapsed between the start and ending of an event e.g. time taken for a liquid to reach boiling point.
- **Temperature** is recorded by putting a thermometer in a vessel to determine the heat released or absorbed, the melting point or melting point of a liquid.
- **Mass** is recorded by putting a substance to be measured on an electronic balance.
- **Volume** is measure using measuring cylinder, by pouring the desired liquid into the cylinder and taking the reading

10.3.1.2 Identify different measuring apparatus used in chemistry.

Some of the Measuring apparatus are

- **Stopwatch or stop clock-** for time
- **thermometers**,- for temperature
- **electronic/beam balances**, for mass
- **Burettes, pipettes, volumetric flask, measuring cylinder, and gas syringes** for recording volume.

10.3.1.3Identify various measuring instrument and other apparatus used in chemistry

- **Stopwatch or stop clock**- for time
- **thermometers**,- for temperature
- **electronic/beam balances**, for mass
- **Burettes, pipettes, volumetric flask, measuring cylinder**, **and gas syringes**- for recording volume.
- **Other apparatus are**

  Spatula, stands and clamp, test-tubes, burners, glass rods, evaporating dish, funnel beaker, conical flask etc.

Figure; showing some common laboratory apparatus

10.3.2 CRITERIA F PURITY

10 3.2 Describe the differences between a pure substance and a mixture.

- A pure substance has a sharp melting point and fixed density while an impure substance has melt or boil over a range of temperatures.

10.3.2.2 Demonstrate how to determine the purity of a substance.

- Put a thermometer into an ice or a liquid and apply heat. For a pure substance the temperature will remain at the at the same at boiling or melting point until complete change of state take place as the liquid change its state to gas or ice change state to liquid at melting or boiling points while a mixture melt or boil at a range of temperature.

10.3 2 3 Explain the importance of purity of a substance.

- purity in substances such as foodstuffs, medicines, drinks is very important because impurities can affect the health of animals e.g. human
- In industry, the computer processor chip must be in purest of silicon otherwise it will not work.

10.3.3 SEPARATING MIXTURE

10.3.3.1 Distinguish between physical and chemical changes

- **Physical change**
  - No new substances formed
  - Is reversible
  - No absorption release of energy
- **Chemical change**
  - New substances are formed
  - Not reversible by simple means
  - Energy is released or absorbed

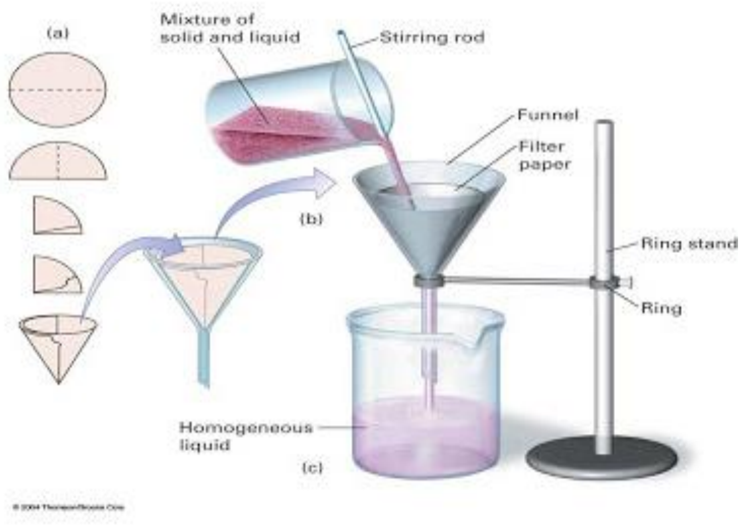
10.3.3.1 Demonstrate different methods of separating mixtures.

1. **Decantation:**

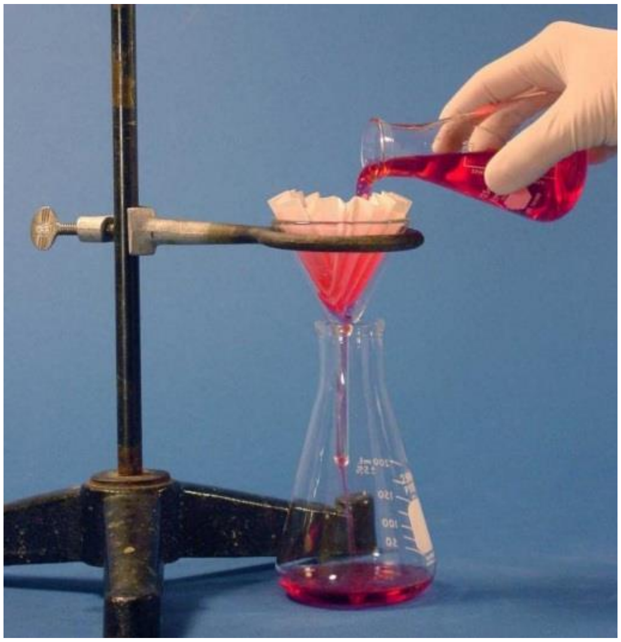
   If a mixture contains solid particles solid particle that can be seen, and that will settle to the bottom of the container, if left to stand ' it can be separated be decantation. You can. Carefully pour off the all the quid so that the solid substance remains

2. **Filtration:**

   The act or process of filtering, especially the process of passing a liquid or gas, such as air, through a filter in order to remove solid particles, is used to separate soluble and insoluble substances. A mixture of soluble and insoluble substances is dissolved in water and thereafter filtered to remove the insoluble substance. The filtrate can evaporated and crystallized to recover the solute.

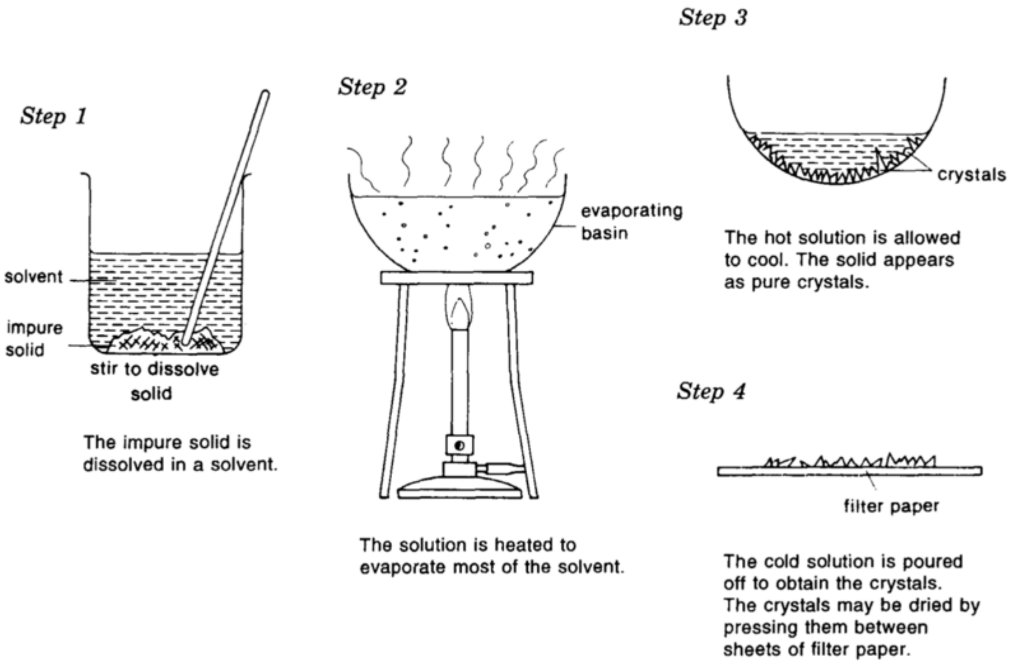


Figure; showing filtration.



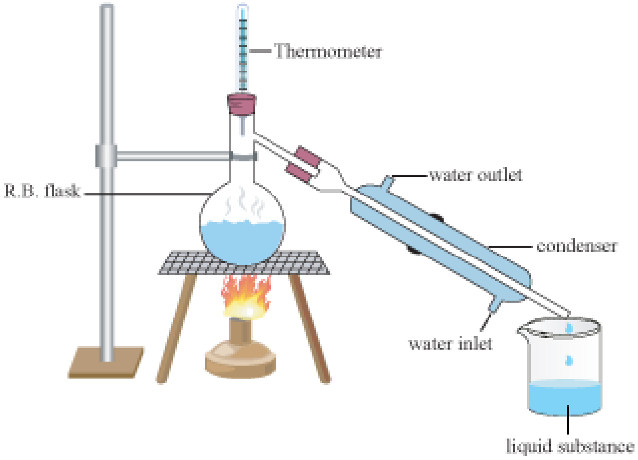
Figure; showing an experiment on filtration

3. **Crystallization** is the process used to recover a solute from. A solution is heated to saturate it and is later let to cool down. When a solution becomes saturated, it means that the solvent cannot ‘‘hold’, any more solute in solution at that temperature, so the solute is forced out of the solution in the form of solid crystals.



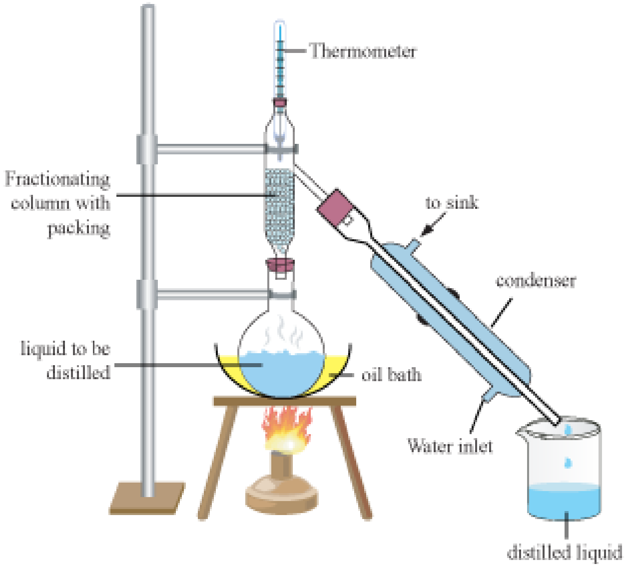
Diagram; showing the four steps of carrying out crystallization.

4. **Simple distillation** is the process used to separate a liquid from a solution by evaporating the mixture and condensing the mixture back to liquid.



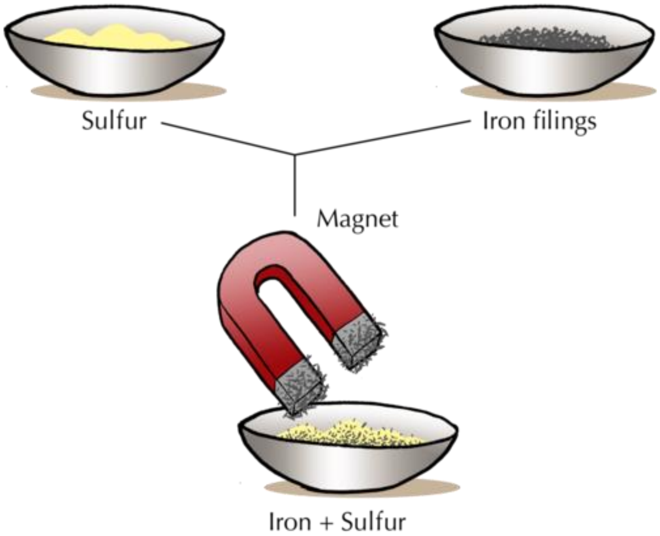
Figure; showing simple distillation.

5. **Fractional distillation**, is the process used to separate mixtures with different boiling point. e.g. pure water, boiling point 100°C and ethanol with boiling point 73°C are heated to 100°C and 73°C respectively at the temperature is let to at constant until all the ethanol evaporate and the raised to 100°C



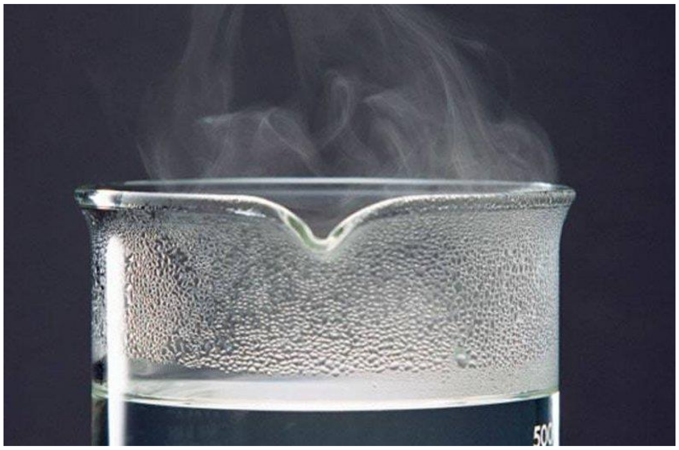
Figure; showing fractional distillation

6. **Magnetic separation**, is used to separate magnetic materials from non-magnetic material. Iron metal can be separated from a mixture of sulphur using a magnet.



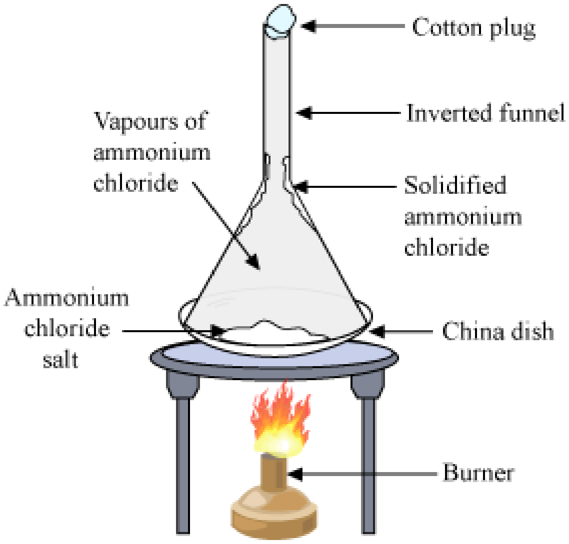
Figure; showing Magnetic separation of iron from sulphur.

7. **Evaporation,** is the process used to recover a solvent from a solution. Salt crystals can be recovered from a solution water and salt by evaporating the mixture.



Figure; showing evaporation

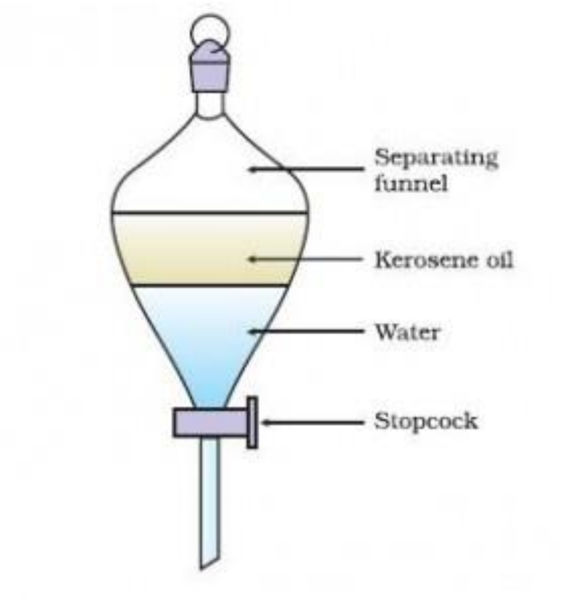
8. **Sublimation** is the change of state from gas to solid or solid to gas directly. It is a technique used to separate substances that undergo sublimation. It works well when a non-sublimating substance has a high melting point.



Figure; showing Sublimation

9. **Separating funnel; is** use to separate liquids that are immiscible, they quickly separate into layers once poured into the funnel. The tap at the bottom of the funnel is opened to pour off the bottom layer, leaving the upper layer or layers. Immiscible

Picture; showing an experiment on the use of separating funnel



Using a separating funnel makes it easier to see exactly where the oil layer starts. Therefore you can turn off the tap just before any oil contaminates the water sample. You can discard a small amount of oil to make sure that all the water particles have been removed from the funnel, before you collect the pure oil in another container.

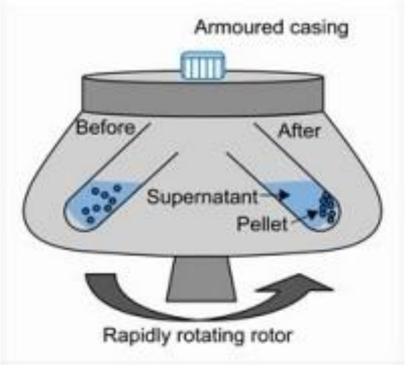
10. Centrifugation

Use a centrifuge

it is another method that can separate two or more immiscible liquids. This method uses a piece of apparatus known as a centrifuge, which rotates the container holding the sample liquid around a fixed axis, but at an angle to the vertical. As the centrifuge rotates, the more dense liquids are pushed outwards and towards the bottom of the container, while the less dense liquids are pulled inwards and upwards. The liquid on top can be poured off or using a pipette too remove it.

Figure; showing a laboratory centrifuge

A laboratory centrifuge: As the rotor rotates, the more dense substance collects at the bottom of each container



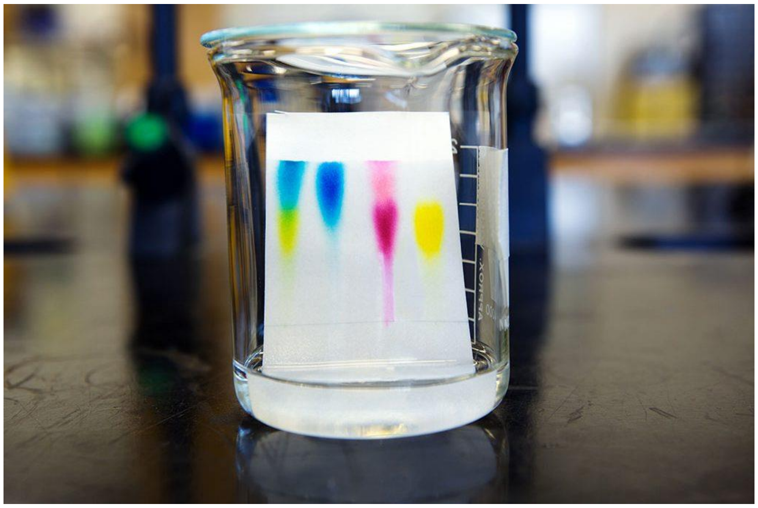
Figure; showing a laboratory centrifuge rotating.

11. **Chromatography**; this is a separating technique in which a mixture is separated as it moves across porous surface, carried by an appropriate developing solvent. Different substance in the mixture will interact differently with both the solvent and the surface, therefore they will separate. The substance that dissolves more readily and is least attracted to the surface, moves the fastest. It therefore travels furthest from the starting line, separating itself from the other substances in the mixture.

10.3.3 Interpret simple paper chromatograms.

- **Ascending paper chromatography.**

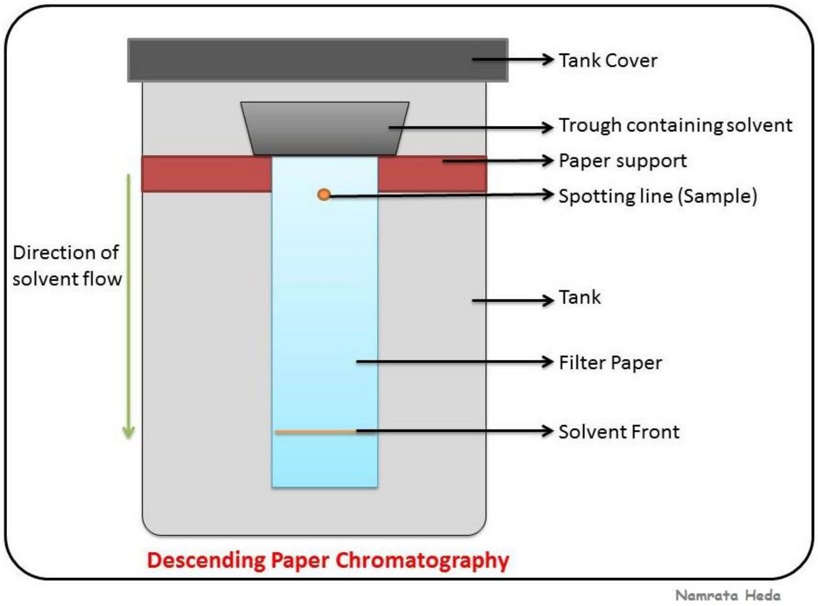
The bottom edge of the chromatogram is in the solvent. The solvent moves upwards through the paper, because the solvent's particles are attracted to the tiny spaces between the particles of paper. The solvent dissolves the sample, so the sample moves upwards along with the solvent. Ascending paper chromatography is quite a slow process.



Picture; showing ascending paper chromatography

- **Descending paper chromatography**

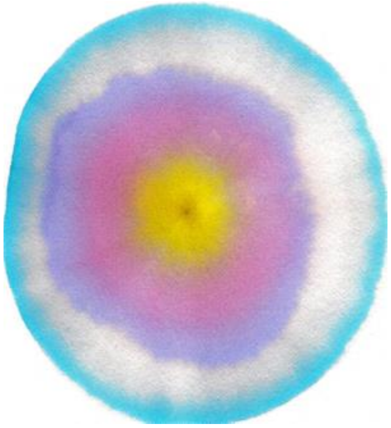
The moving solvent moves vertically downwards as the mixture separates. The chromatography paper is saturated with the solvent beforehand, and more is kept in a chamber at the top of the apparatus. The separation of the mixture occurs faster than in ascending paper chromatography, because the moving solvent is helped by gravity.



Figure; showing descending paper chromatography

- **Radial chromatography**

  The moving solvent moves outwards from a central point and separates the mixture into concentric circles (rings). The end result looks like what is the figure below.



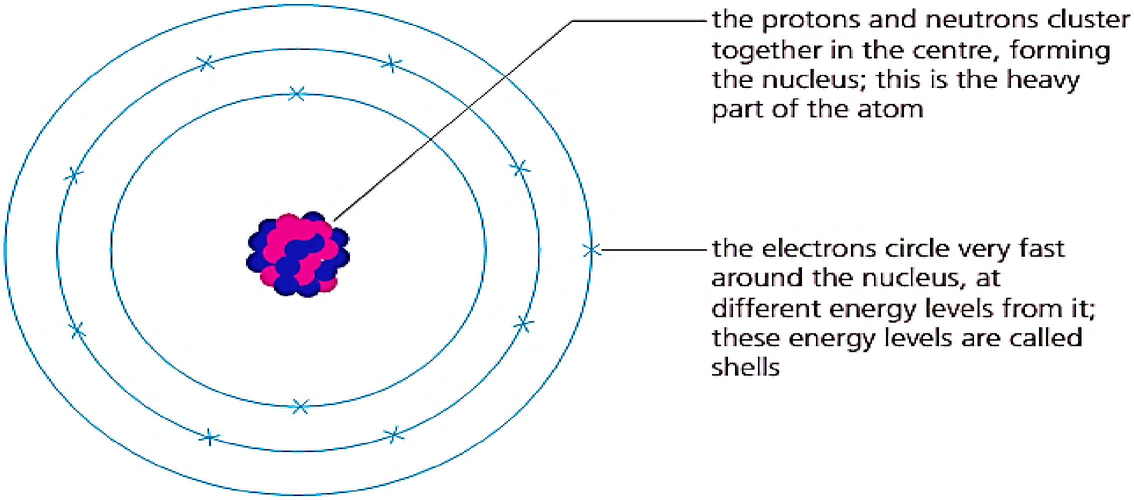
Figure; showing radial chromatography

10.4 ATOMS, ELEMENTS, COMPOUNDS AND MOLECULES

10.4.1. ATOMIC STRUCTURE AND PERIODIC TABLE.

10.4.1.1Describe an atom and its structure.

- An atom is the smallest particle of an element which takes part in a chemical reaction.
- All atoms are made up of three sub-atomic particles: **Protons: Neutrons: Electrons**: The protons and neutrons are found in the centre of the atom, which is called the **nucleus**. The neutrons have no charge and protons are positively charged. The electrons are negatively charged and move around very quickly in **electron shells** or **energy levels**



10.4.1.2 Describe the relative charges and approximate relative masses of protons, neutrons and electrons

About 1840 electrons are equal in mass to the mass of one proton or one neutron. A summary of each type of particle, its mass and relative charge is shown in the table below. You will notice that the masses of all these particles are measured in **atomic mass units (amu)**. This is because they are so light that their masses cannot be measured usefully in grams.

`

Particle	Symbol	Relative mass/amu	Relative Charge
Protons	P	1	+1
Neutrons	N	1	0
Electrons	E	$\frac{1}{1840}$	-1

10.4.1.3 Describe the proton (atomic) number and nucleon (mass) number and nuclide notations

- Proton (atomic) number is represented by letter: Z, the relative atomic mass or nucleon number is represented by letter- A. A is the sum of (protons + neutrons) and $^{A}_{Z}X$ is the nuclide notation, X represent the symbol of an element.

10.4.1.4Describe an element.

- An element contains is a substance that contains only one kind of atoms.
- Atoms are the smallest particles of matter that we cannot break down further by chemical means.

10.4.1.5 Identify elements using their chemical symbols

The atoms of the first 20 elements

In the Periodic Table, the elements are arranged in order of increasing proton number. Here are the first 20 elements, shown as a list.

Element	Symbol	Proton number	Electrons	Neutrons	Nucleon Number (protons + Neutrons)
Hydrogen	H	1	1	0	0
Helium	He	2	2	2	4
Lithium	Li	3	3	4	7
Beryllium	Be	4	4	5	9
Boron	B	5	5	6	11
Carbon	C	6	6	6	12
Nitrogen	N	7	7	7	14
Oxygen	O	8	8	8	16
Fluorine	F	9	9	10	19
Neon	Ne	10	10	10	20
Sodium	Na	11	11	12	23
Magnesium	Mg	12	12	12	24
Aluminium	Al	13	13	14	27
Silicon	Si	14	14	14	28
Phosphorous	P	15	15	16	31
Sulphur	S	16	16	16	32
Chlorine	Cl	17	17	18.5	35.5
Argon	Ar	18	18	22	40
Potassium	K	19	19	20	39
Calcium	Ca	20	20	20	40

10.4.1.6 **Describe the basis of the Periodic Table. The periodic has periods and groups**.

- It gives the names and symbols for the elements.
- The column (groups) and rows (periods) of an element **is in** gives us lots of clues about it.

  For example, look at the columns (groups) numbered I, II, III … The elements in these form families or **groups**, with similar properties.

  So if you know how one element in Group (I) behaves, for example, you can make a good guess about the others in the same group.

10.4.1.7 Describe isotopes

- Isotopes are atoms of the same element, with different numbers of neutrons

  Example of isotopes

  Oxygen atoms have 8 protons and 8 electrons. The majority of naturally occurring oxygen atoms have 8 neutrons, but a small percentage have 9 neutrons, while others have 10 neutrons. But all oxygen atoms have a proton number of 8. The table bellows shows some examples of isotopes.

Name of isotope	Proton number(Z)	Neutron (N)	Atomic Mass(A)
Oxygen-16	8	8	16
Oxygen-17	8	9	17
Oxygen – 18	8	10	18
Chlorine-35	17	18	35
Chlorine-37	17	20	37

10.4.1.8 Calculate relative atomic mass of an element given the % abundances of isotopes and from mass spectrum.

Relative Atomic mass when given % abundance.

The calculation of the average (relative) atomic mass of an atom is performed using the relative abundance data from the isotope of each atom.

Average atomic mass =

$$\frac{[(\%\text{ abundance of isotope})\times(\text{mass of isotope})] + [(\%\text{ abundance of isotope})\times(\text{mass of isotope})] + [.]}{100}$$

For example:

The natural abundance for boron isotopes is 19.9% ^{10}B (10.013 amu*) and 80.1% ^{11}B (11.009 amu*). Calculate the atomic mass of boron.

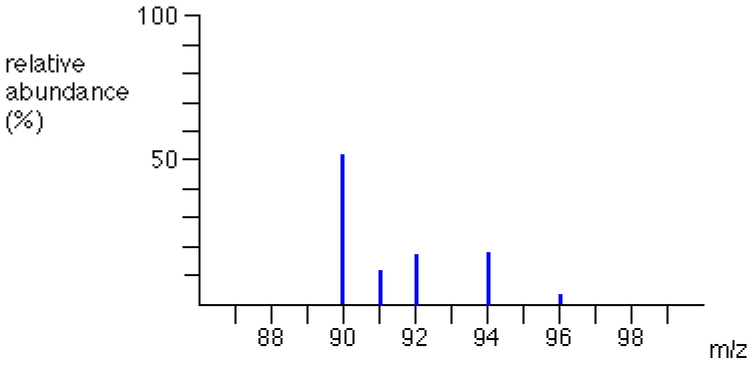
$$\text{Average atomic mass} = \frac{[(19.9\%)(10.013)] + [(80.1\%)(11.009)]}{100}$$

= 10.811 (note that this is the value of atomic mass given on the periodic table)

*amu is the atomic mass unit (u, μ or amu), which is defined as 1/12th the mass of a carbon-12 atom. This value is arbitrary and simply provides a reference point for measuring relative atomic masses.

Relative atomic mass from atomic spectrum

The mass spectrum for Magnesium



The number of isotopes

The 5 peaks in the mass spectrum shows that there are 5 isotopes of magnesium - with relative isotopic masses of 20, 21, 22, 24 and 26 on the ^{12}C scale.

10.4.1.9 Describe the use of radioactive isotope

- **To check for leaks**. Engineers can check oil and gas pipes for leaks by adding radioisotopes to the oil or gas. If a Geiger counter detects radiation outside the pipe, it means there is a leak. Radioisotopes used in this way are called tracers.

  Demonstrate the build-up of electrons in shells

- **To treat cancer** Radioisotopes can cause cancer. But they are also used in **radiotherapy** to *cure* cancer – because the gamma rays in radiation kill cancer cells more readily than healthy cells.
- T**o kill germs and bacteria** Gamma rays kill germs too. So they are used to sterilize syringes and other disposable medical equipment. They also kill the bacteria that cause food to decay.

10.4.1.10 Demonstrate the build-up of electrons in shells

- The number of shells are drawn according to the period where an element is found. If an element is found in period 3, the element has three shells. The first shell must have a maximum number of two electrons. The next shell must have maximum number of 8 electrons. The remaining electron are spilled into the next shell until all electrons are filled into the shell.

10.4.2.1 Describe a compound.

- A compound is made of atoms of different elements, bonded together. The compound is described by a formula, made from the symbols of the atoms in it. (The plural of formula is formulae.)

10.4.2.2 Describe the formation of ions (radicals).

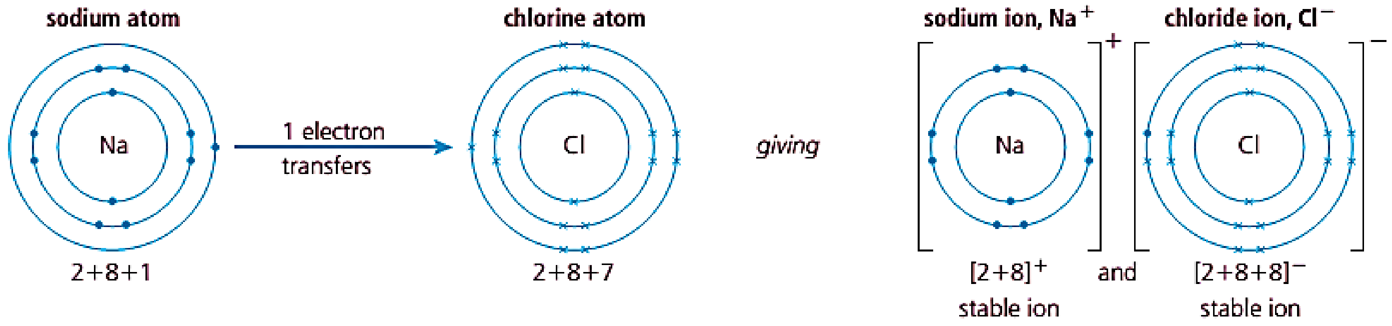
- An atom becomes an ion when it loses or gains electrons.
  An ion is a charged particle. It is charged because it has an unequal number of protons and electrons

10.4.2.3 Describe the formation of ionic (electrovalent) bonds.

- **The ionic bond is the bond that forms between ions of opposite charge.**

  For example

- When a sodium atom and a chlorine atom react together, the sodium atom loses its electron *to the chlorine atom*, and two ions are formed.
  Here, sodium electrons are shown as ° and chlorine electrons as ×



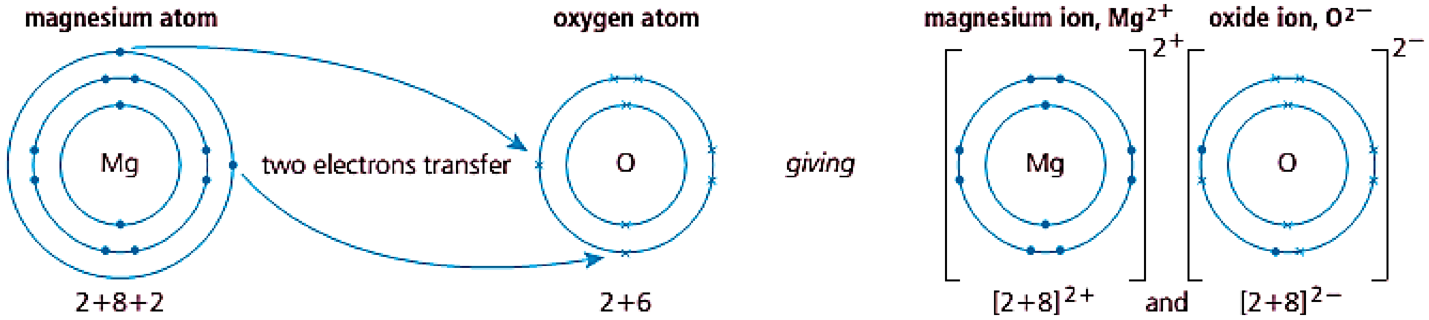
Other metals and non-metals follow the same pattern.

A metal reacts with a non-metal to form an ionic compound. The metal atoms lose electrons. The non-metal atoms gain them.The ions form a lattice. The compound has no overall charge.

Below are two more examples

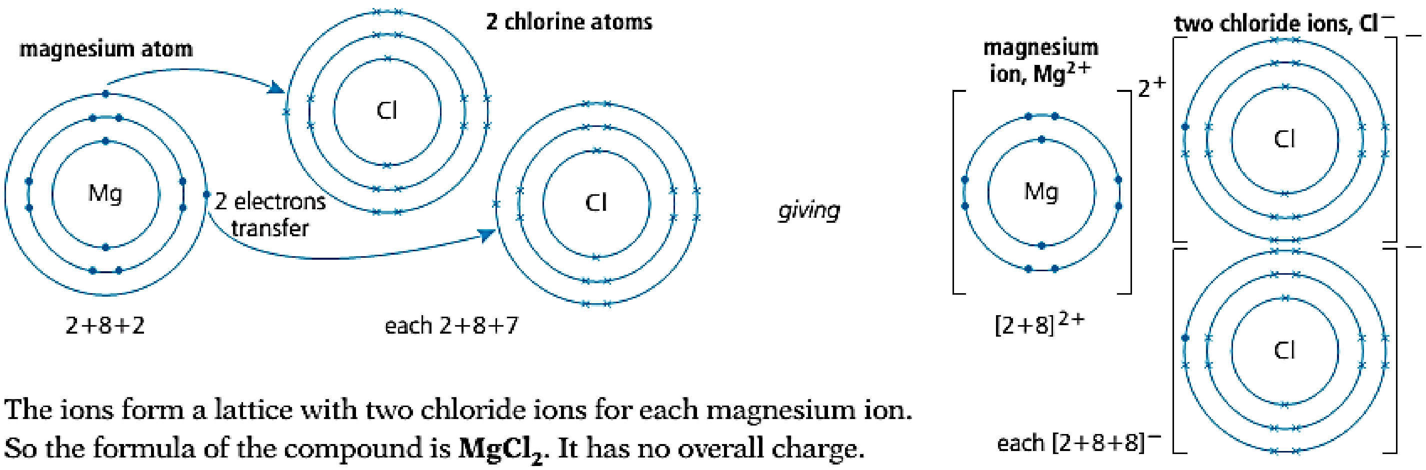
- **Magnesium oxide**

A magnesium atom has 2 outer electrons and an oxygen atom has 6. When magnesium burns in oxygen, each magnesium atom loses its 2 outer electrons to an oxygen atom. Magnesium and oxide ions are formed:



Magnesium chloride

When magnesium burns in chlorine, each magnesium atom reacts with *two* chlorine atoms, to form **magnesium chloride**. Each ion has 8 outer electrons.



The ions form a lattice with two chloride ions for each magnesium ion. So the formula of the compound is **$MgCl_2$**. It has no overall charge.

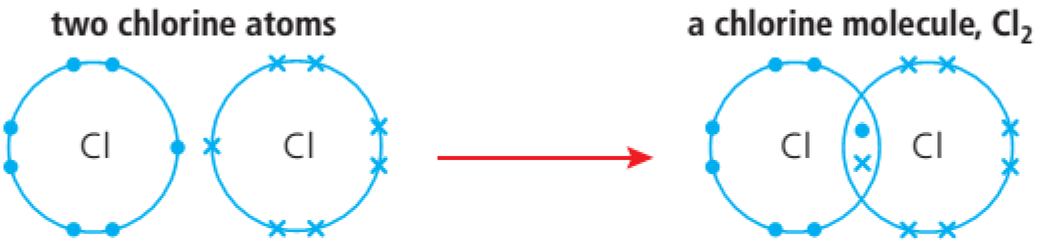
10.4.2.4Describe the formation of covalent bonds

- **Covalent bond** is chemical bond formed by the sharing of one or more pairs of electrons between two atoms.

FORMATION OF COVALENT BOND

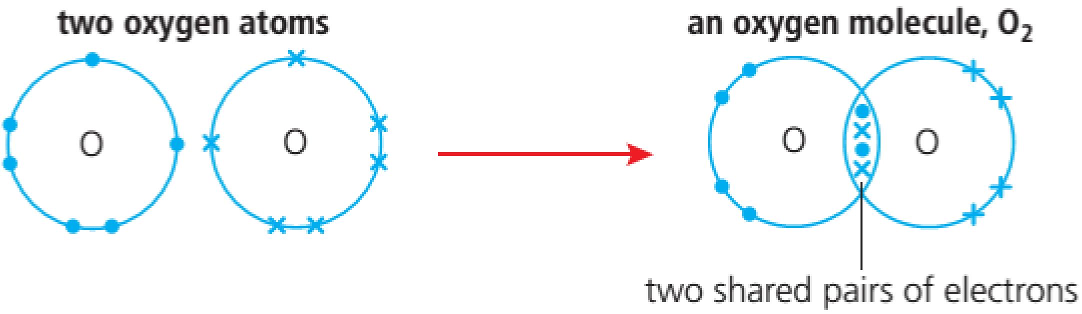
- **Chlorine.**

  A chlorine atom needs a share in one more electron, to obtain a stable outer shell of eight electrons. So two chlorine atoms bond covalently like this:



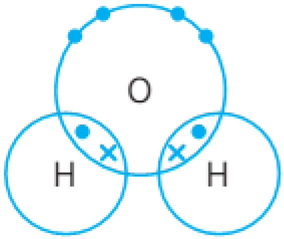
- **Oxygen**

  An oxygen atom has six outer electrons, so needs a share in *two* more. So two oxygen atoms share two electrons each, giving molecules with the formula O_2. Each atom now has a stable outer shell of eight electrons:



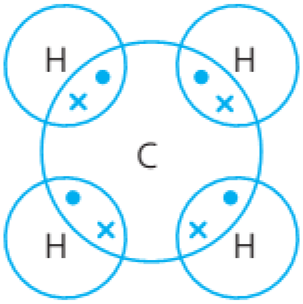
10.4.2.5 Describe the electronic arrangement in simple multiple covalent molecules

❖ WATER (H_2O)



The oxygen atom shares electrons with the two hydrogen atoms. All now have a stable arrangement of electrons in their outer shells: 2 for hydrogen and 8 for oxygen.

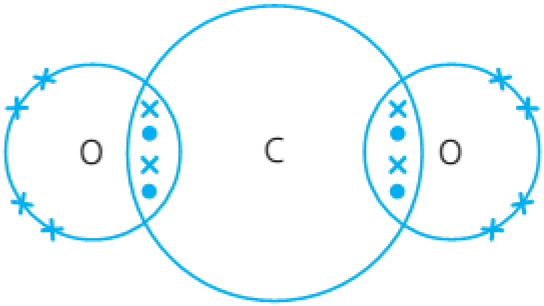
methane, CH_4



a molecule of methane

The carbon atom shares electrons with four hydrogen atoms. All now have a stable arrangement of electrons in their outer shells: 2 for hydrogen and 8 for carbon.

carbon dioxide, CO_2



a molecule of carbon dioxide

The carbon atom shares all four of its electrons: two with each oxygen atom. So all three atoms gain stable shells.

The two sets of bonding electrons repel each other. They move as far apart as they can, giving a linear molecule.

All the bonds are double bonds, so we can show the molecule like this: O = C = O.

10.4.2.6Describe the uses of ionic and covalent compounds

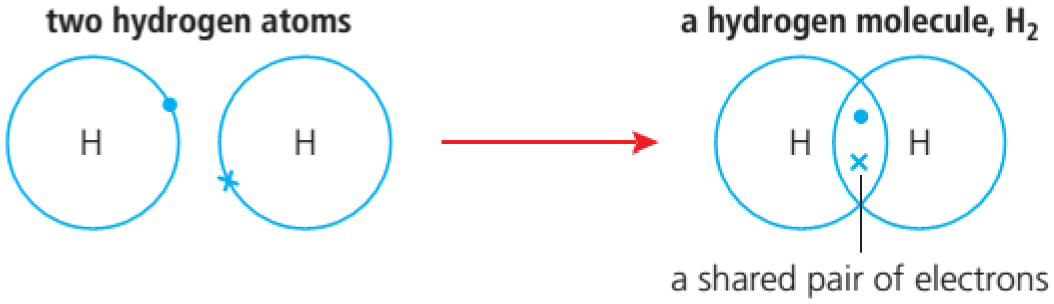
- **Uses of ionic compounds**

  ❖ Because ionic compounds have such high melting points, they can keep their strength and stability at very high temperatures. This means that they are very good refractory materials, which play an important role in some industries. Industries that make use of furnaces need to have linings in those furnaces that can withstand the extremely high temperatures. Some of these are:

- Metals: Whenever metals are processed and formed into products, the metal usually has to be melted down first. Steel mills contain blast furnaces where this happens, and those are often lined with magnesium oxide (MgO).
- Glass and cement: Furnaces used in glass- and cement making are often lined with aluminum oxide (Al_2O_3)

- **Uses covalent compounds**
  - Alcohol: Ethanol is an organic, covalent compound that we know as alcohol. Apart from alcoholic beverages, alcohol is used in medical science as a preservative and an antiseptic, and it is used in some cleaning agents and fuels.
  - Glass making: The main ingredient of glass is silicon dioxide (SiO_2)
  - Are used to make nylon ropes
  - Because they are bad conductors of electricity, covalent compounds are used as insulators
  - They used to make handles for cooking pans.

10.4.2.7 Describe a molecule



- **MOLECULES**

The two bonded hydrogen atoms above form a molecule.

A molecule is a group of atoms held together by covalent bonds.

Since it is made up of molecules, hydrogen is a molecular element. Its formula is H_2. The $_2$ tells you there are 2 hydrogen atoms in each molecule. Many other non-metals are also molecular. For example: iodine, I_2 oxygen, O_2 nitrogen, N_2 chlorine, Cl_2

Elements made up of molecules containing two atoms are called <u>diatomic</u>. So iodine and oxygen are diatomic

10.4.2.8 Describe valence and valence electrons.

- The valency of an element is the number of electrons its atoms lose, gain or share, to form a compound.
- Valence electrons are the number of electrons in the outer most shell of an atom. The group number from 1 to 7 except the noble gases represent valence electrons. Elements in the same group has the same number of valence electrons.

10.4.2.9 Demonstrate how to deduce valence of an element.

The number of electrons that an atom of any element loses, gains or share to form a compound, is called valency number, or combining power. Elements that have: 1 to 3 valence electrons lose those electrons (aluminum’s valence number is 3)

- 4 valence electrons tend to share those electrons (carbon's valence number is 4).
- 5 to 7 valence electrons tend to gain electrons to complete the noble gas.

The table below illustrate how the valency of an atom is obtained from the groups of periodic table

Group	i	ii	iii	iv	v	vi	vii
valency	1	2	3	4	3	2	1

10.4.2.10 Identify the differences in properties of ionic and covalent compounds.

Covalent compounds

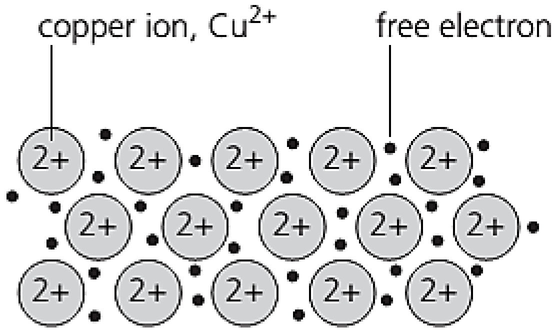
Ionic compounds	Covalent compound
Most ionic compounds are not volatile **They are made up of ions. (they do not easily evaporate at room temperature and pressure)**	Most covalent compounds are volatile. that is why they usually have distinctive smells
When melted or in solution, ionic Compounds conduct electricity.	They do not conduct electricity
Ionic compounds are usually solids at Room temperature and pressure.	Covalent compounds are usually gases or liquids at room temperature and pressure
Ionic compounds have high boiling points and high melting points,	They vaporize easily. This is because of the very weak electrostatic forces of Attraction between the molecules.
They do not vaporize easily. This is because of the very strong electrostatic Forces of attraction between the oppositely charged ions.	Many covalent compounds are insoluble in water, but soluble in other covalent liquids such as alcohol and Tetra chloromethane.
Many ionic compounds are soluble in water	Generally, covalent compounds have Lower densities than ionic compounds.

10.4.2.12 Describe metallic bonding.

Metallic bonding the lattice of positive ions in a **'sea'** of delocalised electrons. This is the type of bonding in which electrons in the outer energy level of the atom of a metal lose electrons and the lost electrons move freely throughout the structure (they are delocalised forming a mobile 'sea' of electrons), in doing so , they form a lattice of positive ions. **Therefore, metals consist of positive ions embedded in moving clouds of electrons.**

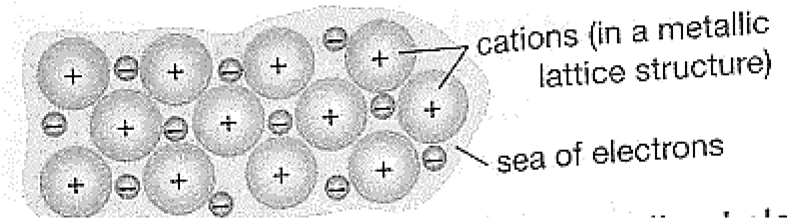
In metals, the atoms are packed tightly together in a regular lattice. The tight packing allows outer electrons to separate from their atoms. The result is a lattice of ions in a ‘sea’ of electrons that are free to move.

Look at copper:



The metallic bond is the attraction between metal ions and free electrons.

It involves the forces of attraction between cat ions (positive ions) and electrons in a metallic lattice structure. In other words, when metal-Atoms are packed together tightly in a metal lattice, their valence electrons break Free from the shells and move around the structure. The electrons become **delocalized** and form a sea" of electrons, which then acts as a glue that hold the lattice together.



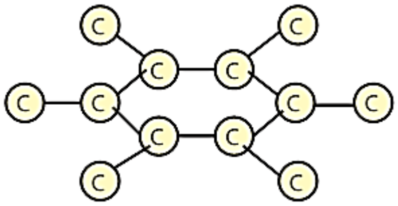
10.4.2.13 Describe the electrical / thermal conductivity of metals.

Metals are both good conductors of electricity and heat. They have very high melting points because very high heat is required to break the lattice structure. They conduct electricity because they have a sea of delocalised electrons around the lattice. The electrons can carry electrical charge throughout the lattice structure. They conduct heat because the electrons both absorb and transfer heat to the cations very well. Metals are ductile (they can be drawn into a wire) and malleable (they can be hammered into a new shape). This is because the positive ions are able to slide over each other, without breaking the bond is he electrical conductivity of free electron movement/delocalized electrons

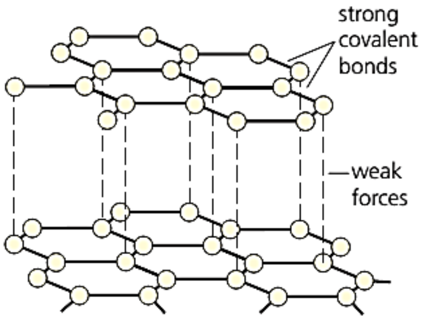
10.4.3 MACROMOLECULES

10.4.3.1 Describe the giant covalent structures of graphite and diamond

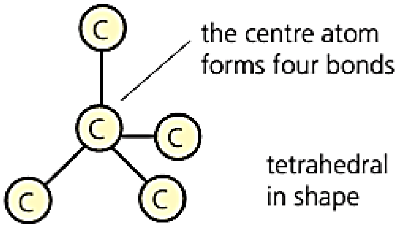
Graphite are giant structures of carbon atoms arranged in Hexagonal layers while diamond are giant structure of carbon atoms arranged tetrahedrally



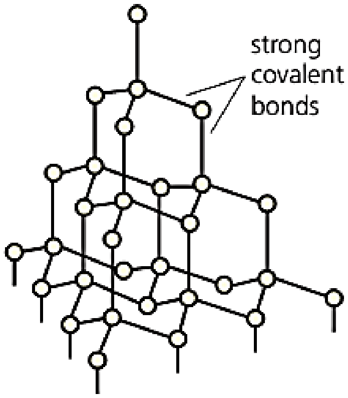
In graphite, each carbon atom forms covalent bonds to *three* others. This gives rings of *six* atoms.



The rings form flat sheets that lie on top of each other, held together by weak forces.



A carbon atom forms covalent bonds to *four* others, as shown above. Each outer atom then bonds to three more, and so on.



Eventually billions of carbon atoms are bonded together, in a giant covalent structure. This shows just a very tiny part of it.

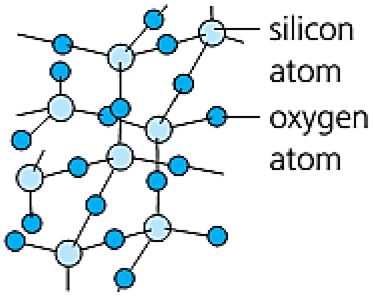
10.4.3.2 Describe the uses of graphite and diamond in relation to their structures

- Graphite is used as a lubricant, pencil leads, electrodes and diamond is used in cutting, jewelry

Substance	Properties	Uses
Diamond	**hardest known substance** **does not conduct**	**in tools for drilling and cutting**
	sparkles when cut	**for jewellery**
Graphite	**• soft and slippery**	**for pencil 'lead' (mixed with clay**
	soft and dark in colour	**as a lubricant for engines and locks**
	conducts electricity	**for electrodes, and connecting brushes in generators**

10.4.3.3 Describe the macromolecular structure of silicon (IV) oxide (silicon dioxide)

- Oxygen atoms are bonded to silicon atoms tetrahedrally despite the formula being SiO2.



▲ Silicon dioxide is made up of oxygen atoms ● and silicon atoms ○. Billions of them bond together like this, to give a giant structure.

10.4.3.4 Identify the similarities in properties between diamond and silicon dioxide

- Similarities in properties between diamond and silicon dioxide are such that

  Atoms are held together by covalent bonds tetrahedrally

10.4.3.5 Demonstrate how to deduce a chemical formulae from valences.

The table below describes how a chemical formula can be derived from valences

Name of a compound	**Step 1**	**Step 2**	**Step 3**	**Chemical formulae**
Calcium hydroxide	Ca OH	2 1 Ca OH	2 1 Ca OH	$Ca(OH)_2$
Carbon dioxide	C O	4 2 C O	~~4~~ 2 ~~2~~ 1 2 1 C O	CO_2
Aluminium carbonate	Al CO_3	3 2 Al CO_3	3 2 Al CO_3	$Al_2(CO_3)_3$

Some common radicals, their valencies and some compounds formed

Name of a radical	Chemical symbol	Valency of a radical	Compound formed with a cations
Hydroxides	OH^-	1	Calcium hydroxide—($Ca(OH)_2$
Ammonium	NH_4^+	1	Ammonium sulphate—$(NH_4)_2SO_4$
Carbonates	CO_3^{2-}	2	Ammonium carbonate—$(NH_4)_2CO_3$
Sulphates	SO_4^{2-}	2	Iron(iii) sulphate—$Fe_2(SO_4)_3$
Phosphates	PO_4^{3-}	3	Ammonium phosphate—$(NH_4)_3PO_4$

10.4.41 CHEMICAL FORMULAE AND EQUATIONS

10.4.4.1 DEMONSTRATE HOW TO CONSTRUCT WORD EQUATIONS.

How to write the equation for a reaction

- These are the steps to follow, when writing an equations
  - Write the equation in words.
  - Now write it using symbols. Make sure all the formulae are correct.

- Check that the equation is balanced, for each type of atom in turn.
  Make sure you do not change any formulae.
- Add the state symbols

10.4.4.2 Formulate balanced chemical equations.

Example 1

- Calcium burns in chlorine to form calcium chloride, a solid. Write an equation for the reaction, using the steps above.
  1 calcium atom reacts with,1 chlorine molecule giving a compound calcium chloride

  2 $Ca_{(s)} + Cl_{2\,(g)} \longrightarrow CaCl_2$

  3 Ca: 1 atom on the left and 1 atom on the right.

  Cl: 2 atoms on the left and 2 atoms on the right.
  The equation is balanced.

Example 2

- Calcium carbonate react with hydrochloric acid to produce calcium chloride, water and carbon dioxide.

  $CaCO_3 + HCl \longrightarrow CaCl_2 + H_2O + CO_2$

  All atoms above are balanced except hydrogen

  $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

  Now the equation is balanced

10.4.4.3Constract net ionic equations from balanced chemical equations.

- Steps for writing an ionic equation.
  - Ensure that the equation is balanced.
  - Leave intact (do not break down) solids-either ionically or covalently bonded, liquid and gases.
  - All compounds with state symbol *aq* must be separated into a cation and anion part introducing an addition sign **+** *as* a separator.
  - The subscript numbers should be written as a coefficient of the atoms of groups of atoms it represent. For example

$CaCl_{2\,(aq)}$ ---------------------- $Ca^{2+} + 2Cl^-$

$(NH_4)_2SO_4$ ------------------- $2NH_4^+ + SO_4^{2-}$

$$CaCO_3 \quad + \quad 2HCl \longrightarrow CaCl_2 \quad + \quad H_2O \quad + CO_2$$

$$CaCO_{3(s)} \quad + \quad 2HCl_{\,(aq)} \longrightarrow CaCl_{2\,(aq)} \quad + \quad H_2O_{\,(l)} \quad + CO_{2\,(g)}$$

The ionic equation is written as follows

$$CaCO_{3(s)} \quad + \quad 2H^+_{\,(aq)} + \cancel{2Cl}^-_{(aq)} \longrightarrow Ca^{2+} + \cancel{2Cl}^-_{(aq)} \quad + \quad H_2O_{(l)} \quad + CO_{2(g)}$$

Finally we cancel out spectator ions i.e. ions that appear on both side of the chemical equation.

There are called spectator ions because they did not participate in the reaction.

$$CaCO_{3(s)} \quad + \quad 2H^+_{\,(aq)} \longrightarrow Ca^{2+}_{\,(aq)} \quad + \quad H_2O_{\,(l)} \quad + CO_{2\,(g)}$$

GRADE 11 TOPICS

11.1 ACIDS, BASES AND SALTS

11.1.1 Characteristic properties of acids and bases

11.1.2 Preparation of salts

11.1.3 Types of oxides

11.1.4 Identification of ions and gases (Qualitative analysis)

11.2 THE MOLE CONCEPT

11.2.1 Relative Masses

11.2.2 The Mole

11.2.3 Empirical and Molecular formulae

11.3 CHEMICAL REACTIONS

11.3.1 Rates of chemical reactions

11.3.2 Chemical Equilibrium

11.3.3 Redox reactions

11.3.4 Energetics of reactions

11.4 THE PERIODIC TABLE

11.4.1Groups and Periods

11.4.2 Groups and Periodic trends

11.4.3 Transition metals

11.5 ACIDS, BASES AND SALTS

11.5.1 CHARACTERISTIC PROPERTIES OF ACIDS AND BASES

11.5.1.1 Describe acids, bases or alkalis in terms of ions they contain or produce in Aqueous solution

- **An acid** is a compound that produces hydrogen ions as the only positively charged Ions in aqueous solution
- **A base** is an oxide or hydroxide of a metal including ammonium hydroxide
- **Alkalis** are soluble bases that produce hydroxide ions in aqueous solution as the only negatively charged ions

11.5.1.2 Describe the meaning of weak, strong, dilute and concentrated acids and alkalis.

ACIDS

- **Weak acids** are acids that partially ionize in water.
- **Strong acids** are acids that completely ionize in water to form hydronium ions.
- **Dilute acid** is one where a pure acid is mixed with significant percentage of water.
- **Concentrated acid** is the amount of acid dissolved in $1dm^3$ of a solution.

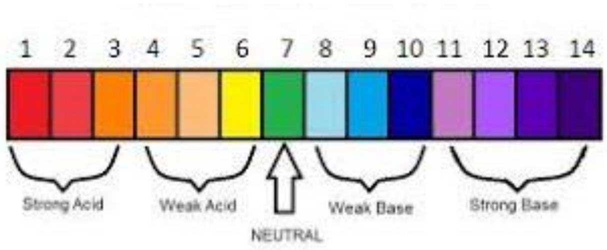
ALKALIS

- **Weak alkalis**. A weak alkali partially disassociate in water.
- **Strong alkalis** – A strong alkali dissociates completely in water.
- **Dilute alkalis** – is an alkali that is not very strong it has a P^H of 8,9,10…
- **Concentrated alkalis** – The amount of alkali dissolved in $1dm^3$ of a solution.

11.5.1.3 Describe the P^H scale (P^H is an abbreviation “potential hydrogen”).

P^H scale is scale ranging from 0 to 14 showing the degree of acidity and alkalinity.

P^H SCALE



11.5.1.4 **Describe neutrality, acidity and alkalinity in terms of P^H value**

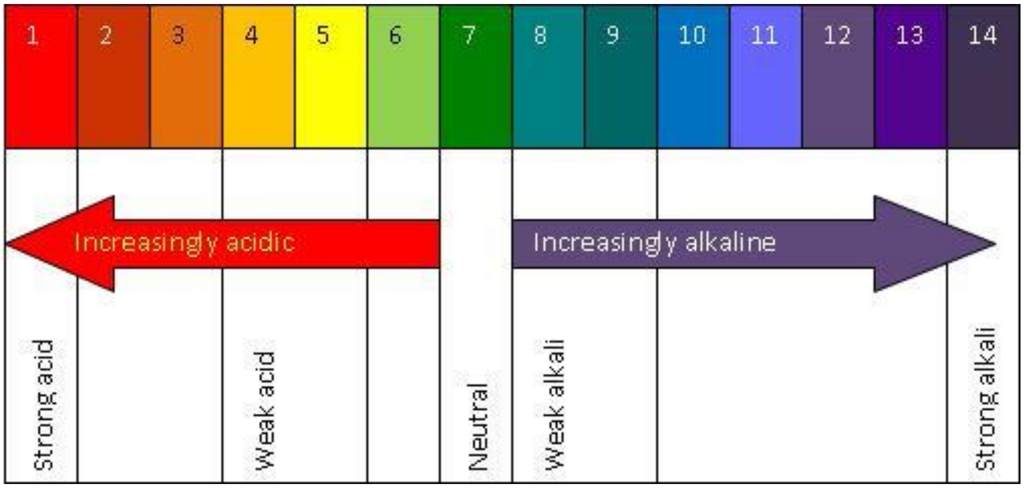
- **Neutrality ;** The P^H of 7 indicates a neutral solution

  (The P^H values: 7 for neutrality)

- **Acidity;** Solution that have a P^H less than 7 are acidic.

  Below 7 for acidity (from 0 to 6.9 is acidity)

- **Alkalinity;** solutions that have a P^H more than 7 are alkaline. ***(**P^H ranging from 8 to 14)*



The P^H scale

11.5.1.5 Determine the P^H value of a solution.

- **Using universal indicator**: which gives different colours at different P^H values.
- **Using P^H meter**: which gives precise values.

11.5.1.6 Demonstrate the characteristic properties of acids

- Acids have a sour taste
- Acids are corrosive (strong acids are dangerous and can burn the skin).
- Acids turn dump blue litmus paper red.

acids reacts with a metal to form a salt and hydrogen gas

Acid + metal ⟶ salt + hydrogen gas

$2HCl_{(aq)} + Mg_{(s)} \longrightarrow MgCl_{2\,(aq)} + H_2$

- **An acid reacts with a base to form salt and water only**,

Acid + base ⟶ salt + water

$HCl_{(aq)} + NaOH_{(AQ)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

- **Acids react with metal carbonate, and hydrogen carbonate to produce salt, water and carbon dioxide**

Acid + carbonate ⟶ Salt + water + carbon dioxide

$2HCl_{(aq)} + CaCO_{3\,(S)} \rightarrow CaCl_{2\,(aq)} + H_2O_{(l)} + CO_{2\,(g)}$

- **State the effect of acids on indicators;**

❖ blue litmus paper turns red
❖ methyl orange change red
❖ Bromethymol blue change yellow
❖ Phenolphthalein change colourles

11.5.1.7 Demonstrate the characteristic properties of bases.

Properties of bases

- Have a bitter taste
- Have a soapy or slippery feel
- Are corrosive
- Are neutralized by acids

11.5.1.8 Illustrate the importance of acid- base reactions

Importance of acid- base reactions

- **Used in controlling the acidity in the soil** – The quality of a crop depends on the quality of the soil. Acidity in soils comes from the build-up of H^+ in the soil solution. For example, acid comes from:
- air pollution that mixes with water vapour and eventually comes back down to earth in the form of rain (in particular, acidic gases from mines and other industrial activities)
- **Treatment of indigestion** – some cells in your stomach produce an acid to help with the digestion of food, while others produce a base to neutralize the acid. Sometimes there is too much acid in the stomach. This acid travels up the esophagus (food pipe) causing heartburn. The symptoms of

heartburn can be relieved by taking antacids (bases or basic salts). Baking soda can also be used to neutralize the excess acid in the stomach.

- **Brushing teeth with toothpaste** – Tooth decay is usually caused by the bacteria in our mouths which converts sugars into lactic acid. It is the build-up of this acid that dissolves the enamel, or outer surface of our teeth. Toothpaste contains sodium bicarbonate (baking soda), which removes stains, provides the foaming action and neutralizes acids.

11.5.1.9State the uses of acids and bases.

- Used in control of P^H in agriculture,
- making of soap,
- in car batteries

11.5.2 PREPARATION OF SALTS

11.5.2.1 Describe a salt

- A **salt is** a compound formed when the hydrogen ions of an acid are fully or partially replaced by a metal or ammonium ions.
- Or **a salt is** compound made of positive metallic/ammonium ions and any negative ion of an acid.

11.5.2.2 Classify salts according to their nature and solubility in water

Types of salts;

- Acid salt
- basic salt
- Normal salts.

Table; solubility rules of salts

SOLUBLE SALTS	**INSOLUBLE SALTS**
All nitrates are soluble	
Carbonates of sodium, potassium, ammonium are soluble (all Carbonates Group I elements)	All other Carbonates are insoluble.
Most sulphate are soluble	Lead (II)sulphate, barium sulphate, calcium sulphate
Most chlorides are soluble	Silver chloride, lead (II) chloride, mercury (I) chloride
Hydroxides and oxides of alkali metals and ammonium	All other oxides and hydroxides are insoluble
All salts of alkali metals and ammonium	

11.5.2.3 Demonstrate the preparation of an insoluble salt.

These are the step for obtaining the Barium sulphate:

Step 1: Mix solutions of barium chloride and magnesium sulphate. A white precipitate of barium sulphate forms at once.

Step 2: The mixture is filtered, barium sulphate gets trapped in the filter paper. (Barium sulphate remains on the filter paper as a residue).

Step 3: It is rinsed with distilled water.

Step 4: Then it is put in a warm oven or porous filter paper to dry.

11.5.2.4 Demonstrate the preparation of a soluble salts.

Step 1: Pipette 25.0cm^3 of 1.0 mol/dm^3 hydrochloric acid into a conical flask.

Step 2: Add 2 droppers of phenolphthalein indicator into the flask and swirl.

Step 3: Slowly, release from the burette aqueous sodium hydroxide of a known concentration into the conical flask and swirl the flask constantly to mix the contents.

Step 4: When the pink colour first appears in the flask, stop the addition of sodium hydroxide and note the volume of alkali used.

Step 5: Repeat the titration steps, this time without the addition of the indicator, adding the noted volume of alkali into the flask of $25.0cm^3$ of 1.0 mol/dm^3acid.

Step 6: The solution formed will be neutral solution of sodium chloride.

Step 7: Heat the solution so that the water will evaporate and the solution become saturated.

Step 8: Crystallize the cooled saturated solution.

Step 9: Filter out the pure sodium chloride crystals.

The steps can be used in the preparation of other salts such as, Zinc sulphate, copper (II) sulphate.

11.5.2.6 Demonstrate the existence of hydrated salts and differentiate from anhydrous salt

- Hydrated salts are salts containing water of crystallization.
- Anhydrous salts do not contain water of crystallization.

11.5.2.7 Describe the behaviour of salts with reference to the atmosphere.

- Salt is hygroscopic, efflorescent, deliquescent.
- Hygroscopic is when a salt absorbs moisture from the atmosphere, but not enough to form a solution.
- Deliquescent is when a salt absorbs enough moisture to form a solution,
- Efflorescent is when a salt gives out the water of crystallization to the atmosphere

11.5.3 TYPES OF OXIDES

11.5.3.1 Describe the various types of oxides.

- **Acidic oxides**; are oxides with acidic properties such as SO_2 and CO_2.
- **Basic oxide**s; are oxides of Group I and Group II metals such as K_2O and MgO.
- **Neutral oxides**; are oxides of some non-metals with neither acidic nor basic properties such as CO, H_2O.
- **Amphoteric oxides**; are oxides with both acidic and basic properties such as ZnO, Al_2O_3 and PbO.

11.5.4.1 Demonstrate the identity of aqueous cations and anions.

Test for aqueous cations

- **Using aqueous sodium hydroxide solution and aqueous ammonia.**

Table: Test for cations

Cations	**Effect of aqueous sodium hydroxide solution**	**Effect of aqueous ammonia**
Aluminium ions (Al^{3+})	White ppt formed, soluble in excess giving a colourless solution	White ppt formed, insoluble in excess aqueous ammonia
Ammonium ions (NH_4^+)	Ammonia gas produced on warming	-
Calcium ions (Ca^{2+})	White ppt formed, insoluble in excess sodium hydroxide	No change
Copper ions (Cu^{2+})	Light blue ppt formed., insoluble in excess sodium hydroxide	Light blue ppt., soluble in excess, giving a dark blue solution
Iron(II) ions (Fe^{2+})	Green ppt formed, insoluble in excess sodium hydroxide	Green ppt., insoluble in excess, turns reddish-brown on standing
Iron (III) ions (Fe^{3+})	Red-brown ppt.formed, insoluble in excess	Red-brown ppt., insoluble in excess
Zinc ions (Zn^{2+})	White ppt formed, soluble in excess giving a colourless solution	White ppt. soluble in excess giving a colourless solution.

Anions using various reagents.

- carbonate,
- chloride,
- iodide,
- nitrate
- sulphate

Table: Test for anions

Anions	**Test**	**Test result**
Carbonate (CO_3^{2-})	Add dilute acid	Effervescence occurs, carbon dioxide gas produced
Chloride (Cl^-) [in solution]	Acidify with dilute nitric acid, then add aqueous silver nitrate.	White ppt formed.
Iodide (I^-) [in solution]	Acidify with dilute nitric acid, then add aqueous lead (II) nitrate	Yellow ppt formed.
Nitrate (NO_3^-) [in solution]	Add aqueous sodium hydroxide, then aluminium foil, warm carefully.	Ammonia gas produced
Sulphate (SO_4^{2-}) [in solution]	Acidify with dilute nitric acid, then add aqueous barium nitrate	White ppt formed

11.5.4 .1 Demonstrate the identity of gases.

Gases being ammonia,

- Place damp universal indicator paper in the gas
- It will turn blue (ammonia gas is alkaline).
- A sharp, pungent smell of ammonia is notice.

Carbon dioxide,

- bubble the gas through limewater(an aqueous solution of calcium hydroxide),
- The limewater will turn milky (CO_2 is slightly acidic, so it reacts with $Ca(OH)_2$ to produce a white precipitate of $CaCO_3$).

Chlorine gas,

- place damp universal indicator paper in the gas
- The indicator paper will first turn red (because chlorine gas is the only gas that has a bleaching effect).

Hydrogen gas,

- Collect some hydrogen gas in a test tube, then hold a burning match stick in the mouth of the test tube

- Hydrogen gas puts off a burning flame with a ‘pop’ sound.

Oxygen gas

- Collect some of the gas in a test tube, then hold a glowing or burning splint in the mouth of the test tube.
- The glowing or burning splint will re-lights or re-kindles

Sulphur dioxide.

- Bubble the gas through an acidified potassium (VI) dichromate ($K_2Cr_2O_7$) solution
- The orange solution will react with $K_2Cr_2O_7$ to produce the green precipitate, $Cr_2(SO_4)_3$.

Table; Test for gases

Gas	**Test**	**Test result**
Ammonia (NH_3)	Introduce damp red litmus paper to the gas	Turns damp red litmus paper blue
Carbon dioxide (CO_2)	Bubble the gas through limewater	White precipitate formed
Chlorine (Cl_2)	Introduce damp blue litmus paper to the gas	Turns litmus paper red then bleaches it
Hydrogen (H_2)	Introduce a lighted splint into the gas	Puts out the lighted splint with a ‘pop’ sound
Oxygen (O_2)	Introduce a glowing splint into the gas	Glowing splint relighted
Sulphur dioxide (SO_2)	Bubble the gas through acidified potassium dichromate (VI)	Turns orange potassium dichromate green.

11.6 THE MOLE CONCEPT

11.6.1 RELATIVE MASSES

11.2.1.1 Describe Relative Atomic Mass and relative molecular mass

- **Relative Atomic Mass**- is the average mass of one atom of an element compared to the mass of one twelfth of one atom of carbon -12.
- **Relative molecular mass-** is the average mass of one molecule of a substance compared to the mass of one twelfth of one atom of carbon-12.

11.6.1.2 Calculate the relative formula mass of a compound

- Relative formula mass- is the sum of all the relative atomic masses of the atoms that make up

  Example. Calculate the relative formula mass of calcium carbonate $CaCO_3$

Answer

Ar × atoms = mass

$Ca = 40 \times 1 = 40$

$C = 12 \times 1 = 12$

$O = 16 \times 3 = \underline{48}$

$\underline{100}$

Therefore, Mr. ($CaCO_3$) = 100.

11.6.2 THE MOLE

11.6.2.1 DESCRIBE A MOLE.

- Amount of substance containing the same number of particles as the number of carbon-12 atoms in exactly 12g of carbon -12 or
- Amount of substance containing 6.02×10^{23} particles.

11.6.2.2 Determine the physical masses (m) of any substance using the molar mass (Mr) and the physical volume (v) of any gas at r.t.p (room temperature and pressure) and vice versa.

- Apply n = mass/Mr and n = volume/ Vm where n = number of moles

(a) (i) find the number of moles in 20g of calcium.

(ii) Find the mass of 3.5 moles of water

Answer

(a)(i) Data n =? ; Ar =40; m = 20g

n = mass/Ar

= 20g/40g/mol

$\underline{=0.5\text{mol Ca)}}$

(ii) Data n =3.5mol; Mr = 18 g/mol; m =?

m= n X Mr

= 3.5mol X 18g/mol

= 63g H_2O

(b) Calculate the number of moles in

(i) 3.01 X 10^{23} atoms of Na.

Answer

Data n=? ; Particles = 3.01 X 10^{23} L = 6.02 X 10^{23}/mol

n=3.01 x 10^{23}/6.02 x 10^{23}/mol

= 0.5 mol Na atoms.

(c) A gas cylinder has a capacity of $200cm^3$.

Calculate;

(i) the moles of CO2 that can fully occupy it at r.t.p

(ii) the moles of CO2 that can fully occupy it at s.t.p

Answer

(i) Data n=? v=$200cm^3$ (200/1000 =0.2 dm^3); Vm = $24dm^3$/mol

n= V/Vm

= $0.2dm^3/24dm^3$/mol

= 0.008 mol CO_2

(ii) Data n =? ; V= $200cm^3$ (200/1000 = 0.2 dm^3); Vm = $22.4dm^3$/mol

n = V/Vm

= 0.2 $dm^3/22.4dm^3$/mol

= 0.0089mol.

11.6.2.3 Describe the relationship of Avogadro's law to reacting moles and volumes of gases at r.t.p and s.t.p.

- The Molar gas volume (Vm) of any gas at r.t.p is $24dm^3$ or 22.4 dm^3 at s.t.p.

- Avogadro's Law states that equal volumes of gases under the same conditions of temperature and pressure contain the same number of molecules.
- This means equal amounts of moles of gases occupy the same volume under the same conditions of temperature and pressure.
- So the volumes have equal moles of separate particles (molecules or individual atoms) in them.
- Therefore, one mole of any gas (formula mass in g), at the same temperature and pressure occupies the same volume.
- This is 24 dm^3 (24 litres) or $24000cm^3$, at room temperature of 25^oC (298K) and normal pressure of 101.3 kPa (1 atmosphere), such conditions are often referred to as r.t.p.
- The molar volume for s.t.p is 22.4 dm^3 (22.4 litres) at 0^oC (273K) and 1atmosphere pressure.

11.6.2.4 Determine the concentration of a solution and apply dilution law.

(Concentration as mol/dm^3 or g/dm^3)

- Concentration: it is amount of substance dissolved in a unit volume of a solution.
- Concentration: It is the amount of substance per unit volume of a solution.
- Concentration can be expressed in two ways:
  - (i) Mass concentration and
  - (ii) Mole concentration.

(i**) Mass concentration**: it is the mass of solute per unit volume of solution .It, therefore, follows:

Mass concentration (Mass. conc.) = mass of solute (g)/volume of solution (dm^3).The units are g/dm^3.

Example

(i) Calculate the concentration of the solution containing 20g of sodium hydroxide in $250cm^3$ of the solution.

Answer

Volume=$250cm^3$ convert to dm^3 by dividing by 1000 giving $0.25dm^3$,then substitute variables in the formula .

Mass.conc = mass/volume

= $20g/0.25dm^3$

$= 80g/dm^3$

(ii) **Mole concentration (molarity)**: it is moles of solute per unit volume of a solution. It, therefore, follows:

Molarity = moles (mol)/volume (dm^3). The units are mol/dm^3.

Mol/dm^3 = molar (M) hence 1 mol/dm^3 = 1M

Example

(i)What is the concentration of a solution made d by dissolving 60 g of sodium chloride (NaCl) in half a litre of a solution?

Answer

Convent the mass of sodium chloride into moles.

n=1.0256mol

Therefore, Molarity = moles/volume

$= 1.0256mol/0.5dm^3$

$= 2.0513mol/dm^3$

Note; The number of moles of solute before dilution is the same as after dilution

$M_1V_1 = M_2V_2$.

11.6.2.5. Illustrate calculations involving stoichiometric reacting moles and volumes of gases and solutions.

- Using molar mass and molar volume of a gas using the mole concept. (Questions on gas laws and conversions of gaseous volumes to different temperatures and pressures will not be required). Proportional stoichiometric masses and the given quantities

11.6.2.6 Describe and calculate the percentage yield in a reaction and the percentage purity of a substance.

Answer

- Percentage yield is the ratio of the actual mass of product to that of expected mass expressed as a percentage. It is calculated as follows;
- ❖ Percentage yield (% yield) = actual mass/expected mass ×100

Example

Consider the reaction between carbon and oxygen

$C (s) + O_2 (g) \rightarrow CO_2 (g)$.

If 60g of carbon was burned in oxygen and 160g of carbon dioxide was produced .what was the percentage yield of carbon dioxide in the experiment.

Answer

First calculate the theoretical mass of Carbon dioxide that can be obtained from the equation

1 mol of C: 1 mol of CO_2

Therefore;

12g/ mol C: 44g/mol CO_2

60g of C: m (CO_2)

m (CO_2) = 220g

Therefore 60g C should yield 220g of CO_2, but only 160g was produced.

Percentage yield (% yield) = actual mass/expected mass × 100

= 160g/220g x 100

= **73%**

$C (s) + O_2 (g) \rightarrow CO_2 (g)$

Actual yield is the actual amount of product produced by a reaction.

Theoretical yield is the expected amount of products produced by the reaction based on calculations involving the amounts of reagents given.

Percentage purity as amount of pure substance divided by total amount of the mixture x 100%

Example

Manganese (IV) oxide reacts with concentrated HCl according to the following equation.

$MnO_2(S) + 4HCl (aq) \rightarrow MnCl_2 (aq) + 2H_2O (l) + Cl_2 (g)$

A 4.325g sample of Manganese (IV) oxide was added to 1.0mol/dm ^{3} HCl. 48cm^3 of the acid was needed to react Manganese (IV) oxide in a given sample. Calculate the percentage purity of Manganese (IV) oxide.

Answer

Number of moles of HC l= 1.0mol/dm^3 x 0.048dm3

= 0.048 mol

The balanced equation;

$MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$

Mole ratio from the equation is,

1mole MnO_2 : 4 moles of HCl

Z : 0.048 moles HCl

Z = 0.048/4

= 0.012 moles MnO_2 (pure), hence 0.012 moles of pure MnO_2 were present.

Mass (pure) MnO_2 = mole x Mr

= 0.012mol x 87g/mol

=1.044g.

% purity = mass of pure sample/mass of mixture x100

= (1.044g/4.325g) x100

= **24%**

11.6.2.7. Determine limiting reagent in a given reaction.

- A limiting reagent is a reactant that is insufficient (hence finishes) in a given reaction.
- It determines the amounts of products formed or reactants used in the reaction.

To determine the limiting reagent using given quantities and stoichiometric proportions.

Step 1: write a balanced equation

Step 2: choose any one product in the equation.

Step 3: convert the given amounts (masses, particles, concentrations or volumes of gases) into moles.

Step 4: calculate the amounts of this product using ratios from the balanced equation and the given quantities.

Step 5: compare the amounts of this product formed

Step 6: reagent that produces the least (smallest) amount is the limiting reagent

- ❖ Limiting reagent always produces the least amount of a product.
- ❖ Use the limiting reagent to calculate the amounts of reactants used or product formed.

Example

Silicon is a semi-conductor that is used in computer chips and solar cell. It is made by the following reaction:

$SiCl_4$ (l) +2Mg(s) → Si(s) +2$MgCl_2$(s)

If 225g of $SiCl_4$, and 225g of magnesium is used to prepare the silicon, which reagent is the limiting reagent.

Answer

Step 1: $SiCl_4$ (l) +2Mg(s) → Si(s) +2$MgCl_2$(s)

Step 2: Chosen product Si

Step 3: converting given masses into moles

Moles of Mg = mass/Ar	moles of $SiCl_4$ = mass/Mr
=225g/24g/mol	=225g/170g/mol
= 9.4mol Mg	= 1.3mol $SiCl_4$

Step3: amount (moles) of Si produced

moles of Si by Mg	moles of Si by $SiCl_4$
from balanced equation:	from balanced equation
2mol Mg: 1mol Si	1mol $SiCl_4$:1mol Si
9.4mol Mg: χ	1.3mol $SiCl_4$: χ

$\chi = 4.7$mol Si

$\chi = 1.3$mol Si

Step 4: comparing amounts of Si produced by Mg and $SiCl_4$, $SiCl_4$ has produced smallest or the least amount of Si hence it is the limiting reagent in this reaction.

11.6.2.8 Demonstrate **calculations involving different types of acid–base titration reactions.**

- Titration is a procedure that is often used to determine the concentration of a solution by reacting with a solution of known concentration.
- Solution of known concentration is called a standard solution.

Example

25cm^3 of hydrochloric acid of concentration 0.12 mol/dm^3 reacts with 28.4 cm^3 of sodium hydroxide solution to form water and sodium chloride. Calculate the concentration of sodium hydroxide solution.

Answer

$HCl (aq) + NaOH (aq) \rightarrow NaCl (aq) + H_20 (l)$

Data

$n_a= 1; n_b =1$ (moles of acid and base in balanced equation)

$C_a=0.12 mol/dm^3$; $C_b = ?$ (Concentrations of acid and base)

$V_a =25 cm^3$; $V_b =28.4 cm^3$ (check that both volumes have the same units)

$C_a \times V_a / C_b \times V_b = n_a / n_b$

$0.12 mol/dm^3 \times 25cm^3 / C_b \times 28.4cm^3 = 1/1$

$C_b = 0.12 mol/dm^3 \times 25cm^3 \times 1 / 28.4cm^3 \times 1$

$C_b = 0.11 mol/dm^3$

11.7.1 EMPIRICAL AND MOLECULAR FORMULAE.

11.7.1.1 Calculate the percentage composition of elements in a compound.

- Percentage composition of an element is relative mass of an element divided the relative formula mass of the compound x 100%. i.e.
- $\% \text{ element} = \frac{\text{mass of element}}{\text{Formula mass}} \times 100\%$

Example

Calculate the percentage composition of sulphuric acid (H_2SO_4)

H =1 x 2 =2

S = 32 x 1 =32

O =16 x 4 = $\underline{64}$

$\underline{98}$

%H = 2/98 x 100 = 2.04%

%S = 32/98 x 100 =32.65%

%O =64/98 x 100 =65.31%

11.7.1.2 Determine the empirical formula of a compound given the molecular formula

- **Molecular formula (M.F):** formula that shows the actual number of atoms in a compound.
- **Empirical formula(E.F) :** simplest formula showing atoms in a molecular formula in their simplest ratio
- Numbers in **M.F** are divided by a common factor (or divisor) to give **E.F.**

Example

Write the empirical formula for the following molecular formulae:

a) C_2H_6
b) O_2
c) $C_6H_{12}O_6$

Answers

Molecular formula (M.F) ----------------- Empirical Formula (E.F)

a) C_2H_6 -- CH_3
b) O_2 --- O
c) $C_6H_{12}O_6$ --- CH_2O

❖ Molecular formula and Empirical formula are related by expression

M.F/n = E.F or

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

11.7.1.3. Determine the empirical and molecular formulae using percentage composition or masses.

- Given percentages or masses of elements empirical and molecular formulae can be determined as follows:

Step 1: write the elements present

Step 2: write the percentages (or masses) under the respective elements.

Step 3: divide percentages (masses) by the relative atomic masses of the elements. This is to convert them into moles.

Step 4: divide the moles throughout by the smallest number of them all.

Step 5: round off to whole numbers.

Step 6: write the empirical formula

Step 7: if relative formula mass is given divide the M.F mass by the E.F mass to obtain the divisor (n)

n = Molecular formula mass (M.F)/ Empirical formula mass (E.F)

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

Example

A compound contains 80% carbon and 20% Hydrogen. The relative formula mass (Mr) is 30. Calculate,

a) The empirical formula.
b) The molecular formula.

Elements:	C	H
Percentages (or masses):	80%	20%
Divide by Ar of element:	80%/12	20%/1
	6.6667mol	20mol
Divide all by smallest:	6.6667/6.6667	20/6.6667
	1	3

Therefore, E.F = CH_3

n = (M.Fmass)/ (E.Fmass)

= 30/15

= 2

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

= $(CH_3)_2$

= C_2H_6

11.7. CHEMICAL REACTIONS

11.7.1. DEFINITION OF A CHEMICAL REACTION

- **Chemical reaction** is a chemical change that occurs when two or more substances combine to form a new substance.

Example of a chemical reaction

$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{(aq)} + H_{2\,(g)}$

Reactants: $Mg_{(s)} + 2HCl_{(aq)}$ and Products: $MgCl_{(aq)} + H_{2\,(g)}$

11.7.2. Rates of chemical reactions

11.7.2.1 Describe rate of a chemical reaction

- **The rate of reaction** is the measure of how long the chemical reaction will take place. Or it is the change in the concentration of reactants or products in a given period of time.

Calculations on the rate of chemical reaction

Reaction Rate = Change of amount of substance/ Time taken. Or

= Change in mass or volume or concentration/ Change in tim**e**

Example: Determine the rate of chemical reaction, if 30.0 dm^3 of Hydrogen gas was produced during the steam reforming of Methane in 5 minutes.

Data

V= $0.25cm^3$ Rate of Chemical reaction = Volume/ Time taken

t = 5.0 s = $30.0cm^3/5.0min$

= <u>$6.0cm^3/min.$</u>

11.7.1.2. Demonstrate the factors that affect the rates of chemical reactions.

Factors that affect the rate of reactions are; temperature, concentration, surface area, catalyst, pressure and light intensity.

- **Temperatur**e

At higher temperatures, the reactant particles will have greater kinetic energy, resulting in a higher speed of movement and more frequent effective collisions. Hence, the increase in temperature of reaction will lead to a higher speed of reaction.

Example

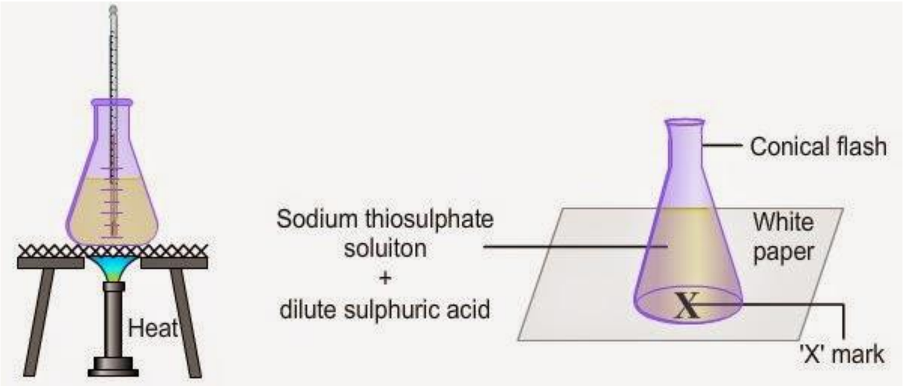
Reaction of sodium thiosulphate and dilute hydrochloric acid:

Chemical equation:

$Na_2S_2O_3 (aq) + 2HCl (aq) \rightarrow 2NaCl (aq) + S(s) + H_2O (l) + SO_{2 (g)}$

Procedure:

- An experiment is set up as shown below*:*



Figure; Showing effect of temperature on speed of reaction

50cm3 of sodium thiosulphate was put into a conical flask, then 1.0 m0l/ dm^3 of dilute Hydrochloric acid was added to sodium thiosulphate in the conical flask, as shown in the diagram above, with a blue cross mark underneath the flask. After sometimes, the blue mark cross had to disappear. And this shows that the chemical reaction has taken place.

A stop watch was started at the point of addition of a dilute acid and the time taken for the cross to disappear was recorded and tabulated.

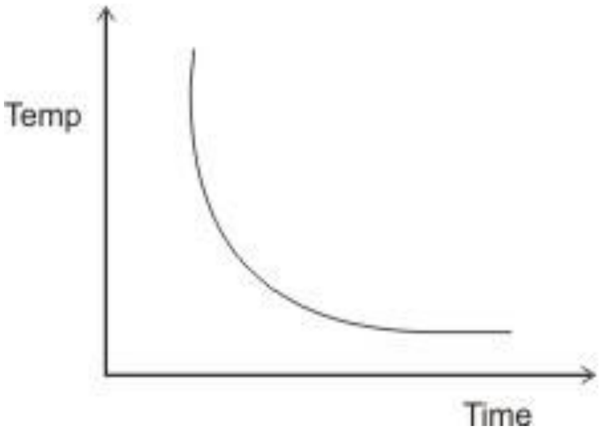
The experiment was repeated for different temperatures and recorded

Data collected

Experiment	**Volume of 0.01 mol/dm3 sodium thiosulphate/cm3**	**Volume of 1 mol/dm3 HCl/cm3**	**Temperature of mixture /°C**	**Time taken for cross to disappear/s**
1	*50*	*5*	*30*	*79*
2	*50*	*5*	*35*	*61*
3	*50*	*5*	*40*	*47*
4	*50*	*5*	*45*	*37*
5	*50*	*5*	*50*	*30*

Data Analysis

Graphical presentation of data collected



Observations & Conclusion:

This means that the higher the temperature, increases the rate of chemical reaction. This also means that the speed of reaction is faster at a higher temperature.

- **Concentration**

An increase in the concentration of one or more of the reactants increases the speed of the chemical reaction. This is due to the fact that there are more particles in a given volume and hence, the frequency of effective collision increases.

Experiment 1

- The chemical reaction of calcium carbonate with hydrochloric acid is altered in the second experiment, keeping all other factors constant.

$CaCO_{3\,(s)} + 2HCl_{(aq)} \rightarrow CaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Experiment: one (1)

Using 2.0g of granules of Calcium Carbonate, reacting with 1.0 mol/dm^3 of dilute hydrochloric acid.

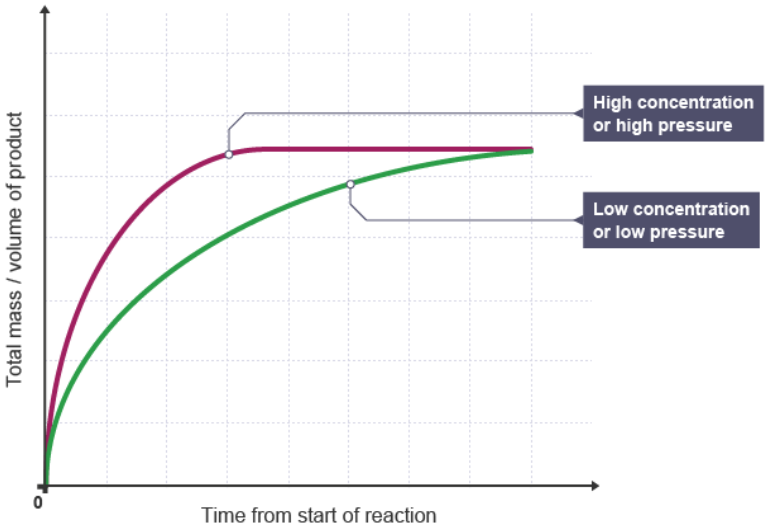
Experiment 2

- Using 2.0g of powder of Calcium Carbonate, reacting with 1.0 mol/dm^3 of dilute hydrochloric acid

Then the volume of carbon dioxide that was given off every after 3 minute's interval for each experiment.

A graph for the two sets of the readings was plotted and interpreted.

Volume of carbon dioxide produced (cm^3) against time (min)



Key

🟪 Experiment 2

🟩 Experiment 1

Observations & Results:

The speed of a chemical reaction will be faster in experiment two (2) than experiment 1, because the concentration of the powdered reactant is higher than that of the granules reactants in experiment one (1).

- **Surface Area (effect of particle size)**

Smaller the size particle, increases the surface area of the reacting particles, hence increasing the rate of chemical reaction.

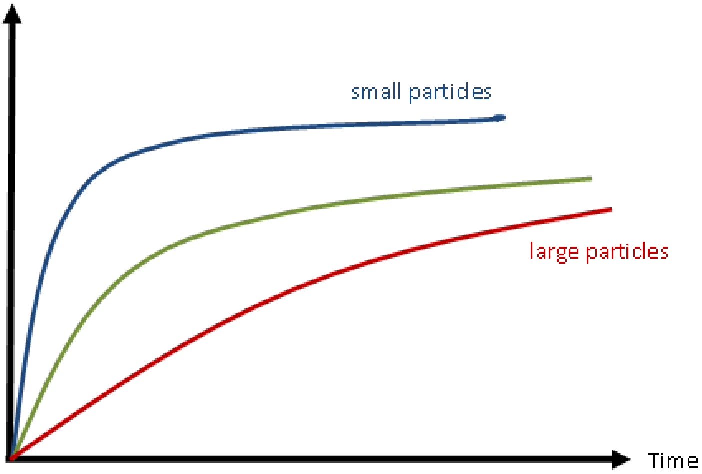
Example

- Reaction of calcium carbonate with hydrochloric acid:

Procedure:

1. Two experiments using the same concentration of hydrochloric acid but with different sizes of calcium carbonate are carried out.

2. The volume of carbon dioxide produced is measured in each experiment and a graph plotted as shown below.



Figure; showing graph of volume of CO_2 per cm^3 against time /min

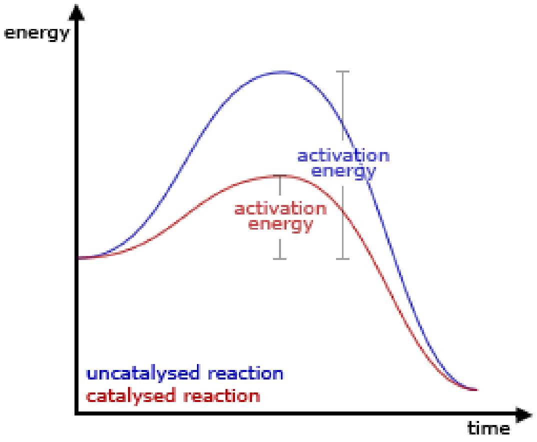
Observations:

The experiment with larger pieces of calcium carbonate was slower in speed while the experiment with the smaller pieces of calcium carbonate was faster.

- **Catalyst**

A catalyst is a substance which changes the rate of reaction, but itself is unchanged chemically at the end of the reaction. A catalyst changes the rate of reaction by lowering the activation energy of a reaction.

Particles react through collisions with each other. When a catalyst is present, the particles can collide with the catalyst as well, causing the particles to react with a lower activation energy.



Figure; showing catalyzed and non –catalyzed reaction

1. Different reactions will require different types of catalysts, i.e. each catalyst is specific to a particular reaction.

2. Enzymes, however are biological catalysts found in living cells. With them, reactions involve the breakdown of giant molecules (such as proteins and starch) into simpler ones like amino acids and sugars can take place.

3. A catalyst does not change the amount of products obtained in a reaction. It only speeds up the chemical reaction. (Note: ΔH is the same for both catalyzed and non-catalyzed reaction).

Example 2

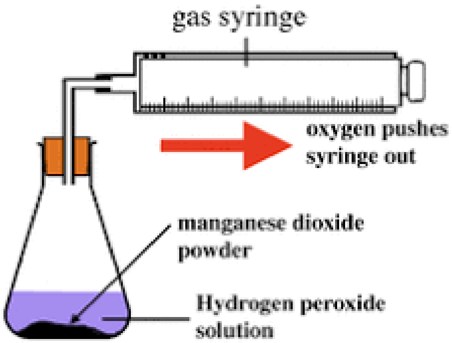
Decomposition of hydrogen peroxide:

Chemical equation:

$2H_2O_{2\ (aq)} \rightarrow 2H_2O_{\ (l)} + O_{\ 2(g)}$

Procedure:

1. The experiment is set up as shown.



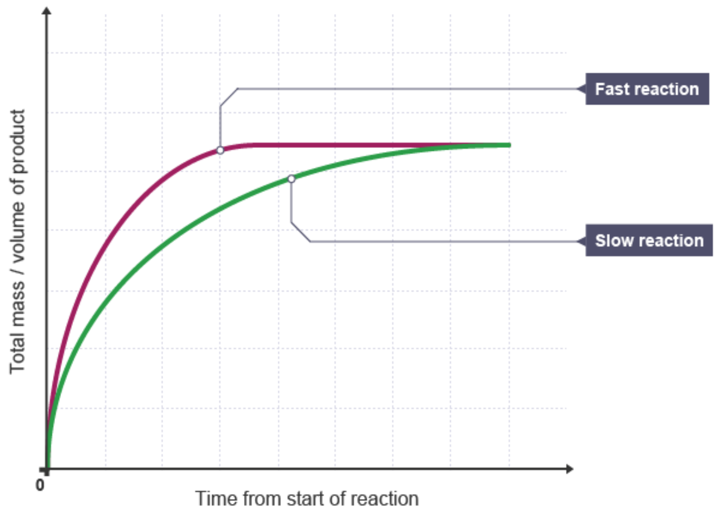
Figure; showing decomposition of hydrogen peroxide using manganese (IV) oxide as catalyst.

2. One experiment was carried out without a catalyst and the other with a catalyst.

<u>Experiment 1</u>	**<u>Experiment 2</u>**
$50cm^3$ of 0.1 mol/dm^3 H_2O_2	$50cm^3$ of 0.1 mol/dm^3 H_2O_2
No catalyst	NbO_2 as catalyst
Room temperature and pressure	Room temperature and pressure

3. The volume of oxygen given out at every 30 s was measured and recorded for both experiments.

11.7.4. The data was collected and a graph for the two experiments was plotted.



Key

--- experiment 2

-- experiment 1

Figure; showing a graph of volume of O2 formed per cm^3 against time/ min

Observations

1. Same volume of oxygen was collected in both experiments.

2. The speed of reaction is higher in experiment 2 as it was a catalyzed reaction.

- **Pressure**

Pressure is the volume of s inversely proportional to the volume of the reacting particle, hence the small the volume, increases the pressure which in turn increases the rate of chemical reaction.
This is explained by the reacting particles being packed into a smaller volume, resulting in more frequent effective collisions between reactants.

- **Light Intensity**

Light (when viewed as an electromagnetic wave) is considered to be an energy source and has sufficient impact energy to break chemical bonds. This energy is more than enough to overcome the activation energy. The greater the intensity or energy of light, the more reactant molecules are likely to gain kinetic energy, so the faster the reaction should be. Methane reacts very slowly with chlorine in the dark, but the rate of reaction is much faster in the presence of ultraviolet light.

11.7.1.4. Describe methods of controlling the rate of chemical reactions.

- Reducing the frequency of collisions between reacting particles such as explosions in flour mills or coal mines when ignited to surface area

11.7.1.5 **Describe the effect of a catalyst on the activation energy**

- Catalyst lowers the activation energy thus increasing the rate of a chemical reaction.

11.7.2. CHEMICAL EQUILIBRIUM

11.7.2.1 Describe what chemical equilibrium is

- Chemical equilibrium:
  - ❖ the rate of the forward reaction equalises with rate of the backward reaction and the concentrations of the substances remain constant. It occurs in reversible reactions.

11.7.2.2 Describe the effect on the position of equilibrium of a reaction upon changing conditions.

- As changes in temperature, pressure, concentration. Apply Le Chartelier's Principle (candidates may **not** be required to state the principle)

NB: Catalysts have no effect on the position of equilibrium

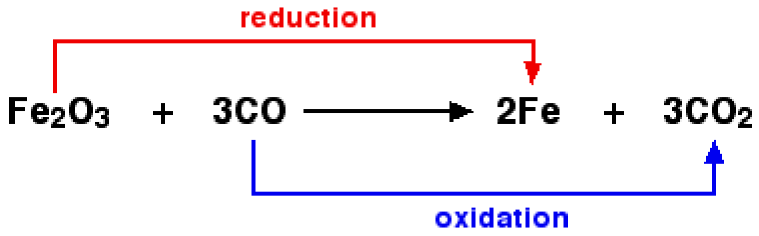
11.7.3 REDOX REACTIONS

11.7.3.1 Describe what oxidation and reduction is

Redox in terms of electron transfer, hydrogen/oxygen transfer, changes in oxidation state.

- Oxidation is gain of oxygen. Reduction is loss of oxygen.

For example, in the extraction of iron from its ore:



$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

Because both ***red***uction and ***ox***idation are going on side-by-side, this is known as a ***redox*** reaction.

Oxidising and reducing agents

An oxidising agent is substance which oxidises something else. In the above example, the iron(III) oxide is the oxidising agent.

A reducing agent reduces something else. In the equation, the carbon monoxide is the reducing agent.

- Oxidising agents give oxygen to another substance.
- Reducing agents remove oxygen from another substance.

Oxiation and reduction in terms of hydrogen transfer

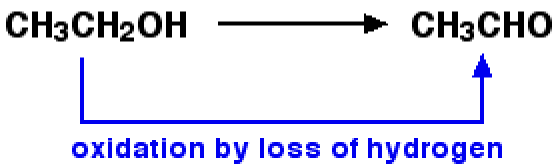
These are old definitions which aren't used very much nowadays. The most likely place you will come across them is in organic chemistry.

Definitions

- Oxidation is loss of hydrogen.
- Reduction is gain of hydrogen.

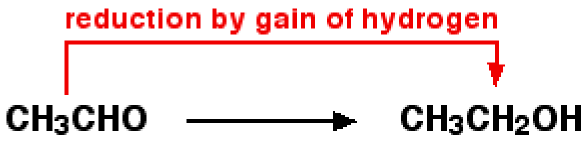
Notice that these are exactly the opposite of the oxygen definitions.

For example, ethanol can be oxidised to ethanal:



You would need to use an oxidising agent to remove the hydrogen from the ethanol. A commonly used oxidising agent is potassium dichromate(VI) solution acidified with dilute sulphuric acid.

Ethanal can also be reduced back to ethanol again by adding hydrogen to it. A possible reducing agent is sodium tetrahydridoborate, $NaBH_4$. Again the equation is too complicated to be worth bothering about at this point.



- Oxidising agents give oxygen to another substance or remove hydrogen from it.
- Reducing agents remove oxygen from another substance or give hydrogen to it.

Oxdation and reduction in terms of electron transfer

This is easily the most important use of the terms oxidation and reduction at A' level.

Definitions

- ***Oxidation is loss of electrons.***
- ***Reduction is gain of electrons.***

It is essential that you remember these definitions. There is a very easy way to do this. As long as you remember that you are talking about electron transfer:

OIL RIG

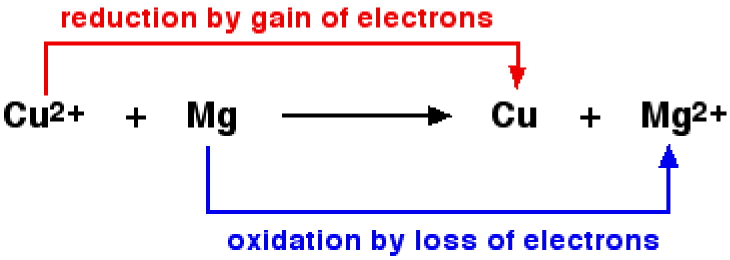
oxidation **i**s **l**oss **r**eduction **i**s **g**ain

A simple example

The equation shows a simple redox reaction which can obviously be described in terms of oxygen transfer.

$$CuO + Mg \longrightarrow Cu + MgO$$

Copper(II) oxide and magnesium oxide are both ionic. The metals obviously aren't. If you rewrite this as an ionic equation, it turns out that the oxide ions are spectator ions and you are left with:



11.7.3.2 Describe what redox reactions is.

- It a reaction that involves both oxidation and reduction reaction.

11.7.3.3 Identify oxidizing and reducing agents in a reaction.

- Oxidizing agent as a reactant that gains electrons and/or reduces oxidation state
- Reducing agent as a reactant that loses electrons and/or increases oxidation state

11.7.3.4 Demonstrate how to determine the oxidation number of an element with variable valency in a compound/ion.

- Determining oxidation numbers using standard rules.

11.7.3.5 Deduce a redox reaction using oxidation numbers.

- Changes in oxidation numbers of reactants and products.

11.7.3.6Describe what non-redox reaction is

  - Non redox reaction, there is neither oxidation nor reduction involved.

11.7.3.7 Identify the characteristics of oxidizing and reducing agents

- Oxidizing agents identified using potassium iodide solution as reducing agent in the presence of starch or starched potassium iodide paper.
- Reducing agents identified using acidified potassium dichromate or potassium permanganate as oxidizing agents and observe colour changes only.

NB: No equations involving potassium dichromate and potassium permanganate will be required.

11.7.4 ENERGETICS OF REACTIONS

11.7.4.1Describe what endothermic and exothermic reactions are

- Types of Energetics of reactions:
  - ❖ As energy in(endothermic)
  - ❖ and energy out(exothermic) reactions

11.7.4.2 Determine a reaction which is endothermic or exothermic.

Identification of endothermic or exothermic reactions: Such as changes in enthalpy, energy level diagrams. Calculating ΔH using bond energies.

1. **Monitor temperature change**

   When energy is released in an exothermic reaction, the temperature of the reaction mixture increases. When energy is absorbed in an endothermic reaction, the temperature decreases. You can monitor changes in temperature by placing a thermometer in the reaction mixture.

2. **Calculate the enthalpy of reaction (ΔH)**

   To classify the net energy output or input of chemical reactions, you can calculate something called the enthalpy change (ΔH) or heat of reaction, which compares the energy of the reactants with the energy of the products.

   Enthalpy is a measure of internal energy. So, when you calculate the difference between the enthalpy of the products and the enthalpy of the reactants, you find the enthalpy change (ΔH), which can be represented mathematically as:

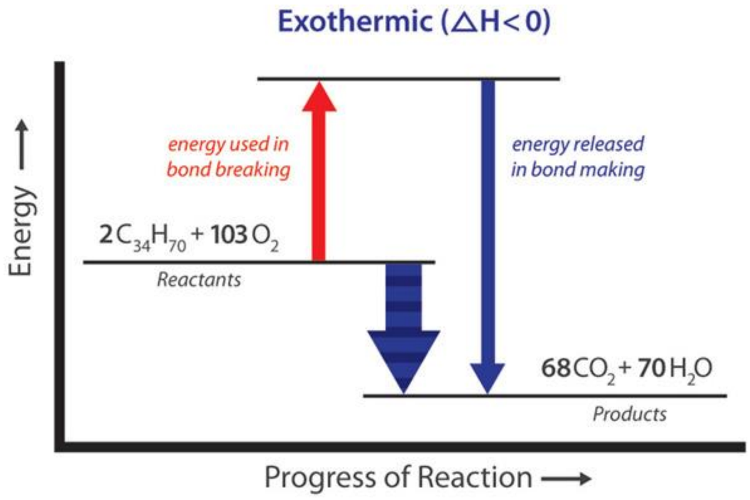
   ΔH = energy used in reactant bond breaking + energy released in product bond making

   Wait, how can you find a difference by adding? The enthalpy values are added in the equation above because, by definition, energy used in reactant bond breaking is always positive (+) and energy released in product bond making is always negative (−).

If ΔH is negative (−) then the chemical reaction is exothermic, because more energy is released when the products are formed than energy is used to break up the reactants. If ΔH is positive (+) then the chemical reaction is endothermic, because less energy is released when the products are formed than the energy is used to break up the reactants.

You can also use energy level diagrams to visualize the energy change during a chemical reaction as a result of the energies used and released according to the above equation for ΔH. To understand these diagrams, compare the energy level of the reactants on the lefthand side with that of the products on the right-hand side.

The graph below charts the energy change when a candle burns. The wax ($C_{34}H_{70}$) combusts in the presence of oxygen (O_2) to yield carbon dioxide (CO_2) and water (H_2O). Because more energy is released when the products are formed than is used to break up the reactants, this reaction is exothermic, and ΔH for the reaction is negative.



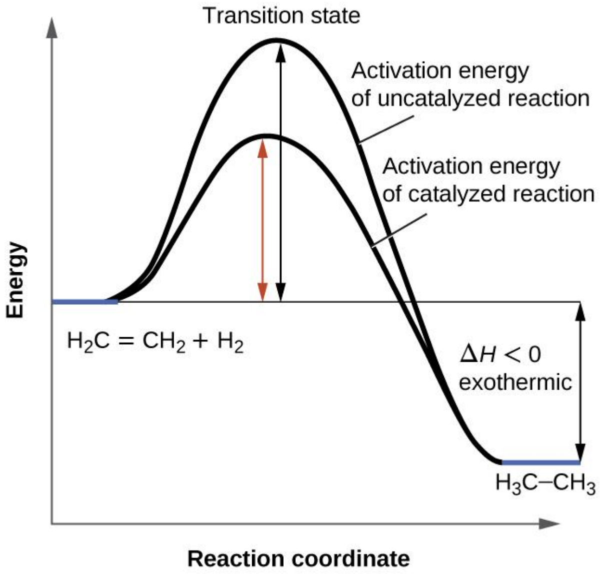
Energy level diagrams.

In this investigation, you will observe whether energy is absorbed or released in two different chemical reactions and categorize them as exothermic and endothermic. You will also explore the relationship between energy changes and chemical reactions.

11.7.4.3 Describe endothermic and exothermic reactions in relation to bonds.

- Endothermic as bond breaking
- and exothermic as bond formation.

11.7.4.4 Identify activation energy for a catalysed and uncatalysed reaction on an energy level diagram



- As in energy level representations: lower for catalysed and higher for uncatalysed.

11.7.4.5 Explain the advantages and disadvantages of energy sources (fuels).

- Advantages of energy sources:
  - ❖ safety,
  - ❖ cost of available reserves,
  - ❖ renewable/
- disadvantages of energy sources:
  - ❖ non-renewable sources

11.7.4.6 Describe the effects of the use of fuels on the environment

- Effects of the use of fuels on the environment:
  - ❖ pollution
  - ❖ greenhouse effect (global warming).

11.7.4.7 Describe the use of silver halide in photography

- Use of silver halide in photography:
  - ❖ As reduction of silver ions to metallic silver by absorption of light.(endothermic reaction)

11.7.4.8 Describe respiration and photosynthesis in terms of energy changes

- Respiration as exothermic process between oxygen and glucose producing carbon dioxide and water
- Photosynthesis as endothermic process between water and carbon dioxide through absorption of light producing glucose and oxygen.

11.7.4.9 Describe use of radioactive isotope in relation to energy changes.

- ❖ As a source of nuclear energy.

11.7.4.10 Explain batteries as convenient source of electrical energy.

- The convenient source of electrical energy:
  - ❖ Batteries are potable.

11.8 THE PERIODIC TABLE

Groups and Periods

11.8.1.1 Describe the Period Table

- The periodic table is a tabular arrangement of chemical elements by order of their atomic numbers.
- It can also be described as a method of classifying chemical elements by predicting the properties of elements.

The periodic table is a way of arranging and classifying elements according to their Atomic numbers. The table consists of rows called **periods** and columns called **groups.** There three other blocks of elements, namely the transition metals, the lanthanides and the actinides. The various section of the periodic table are summarized in the figure below.

Period	GROUP																	
	I	II											III	IV	V	VI	VII	0
1							atomic # → atomic symbol → English element name → Relative atomic mass →	1 **H** hydrogen 1										2 **He** helium 4
2	3 **Li** lithium 7	4 **Be** beryllium 9											5 **B** boron 11	6 **C** carbon 12	7 **N** nitrogen 14	8 **O** oxygen 16	9 **F** fluorine 19	10 **Ne** neon 20
3	11 **Na** sodium 23	12 **Mg** magnesium 24											13 **Al** aluminum 27	14 **Si** silicon 28	15 **P** phosphorus 31	16 **S** sulfur 32	17 **Cl** chlorine 35.5	18 **Ar** argon 40
4	19 **K** potassium 39	20 **Ca** calcium 20	21 **Sc** scandium 45	22 **Ti** titanium 48	23 **V** vanadium 51	24 **Cr** chromium 52	25 **Mn** manganese 55	26 **Fe** iron 56	27 **Co** cobalt 59	28 **Ni** nickel 59	29 **Cu** copper 64	30 **Zn** zinc 65	31 **Ga** gallium 70	32 **Ge** germanium 73	33 **As** arsenic 75	34 **Se** selenium 79	35 **Br** bromine 80	36 **Kr** krypton 84
5	37 **Rb** rubidium 85	38 **Sr** strontium 88	39 **Y** yttrium 89	40 **Zr** zirconium 91	41 **Nb** niobium 93	42 **Mo** molybdenum 96	43 **Tc** technetium 98	44 **Ru** ruthenium 104	45 **Rh** rhodium 103	46 **Pd** palladium 106	47 **Ag** silver 108	48 **Cd** cadmium 112	49 **In** indium 115	50 **Sn** tin 119	51 **Sb** antimony 122	52 **Te** tellurium 128	53 **I** iodine 127	54 **Xe** xenon 131
6	55 **Cs** cesium 133	56 **Ba** barium 137	57 **La** lanthanum 139	72 **Hf** hafnium 178	73 **Ta** tantalum 181	74 **W** tungsten 184	75 **Re** rhenium 186	76 **Os** osmium 190	77 **Ir** iridium 192	78 **Pt** platinum 195	79 **Au** gold 197	80 **Hg** mercury 201	81 **Tl** thallium 204	82 **Pb** lead 207	83 **Bi** bismuth 209	84 **Po** polonium 209	85 **At** astatine 210	86 **Rn** radon 222
7	87 **Fr** francium 223	88 **Ra** radium 226	89 **Ac** actinium 227															

lanthanides **(rare earth metals)**	58 **Ce** cerium 140	59 **Pr** praseodymium 141	60 **Nd** neodymium 144	61 **Pm** promethium 145	62 **Sm** samarium 150	63 **Eu** europium 152	64 **Gd** gadolinium 157	65 **Tb** terbium 159	66 **Dy** dysprosium 162	67 **Ho** holmium 165	68 **Er** erbium 167	69 **Tm** thulium 169	70 **Yb** ytterbium 173	71 **Lu** lutetium 175
actinides	90 **Th** thorium 232	91 **Pa** protactinium 231	92 **U** uranium 238	93 **Np** neptunium 237	94 **Pu** plutonium 244	95 **Am** americium 243	96 **Cm** curium 247	97 **Bk** berkelium 247	98 **Cf** californium 251	99 **Es** einsteinium 252	100 **Fm** fermium 257	101 **Md** mendelevium 258	102 **No** nobelium 259	103 **Lr** lawrencium 262

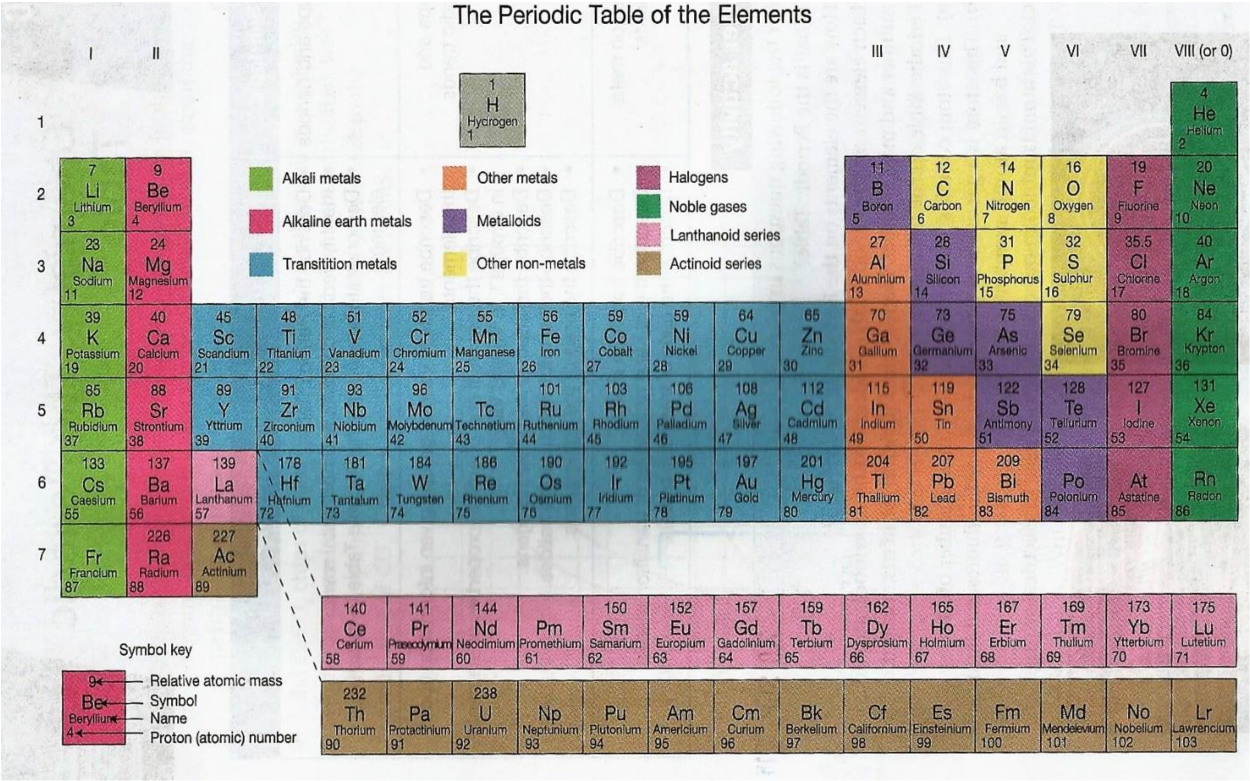
Figure; showing the periodic table of elements

11.8.1.2. Identify vertical columns and horizontal rows.

- **Vertical columns** of elements on the Periodic Table represent **Groups.** There are eight groups on the Periodic Table .The group to which an element belongs depends on the number of electrons in an elements outermost shell.
- **Horizontal rows** of elements on the Periodic Table are **Periods.** There are seven periods on the Periodic Table. The period to which an element belongs depends on the number of occupied shells it has in its electronic structure.

11.8.1.3. Demonstrate how to use the Periodic Table to classify elements.

Elements are classified as non- metals, metals and metalloids. All the metals are found to the left of the thick line in the figure below the non-metals are found on either side of the line as shown below



Figure; shows classification of elements on the periodic table

11.8.2.1. Describe trends in various Groups given information about the elements

GROUPS AND PERIODIC TRENDS

- **GROUPS TRENDS**
  Elements in the same group;
  - Have the same number of electrons in the outermost shell
  - Form ions with similar formulae

- Form compounds with similar chemical properties
- Have similar physical and chemical properties

- **PERIODIC TRENDS**
  - The period to which an element belongs depends on the number of occupied shells it has in its electronic structure.
  - Elements are arranged in order of increasing proton number (atomic number).

11.8.2.2. Describe the physical and chemical properties of elements in Group I, II, VII and VIII.

GROUP I

Group I elements are known as the alkali metals.

Physical properties

- Are soft and easy to cut with a knife
- Are silvery when freshly cut
- Float on water because of their low density
- Have low melting points and boiling points
- Are good conductor of heat and electricity because they are metals
- React spontaneously and vigorously with water and oxygen ,so they must be stored under oil

Chemical properties

- These elements all have one valence electron, so, they react by losing electron to form a single positive ion.
- They are the most reactive metals and their reactivity increase down the group.
- React vigorously with cold water forming H_2 (Hydrogen gas)
- Burn in air with colored flames to form oxides
- Alkali metals react with Halogens producing a neutral salt which dissolve to form a colorless solution e.g. Sodium Chloride

SUMMARY TRENDS.

- Reactivity of elements increases as you go down the group
- Density increases as you go down the group
- Melting and boiling points decreases as you go down the group
- They become softer to cut as you go down the group

Group II

Group II elements are known as alkaline earth metals.

Physical properties

- Are soft and can easily be cut with a knife (although they are slightly harder than the alkali metals) but harder than group one elements.
- Show a silvery color when freshly cut
- Are low- density metals.
- Have higher melting points and boiling points
- Are good conductor of both heat and electricity
- They tarnish quickly when left in air due to formation of an oxide on their surface

Chemical properties

- These elements all have two valence electrons, so, they react by losing two electrons to form an ion with a charge $2^{+.}$
- Their reactivity increase down the group, but they are less reactive than the alkali metals.
- They react with cold water but less vigorously than alkali metals

SUMMARY TRENDS.

- They all have two electrons in their outer most shell
- Reactivity increases as you go down the group
- Melting and boiling points decrease as you go down the group.

Group VII

Group seven elements are also known as salt Halogens. Halogen is derived from Greek meaning salt maker. When halogens gain electrons they form ions known as Halides.

They all have seven electrons in their outer most shell.

Physical properties

- They are non-metals
- Exist as diatomic molecules e.g. F_{2}, Cl_{2}, Br_{2} and I_{2}.
- Have a high melting and boiling point
- Are poisonous
- Are poor conductors of both heat and electricity
- These are colored elements and they darken as you go down the group

Chemical properties

- These elements all have seven valence electrons. So they react by gaining one electron to form a single negatively charged ion.
- Their reactivity decrease as you go down the group

- Halogens undergo displacement reactions; the more reactive halogen will displace the less reactive halide ion from its salt in a solution. Reactivity decreases as you go down the group
- Melting and boiling points increase as you go down the group.

11.8.2.3. Describe the importance of halogens

Importance of halogens

- Fluorine is used as fluoride in toothpaste,
- chlorine in water treatment, antiseptic,
- bromide in photographic film
- Small amounts of iodine are needed in our bodies to prevent the swelling of the thyroid gland.
- Chlorine is used to bleach wood pulp to make white paper. (used as bleaching agent)

11.8.2.4. Describe the harmful effects of halides.

- When used in drugs may have negative effect on man,
- Used in Pesticides which has negative effect on environment when used in excess,
- CFCs cause ozone layer depletion (CFCs)

11.8.2.5. Describe the use of the noble gases

- providing an inert atmosphere in bulbs
- The significance of their non- reactivity in providing an inert atmosphere. Such as argon in electrical lamps, helium in balloons

Transition metals

11.8.3.1. Describe transition metals.

Answer

- Transition metals are a block elements between Group II and Group III of the Periodic Table

11.8.3.2. Describe general properties of transition metals.

General properties of transition metals.

- have variable valencies,
- high densities,
- high melting points,
- form coloured compounds,
- Are catalysts.

Note: Electronic configuration of transition metals will **not** be required

11.8.3.3. Describe the uses of transition metals

Answer

Uses of transition metals

They are used;

- as catalysts in industries in order to speed up reactions
  E.g. an iron catalysis is used in the manufacturing of ammonia.
  A nickel catalyst is used in the manufacturing of margarine
- to make alloys, because they combine readily with themselves and other elements e.g.
  - mirrors are made with copper –tin alloy.
  - Brass is a copper – zinc alloy.
  - Steel is an iron and carbon alloy.
- As conductors of electricity and heat e.g.
  - Copper is used in electrical wiring
  - Aluminum is used to make pots and pans for cooking
- To manufacture paint

NB: Heavy metals are no longer used to make paint for health reasons

GRADE 12 TOPICS

12.1 CHEMISTRY AND ELECTRICITY

12.1.1 Conductors
12.1.2 Electrolysis
12.1.3 Simple cells (chemical cell)

12.2 METALS

12.2.1 General properties of a metals
12.2.2 Reactivity and Electro Chemical Series
12.2.3 Alloys
12.2.4 Corrosion

12.3 NON -METALS

12.3.1 General properties of non-metals
12.3.2 Hydrogen
12.3.3. Oxygen
12.3.4 Nitrogen
12.3.5. Chlorine
12.3.6 Sulphur
12.3.7 Carbon and carbonates
12.3.8 Silicon

12.4 ORGANIC CHEMISTRY

12.4.1 Saturated and unsaturated Hydrocarbons
12.4.2 Alcohols (Alkanols)
12.4.3 Carboxylic acids (alkanoic acids)
12.4.4 Esters (Alkanoates)
12.4.5 Homologous series
12.4.6 Macromolecules (Polymers)

12.9 CHEMISTRY AND ELECTRICITY

12.9.1.1 Classify conductors and non- conductors

- Conductors are metals such as copper, aluminium, silver and non-conductors being non-metals such as sulphur, phosphorus, except carbon in form of graphite.

Electrolysis

12.9.2.1 Classify electrolytes and non-electrolytes

Difference between electrolytes and non-electrolytes:

- **Electrolytes** are ionic compounds while **non- electrolytes** are covalent compounds.
- **Electrolytes** conduct electricity and are decomposed while **non-electrolytes** do not conduct electricity and are not decomposed.

12.9.2.2 Describe what electrolysis is

- **Electrolysis**: is the decomposition of an electrolyte using electricity in an electrolytic cell.
- **Voltammeter:** apparatus where electrolysis occurs.
- **Electrolyte** is a compound that conducts electricity and gets decomposed.
- **Electrodes:** piece of metal or graphite where oxidation or reduction takes place.

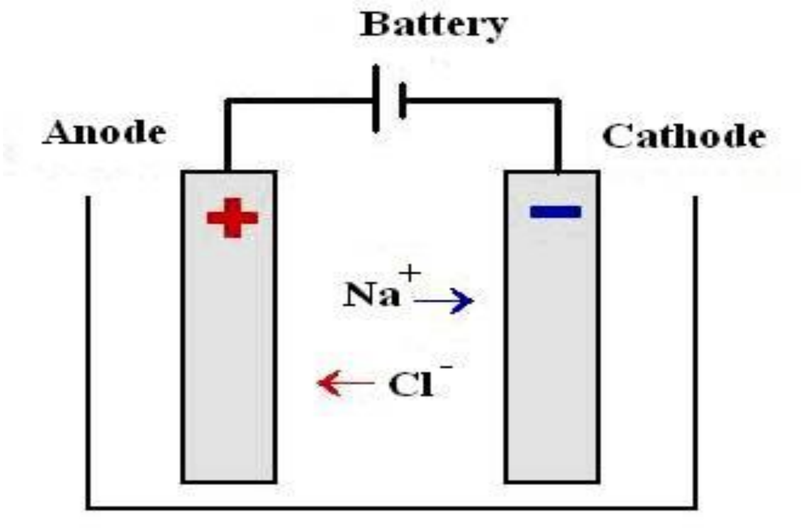
❖ **Anode** (+ve): electrode where oxidation occurs.
❖ **Cathode** (-ve): electrode where reduction occurs.

12.9.2.3 Describe the products at the electrodes during electrolysis of molten binary ionic compounds .

- **Binary compounds** consist of only two elements, **molten** means in liquid state .e.g. molten lead (II) bromide, molten sodium chloride, molten aluminium oxide etc.

- Electrolysis of molten binary ionic compounds always produce:
  - **d)** **Metals** at the cathode by reduction.
  - **e)** **Non- metals** at the anode by oxidation.
- Passing an electric current through molten sodium chloride decomposes the material into sodium metal and chlorine gas. Care must be taken to keep the products separated to prevent the spontaneous formation of sodium chloride.

Electrolysis of molten sodium chloride.



12.9.2.4. Describe the products at the electrodes during electrolysis of aqueous ionic solutions.

- Products formed in the electrolysis of aqueous ionic solutions are predicted by the **selective (preferential) discharge** of ions under the given conditions.
- Electrolytes includes, dilute sulphuric acid (acidified water), concentrated hydrochloric acid, concentrated aqueous sodium chloride (brine) using carbon electrodes, and aqueous copper (II) sulphate using carbon and copper electrodes.

12.9.2.5. Describe the industrial applications of electrolysis

- Applications of electrolysis include:
  - Extraction of metals such as aluminium from its oxide, potassium, sodium, calcium and magnesium from their chlorides.
  - Purification of metals e.g. copper refinery and
  - Electroplating of metals to prevent corrosion and enhancing appearance.

12.9.2.6. Calculate the quantity of electrolytic products.

Quantitative Aspects of Electrolysis

- The amount of product produced at the electrode depends on three factors namely:
  - current (I),
  - time (t) and
  - the charge on the ion.

- Calculations on electrochemistry are based on Faraday's two laws of electrolysis.
- 1st law states that amount of substance produced at the electrode is directly proportional to the quantity of electricity (Q) used.
  Mass (m) ᾱ current (I) x time (t) but current (I) x time (t) = Charge (Q)
- 2nd law states that the amount of substance produced at the electrode is inversely proportional to the charge on the ion when the same quantity of electricity is passed through electrolytes of different charges.
  Mass (m) ᾱ 1/ charge
- Combining the two laws it, therefore, follows that:

Mass (m) **ᾱ** current (I) x time (t) ᾱ 1/ ion charge

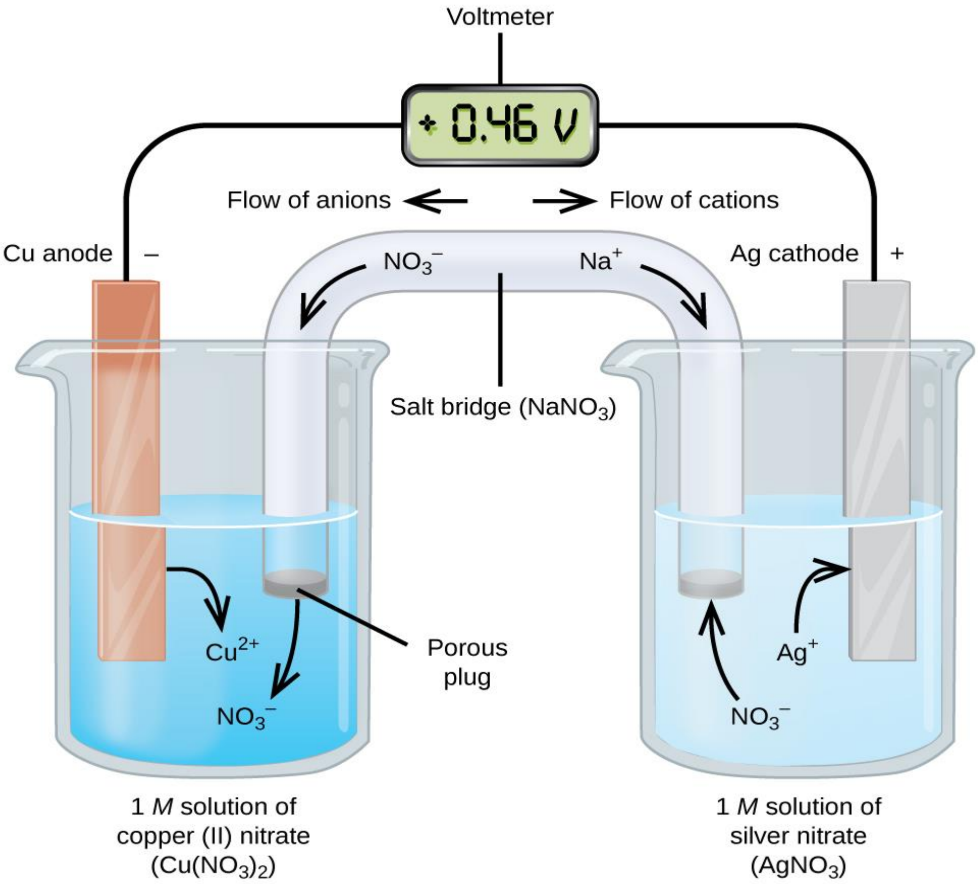
m = Z (I t/ion charge); Z =Ar/F (known as the electrochemical equivalence of the element) where Ar is the relative atomic mass of element and F is the Faraday constant.

Therefore, the mass produced during electrolysis is given by:

m = Ar x I x t /F x ion charge

12.9.3.1. Describe what a chemical cell is

- A chemical cell: consists of two different metals connected together and dipped in an electrolyte to produce electricity.
- Converts chemical energy into electrical energy. Examples include Galvanic and Daniel cells.



12.9.3.2. Compare electrolytic cells and simple cells

d) **Types of cells**: electrolytic cells and electrochemical cells (simple cells).

e) **Similarities**: in both cells, oxidation is at the anode and reduction is at the cathode.

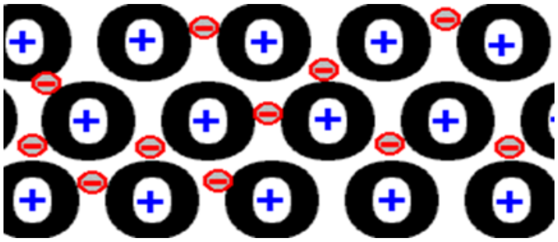
f) **Differences**:

- Cathode is negative in electrolytic cell while positive in simple cell.
- Anode is positive in electrolytic cell while negative in simple cell.
- In electrolytic cells both electrodes (anode and cathode) are dipped in the same electrolyte in the same container while in simple cells the electrodes are in different containers with different electrolytes.
- Electrolytic cells may use same or different material for the electrodes while simple cells always use different materials.

Electrolytic cells consume (use up) electricity while simple cells produce electricity.

12.10 METALS

12.10.1.1 Describe diagrammatic representations of pure metals



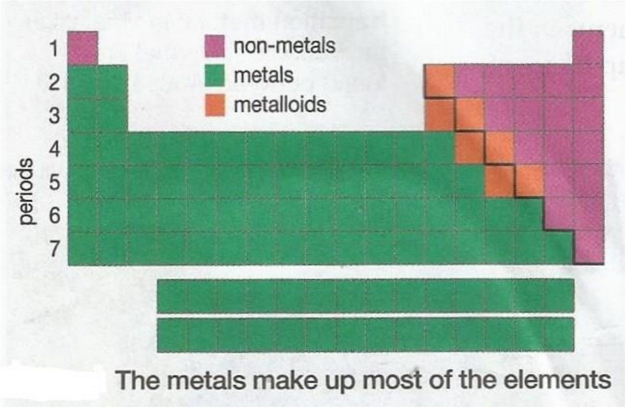
Figure;showing structure of a pure metals

- Metals, atoms are arranged in layers, which allows metals to be bent and shaped without breaking the bonds.
- There is a strong electrical force of attraction between these free electrons (***mobile electrons or 'sea' of delocalised electrons***) (–) and the 'immobile' positive metal ions (+) that form the giant lattice and this is the ***metallic bond***.
- Most metals are good conductors of electricity and heat because the delocalised electrons in the metal carry electrical charge through the metal.

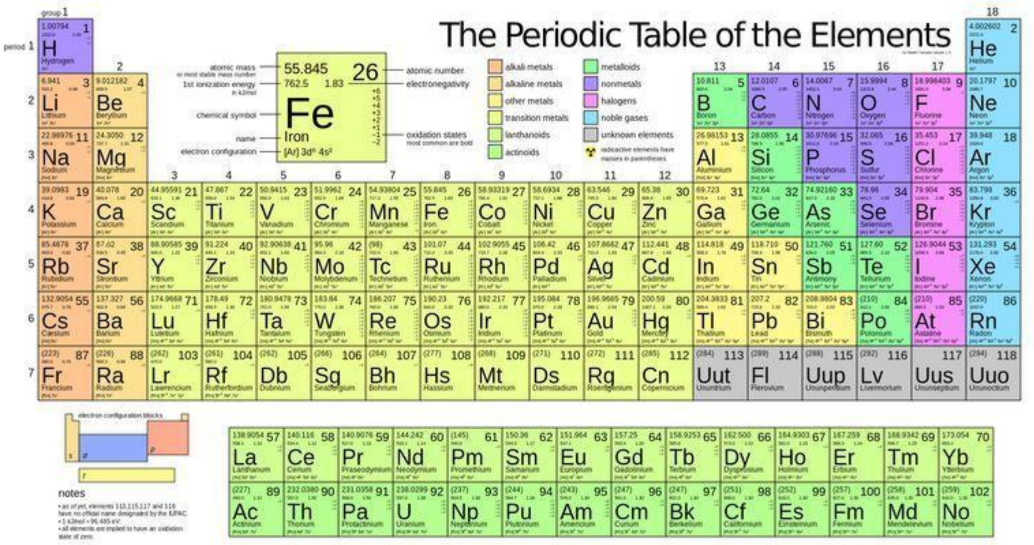
12.10.1.2 Describe the physical properties of metal

- They are hard and strong
- They are malleable (***can be deformed or molded into a different shape***).
- They are ductile (***capable of being drawn out into wire or threads***)
- They are sonorous (***capable of making ringing sound***)
- They are (***lustrous***) shiny but can be easily tarnished by corrosive oxidation in air and water.
- They have high melting and boiling points, though there notable exceptions e.g. Mercury
- Most metals have high densities
- Good conductor of both heat and electricity

Metals, metalloids and non-metals as they appear on the periodic table



The metals make up most of the elements



period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 1																	He 2
2	Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
3	Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
4	K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
5	Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
6	Cs 55	Ba 56	Lu 71	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
7	Fr 87	Ra 88	Lr 103	Rf 104	Db 105	Sg 106	Bh 107	Hs 108	Mt 109	Ds 110	Rg 111	Cn 112	Uut 113	Fl 114	Uup 115	Lv 116	Uus 117	Uuo 118

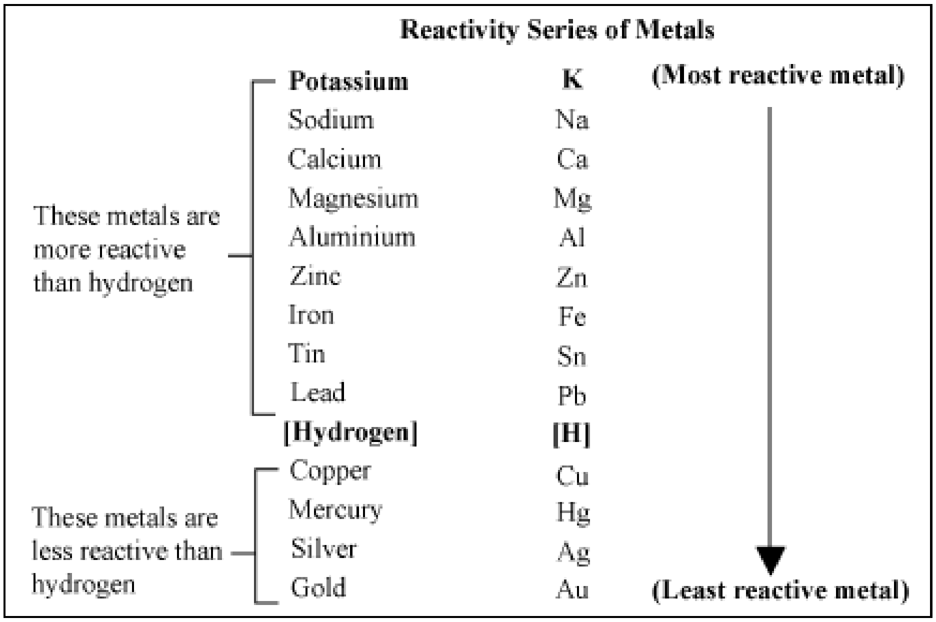
La 57	Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70
Ac 89	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102

12.10.1.3. Describe the chemical properties of metals

- All metals are electropositive (***lose electrons to form positive ions***)
- All metals react with oxygen to form metal oxides
- Metals react with water to produce metal oxide (or metal hydroxide) and hydrogen gas.
- Metals react with a dilute acid to form a metal salt and hydrogen gas.
- All metals react with chlorine to form ionic metal chloride
- Metal are reducing agents (***in a reaction metals are electron donors***)

12.10.2.1 Describe the reactivity series of metals

As arrangement of metals in order of either their increasing or decreasing order of reactivity as being potassium, sodium, calcium, magnesium, aluminum, zinc, ion, lead,)hydrogen), copper and silver.

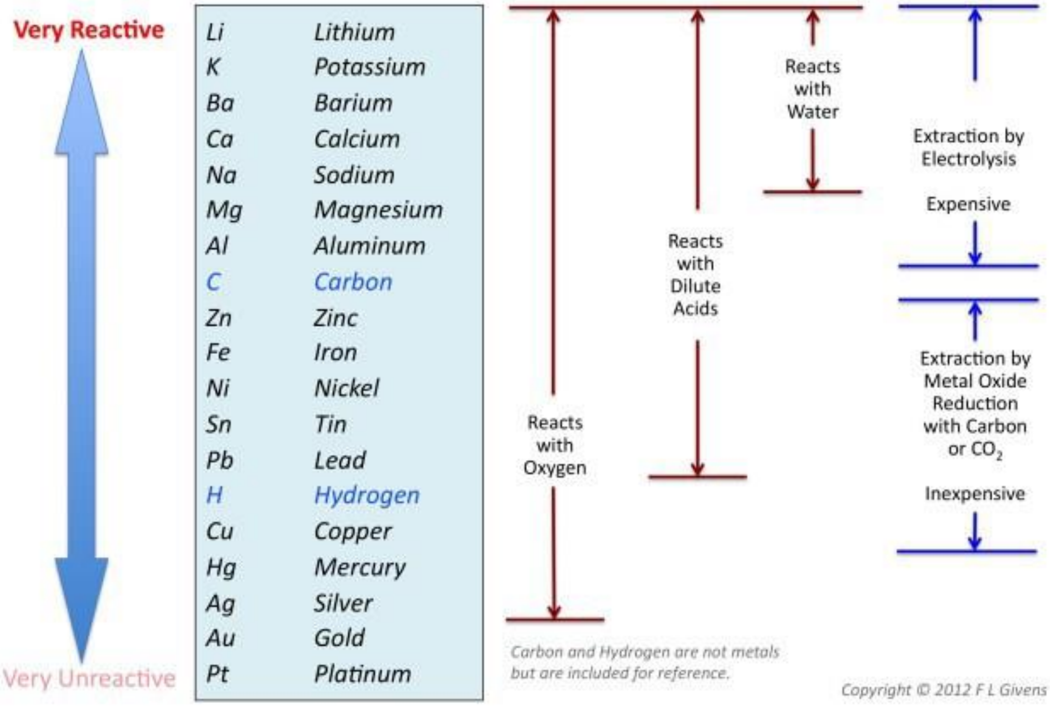


Reactivity Series of Metals

Potassium	**K**	**(Most reactive metal)**
Sodium	Na	
Calcium	Ca	
Magnesium	Mg	
Aluminium	Al	
Zinc	Zn	
Iron	Fe	
Tin	Sn	
Lead	Pb	
[Hydrogen]	**[H]**	
Copper	Cu	
Mercury	Hg	
Silver	Ag	
Gold	Au	**(Least reactive metal)**

These metals are more reactive than hydrogen

These metals are less reactive than hydrogen



Li	*Lithium*
K	*Potassium*
Ba	*Barium*
Ca	*Calcium*
Na	*Sodium*
Mg	*Magnesium*
Al	*Aluminum*
C	*Carbon*
Zn	*Zinc*
Fe	*Iron*
Ni	*Nickel*
Sn	*Tin*
Pb	*Lead*
H	*Hydrogen*
Cu	*Copper*
Hg	*Mercury*
Ag	*Silver*
Au	*Gold*
Pt	*Platinum*

Carbon and Hydrogen are not metals but are included for reference.

Copyright © 2012 F L Givens

12.10.2.2. Explain the apparent non reactivity of aluminium.

The non-reactivity of aluminum is due to formation of thin protective oxide layer, Al_2O_3 on surface of metal which cuts off further reaction.

12.10.2.3 Demonstrate an order of reactivity.

Mnemonics: ***Posocamazic*** as a word assist to remember the reactivity series Po potassium/so sodium/Ca calcium/M magnesium/ A aluminium/Z zinc/ I iron/ C copper.

12.10.2.4 Describe the effects of heat on hydroxides, carbonate, nitrate of metals and ammonium compounds.

Compounds of more reactive metals are difficulty to decompose while compounds of less reactive metals easily decompose

Most nitrates tend to decompose on heating to give the metal oxide, brown fumes of nitrogen dioxide, and oxygen.

Group 2 nitrate like magnesium nitrate decomposes like:

$$2Mg(NO_3)_2(s) \rightarrow 2MgO(s) + 4NO_2(g) + O_2(g)$$

In group 1metals nitrates producing the metal nitrite and oxygen, except lithium nitrate which behave like magnesium nitrate.

$$2XNO_3(s) \rightarrow 2XNO_2(s) + O_2(g)$$

When heated to 512 °C calcium hydroxide into calcium oxide and water.

$$Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(l)$$

Metals above sodium are generally very reactive and form stable carbonates.

Metals below sodium form carbonates that can be decomposed by heat into metal oxide and carbon

$$CaCO_{3(S)} \longrightarrow CaO_{(S)} + CO_{2\,(g)}$$

12.10.2.5 Describe the extraction of copper, iron, aluminium and zinc from their ores.

Extraction of copper, iron, aluminium and zinc: Chemical and electrolytic reduction. Chemical reducing agents being Carbon, carbon monoxide, and hydrogen.

(I) EXTRACTION OF IRON

Iron is widely found in the earth crust but seldom as free element. The common ores of Iron are haematite (Fe_2O_3), magnetite and (Fe_3O_4) and siderite $FeCO_3$.

Extraction takes place in two stages

Stage 1

The roasting of the iron ores to removal of water, carbon dioxide and other impurities of low melting points and form ferric oxide.

$FeCO_3\,(S) \rightarrow FeO\,(s) + CO_2\,(g)$. ***Carbon dioxide evolves in air.***

$4FeO(s) + O_2\,(g) \rightarrow 2Fe_2O_3\,(s)$. ***The ferric oxide is heated in air to produce haematite.***

Stage ii

Reduction of oxide to iron in the blast furnace

The three raw materials are put on top of the blast furnace: the iron oxide (Fe_2O_3), coke (C) and limestone ($CaCO_3$) (***the charge***).

Step 1

Hot air is blown at the bottom of the furnace at the temperature of about 800°C.

The reactions:

Coke burns in air and produce carbon dioxide rises the temperature to about1700°C.

$C\,(s) + O_2\,(g) \rightarrow CO_2\,(g) + \text{heat}$

Further up the furnace carbon dioxide is reduce by carbon (coke) to carbon monoxide.

$CO_2\,(g) + C(s) \rightarrow 2CO\,(g)$

Step 2

Carbon monoxide reduces iron oxide to molten iron and this reaction is reversible.

$Fe_2O_3(s) + 3CO\,(g) \rightleftharpoons 2Fe\,(l) + 3CO_2\,(g)$

$Fe_3O_4(s) + 4CO\,(g) \rightleftharpoons 3Fe\,(l) + 4\,CO_2\,(g)$

The gases leaving the blast furnace contain about 25 percent carbon monoxide, they burn to heat the air blast entering the furnace.

Step 3

The limestone (calcium carbonate) is heated and decomposes to calcium oxide, which combines with silicon dioxide (SiO_2) the impurity found in the ore and forms a slag of calcium silicate.

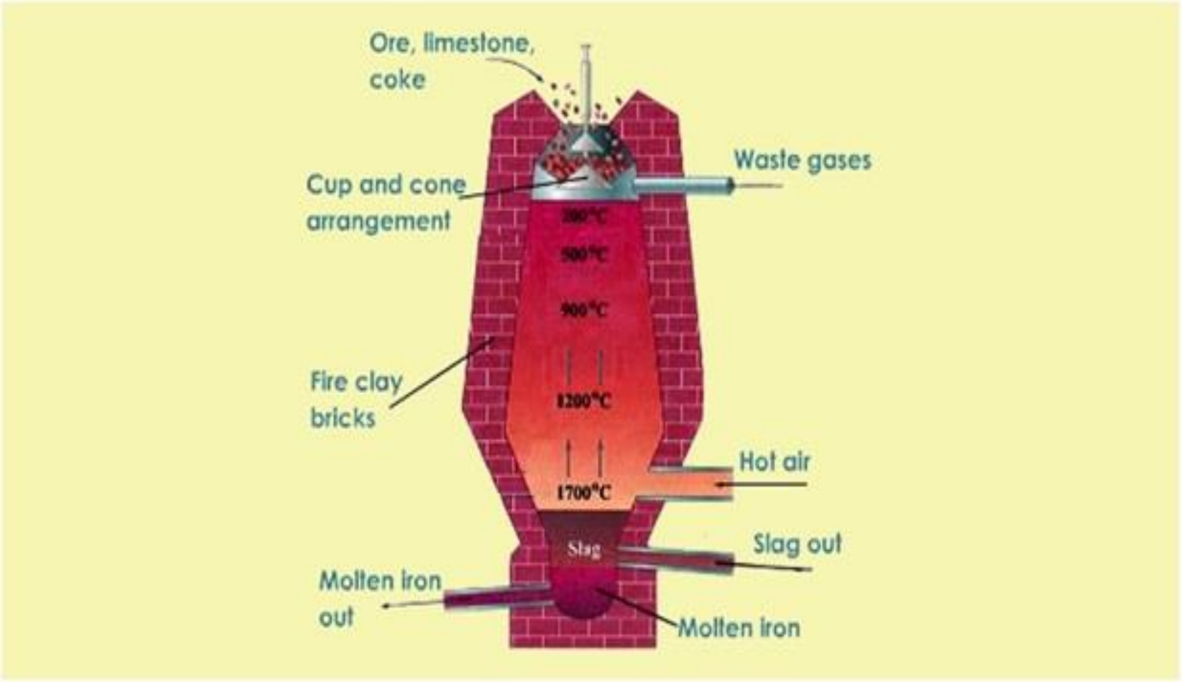
Calcium carbonate (limestone) decomposes to calcium oxide and carbon dioxide

$CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g)$

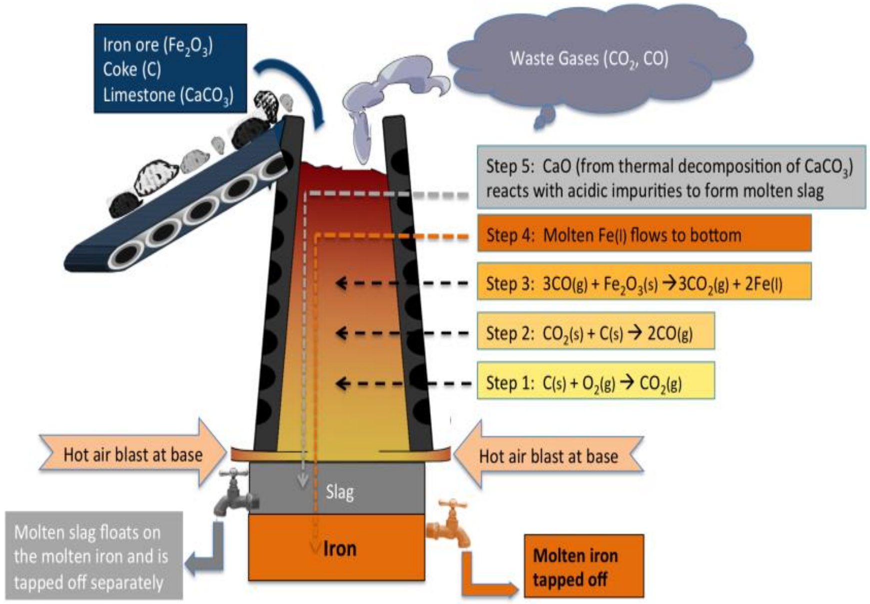
Calcium oxide combines with silicon dioxide to form slag of calcium silicate.

$CaO (s) + SiO2 (s) \rightarrow CaSiO_3 (l)$

THE BLAST FURNACE



A diagram of the summary of the extraction of iron in a blast furnace



The iron trickles to the bottom of the furnace and absorbs some carbon dioxide which lowers is melting point. Both the molten slag and molten iron drop to the bottom of the furnace. The less dense slag floats on top of the iron and prevents the further oxidation of the iron by the hot air blast. The slag and iron are run out by separate holes from time to time. The process is continuous.

(II) EXTRACTION OF ZINC

Zinc is extraction from zinc sulphide or zinc blend. The ore is crushed and put into the tank of water containing chemical frothing agent. Air is blown in the mixture and rocks sink to the bottom of the tank. The zinc sulphide floats on top and it is removed as concentrate.

The concentration ores are first roasted in air to convert them into oxides.

$$2ZnS\ (s) + 3O_2\ (g) \rightarrow 2ZnO\ (s) + 2\ SO_2\ (g).$$

The sulphur dioxide that is produced is used in the manufacturing of sulphuric acid.

$$ZnO\ (s) + C\ (s) \rightarrow Zn\ (g) + CO\ (g)$$

Then zinc oxide is mixed with powdered coke and heated to temperature of about $1400^{o}C$. The coke reduces the zinc oxide to zinc. The zinc vapour and carbon monoxide are passed through the outlet of the furnace where the zinc vapour condenses and solidifies and the carbon monoxide bubbles off.

(III) EXTRACTION OF Copper

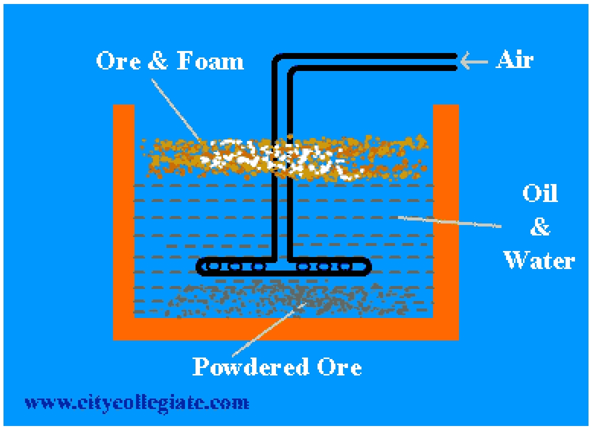
There are three stages of extracting copper

- ***The concentration of the ore to remove the impurities.***
- ***Roasting of the ore and reduction.***
- ***Refining of impure copper.***

Copper is extracted from its common ore copper pyrite ($CuFeS_2$).

Copper pyrite is crushed to powder and mixed with water and oil into the tank.

Air is blown into the mixture, heavy impurities sink to the bottom of the tank.



The concentrate of copper pyrite float as the froth, it is removed and roasted in air.

$$2\ Cu\ FeS_2\ (s) + 4O_2\ (g) \rightarrow Cu_2S\ (g) + 2F\ e\ O\ (g) + 3SO_2\ (g)$$

Sulphur gas escapes in air

Iron oxide is removed by heating it with silicon (IV) oxide in the absence of air.

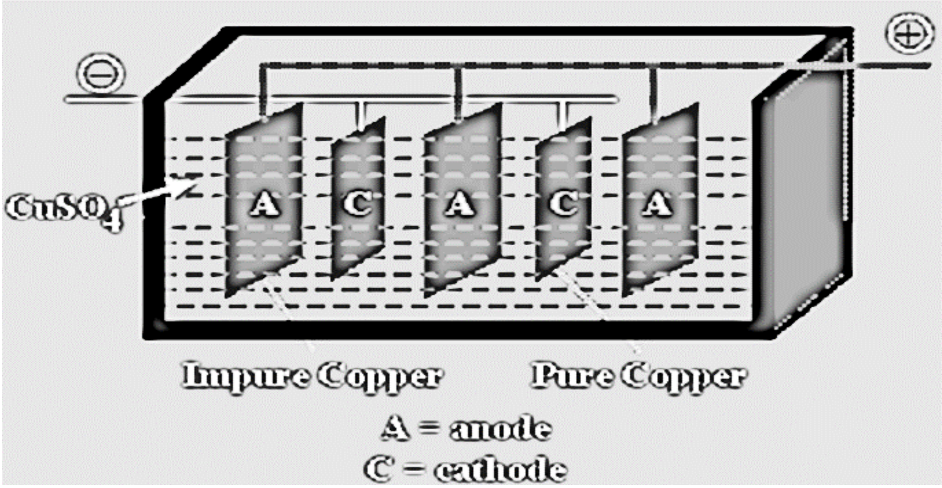
$$FeO\ (s) + SO_2\ (g) \rightarrow FeSiO_3\ (l)$$

A slag of iron silicate is formed.

Copper (l) sulphide is heated in limited supply of air in converter where it is reduced to copper blister and the sulphur dioxide further escapes in air.

$$Cu_2S\ (g) + O_2\ (g) \rightarrow 2Cu(s) + SO_2\ (g)$$

The copper blister which is impure is purified by electrolysis.



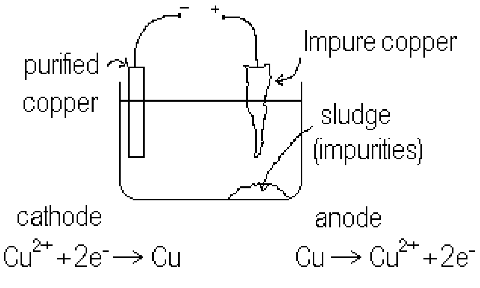
The impure bar is connected to the anode (***positive electrode***) and the thin pure copper metal form cathode (***negative electrode***). The electrolyte (***solution that conduct electricity***) of copper (ii) sulphate is used.

At the anode the copper blister dissolves in the solution and start to reduce in size.

$$Cu\ (s) \rightarrow Cu^{2+}\ (aq) + \tilde{e}$$

At the cathode copper ions are accepted and increases the bar.

$$Cu^{2+}\ (aq) + 2\tilde{e} \rightarrow Cu(s)$$



Properties of copper

- Copper is a heavy reddish metallic luster with a density of 8.9g/cm^3
- malleable and ductile (can form into wires) with very high thermal and electrical conductivity
- Melting point 1083.4 +/- 2^{o}C.
- Corrosion resistance

(IV) EXTRACTION OF ALUMINIUM

The chief ore of aluminium is bauxite (Al_2O_3, $2H_2O$). Aluminium is extracted from its pure oxide by electrolysis.

To obtain the pure oxide, bauxite is dissolved in sodium hydroxide solution under pressure to form sodium aluminate.

$Al_2O_3(s) + 3H_2O\ (l) + 2NaOH\ (aq) \rightarrow 2NaAl\ (OH)_4\ (aq)$ (***sodium aluminate***).

The solution of sodium aluminate is filtered and seeded with pure aluminium oxide causing the aluminium hydroxide to crystalline to pure state.

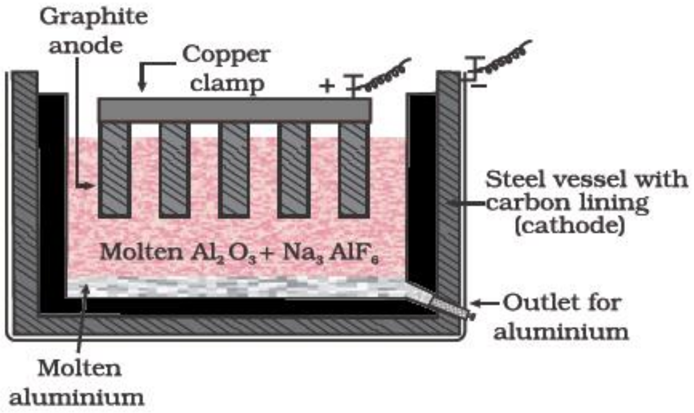
$NaAl\ (OH)_4\ (aq) \rightarrow NaOH\ (aq) + Al\ (OH)_3\ (s)$ (***aluminium hydroxide***).

Subsequent heating to 1200oc in the rotary kiln produces pure aluminum oxide

$2Al\ (OH)_3\ (S) \rightarrow 3H_2O\ (g) + Al_2O_3(s)$ (***pure aluminium oxide***)

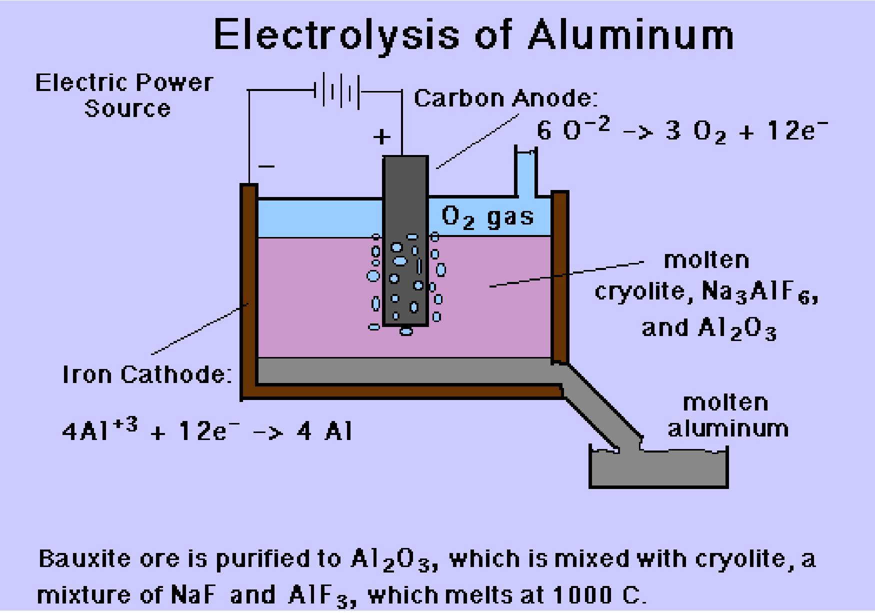
The pure oxide is dissolved in molten cryolite Na_2AlF and the mixture acts the electrolyte in the

Diagrams showing cell for aluminium extraction.



$Al^{3+} + 3\check{e} \rightarrow Al(s)$

$2O^{2-} \rightarrow O_2(g) + 2\check{e}$



Bauxite ore is purified to Al_2O_3, which is mixed with cryolite, a mixture of NaF and AlF_3, which melts at 1000 C.

$$6O^{2}- \rightarrow 3O_2 + 12\tilde{e}$$

$$4Al^{+3} + 12\breve{e} \rightarrow 4Al$$

At the cathode aluminium ions gain electrons to form aluminium metal

At the anode the oxide ions lose electrons to forms the oxygen gas.

12.10.2.6 Describe the uses of copper, iron, zinc and aluminium

Uses of copper, iron, zinc and aluminium: Such as electrical wires, construction, aircraft parts. Making alloys, pipes etc.

12.10.2.7 Explain the harmful effects of some metals.

Harmful effects of metals:

Lead poisoning (brain damaging), sodium ions in raising high blood pressure, alzehermia (hypertension) by aluminium.

12.10.3.1Describe what an alloy is

An alloy is a mixture of metals or a mixture of a metal and another element. An alloy is a mixture of two or more elements, where at least one element is a metal. Examples of mixtures of metals bronze (copper and tin) and brass (copper and zinc). Mixture of a metal with another substance steel (iron and carbon).

12.10.3.2 Describe diagrammatic representations of alloys.

Alloys contain atoms of different sizes which distort the regular arrangements of atoms. This makes it more difficult for the layers to slide over each other, so alloys are harder than the pure metal.

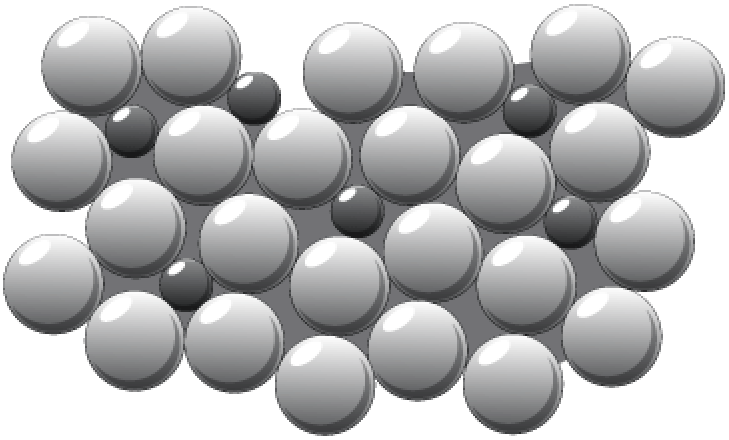


Figure ;showing Structure of an alloy

It is more difficult for layers of atoms to slide over each other in alloys.

12.10.3.3 Explain the advantages of using alloys over pure metals

Advantages of using alloys are that alloys exhibit better properties compared to a pure metal such as:

- Conducting of both electricity and heat.
- They are generally stronger than pure metals.
- Some are lighter than pure metals.
- They have high tensile strength.
- They are harder than pure metals.

12.10.3.4 Identify common uses of alloys

- Silver is used for making coins
- Stainless is used for making car parts, kitchen sinks and cutlery
- Duralium is used for making aircraft parts
- Bronze is used for making statues, ornaments, church bells
- Solders used for joining wires and pipes
- Brass used for making musical instrument.

Composition of common alloys and their properties

Alloys	**compositions**	**properties**	**Uses**
Bronze	➢ 90% copper ➢ 10% tin	Hard and strong. Doesn't corrode easily. Has shiny surface.	To build statues and monuments, Artistic materials
Brass	➢ 70% copper ➢ 30% zinc	Harder than copper	Making; musical instruments and kitchen ware
Steel	➢ 99% Iron ➢ 1% carbon	Hard and strong	Making of bridges, car bodies and railway trucks
Stainless steel	➢ 74% iron ➢ 8% carbon ➢ 18%chromium	Shiny Strong Doesn't rust	Making cutlery and surgical instruments
Duralumin	➢ 93% aluminium ➢ 3% copper ➢ 3% magnesium ➢ 1% manganese	Light Strong	Making bodies of aeroplanes and bullet trains
Pewter	➢ 96% tin ➢ 3% copper ➢ 1% antimony	Luster Strong Strong	In making souvenir

12.10.4 Corrosion

Generally metals corrode in air except (gold and silver) noble metals or unreactive.

12.10.4.1 Describe what corrosion is

- Corrosion is the chemical wearing of metal resulting from attack by atmospheric oxygen in pressure of moisture.
- Corrosion of iron is called rusting. Air and moisture are the conditions for rusting.

12.10.4.2 Relate corrosion to the reactivity of metals.

- More reactive metals easily corrode while less reactive metals do not easily corrode.
- Potassium and sodium tarnish in air. Generally corrosion of metals depends on the position of the metal in the activity series.

12.10.4.3 Describe the different methods of preventing corrosion

The methods of preventing corrosion:

- Sacrificial protection(***protecting iron/steel with more reactive metal***)
- Painting (***coating with paint***)

- Greasing and oiling (***applying oil or grease on the metal***)
- Galvanizing (***coating iron/steel with zinc***)
- Tinning (***coating iron/steel with tin which is less reactive***).

12.11 NON-METALS

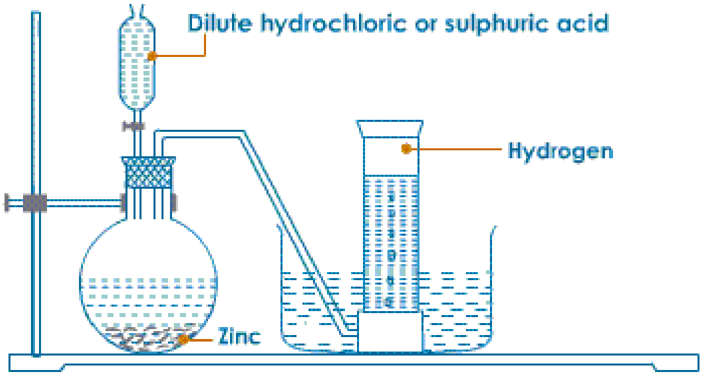
12.11.1. GENERAL PROPERTIES OF NON-METALS

12.11.1 .1 **Describe the physical and chemical properties of non-metals**

- They are not strong, or malleable, or ductile, or sonorous
- They have lower melting and boiling points that metals
- They are poor conductors of electricity; graphite (carbon) is the only exception. They are also poor conductors of heat
- They have low densities
- Like metals, most of them react with oxygen to form oxides
- When they form ions, the ions are negative

12.11.2.1 **Demonstrate the laboratory preparation, collection and test for hydrogen**

- By action of (a) reactive metals on (i) water/steam and (ii) dilute acids (b) aluminiun or zinc on aqueous potassium or sodium hydroxide. Hydrogen is collected by upward delivery method. The test for hydrogen is that when introduced to a lighted splint, the splint is put out with a 'pop' sound.



12.11.2.2 **Describe the physical and chemical properties of hydrogen**

Physical

- It is colorless, odorless and tasteless gas
- No effect both blue and red litmus paper
- It is less than air
- It has a boiling point of 21K and a melting point of 14K

- It is the least dense gas

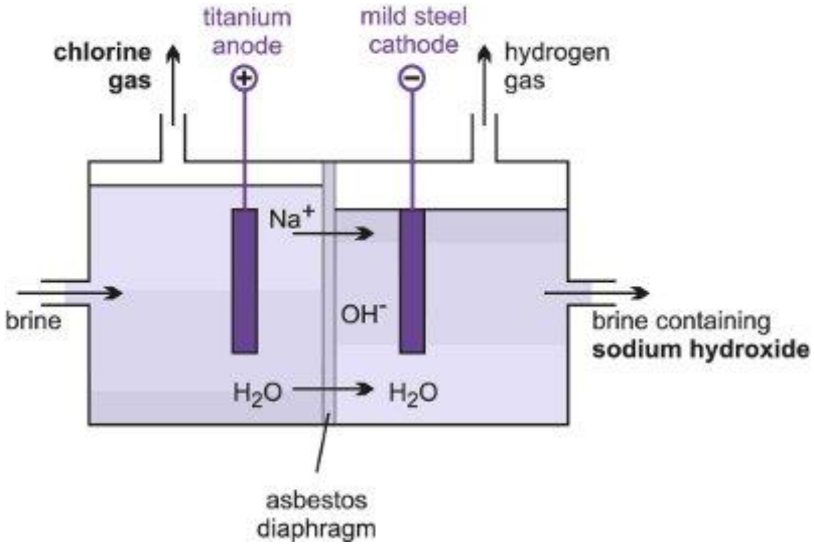
Chemical

- It reacts spontaneously with chlorine explosively in sunlight to form hydrogen chloride gas
- It reacts with oxygen explosively to form water vapour.

12.11.2.3 **Describe industrial manufacture of hydrogen**

- It is manufactured by cracking, electrolysis of brine (salt water) and from natural gas

Electrolysis of brine (salt water)



12.11.2.4 **Describe the uses of hydrogen**

- In Haber process for manufacture of ammonia gas
- In metalogical balloon because it is lighter
- Used in the manufacture of margarine using nickel as a catalyst
- Used as rockets fuel because hydrogen liberates a lot of energy when burned

12.11.3.1. Demonstrate the laboratory preparation, collection and test for oxygen

- By (i) catalytic decomposition of hydrogen peroxide and (ii) catalytic decomposition of potassium chlorate.It is collected above water .The test for oxygen is that when it is introduced to a glowing splint the splint re-lights.

12.11.3.2 Describe the physical and chemical properties of oxygen

Physical properties of oxygen

- It is a colorless gas
- It is tasteless and odorless
- Insoluble in water
- As dense as air

Chemical properties of oxygen

- It relights a glowing splint
- It reacts with metals to form metal oxides
- Reacts with non-metals

12.11.3.3 Describe the industrial manufacture of oxygen

- By fractional distillation of liquid air. The air is distilled into fractions of varying temperature within the tower. The tower is coolest at the top. The air separates into fractions as nitrogen separates first at -190^{0}C followed by argon at -186^{0}C then oxygen at -163 ^{0}C. By so doing oxygen will have been prepared.

12.11.3.4 Describe the uses of oxygen in industry and in natural processes

- It is used for burning
- It is used for welding
- It is used in the manufacture of steel
- In space by astronauts on space missions
- It is used for respiration
- In fuel cell to restore life to polluted lakes and rivers

12.11.3.5 Explain the importance of ozone layer and dangers of its depletion

- **IMPORTANCE-**It traps dangerous ultraviolet (UV) radiation from the sun. **DANGER-** if depleted by CFC's it causes skin cancer, respiratory diseases

12.11.3.6 Demonstrate the chemical test for water.

- Using white anhydrous copper (ii) sulphate which turns blue

12.11.3.7 Describe the importance of water

- **IN HOMES**-For laundry, drinking, cooking .**IN INDUSTRIES**- For cleaning , as a solvent, as a coolant, as a chemical reactant

12.11.5 NITROGEN

12.11.5.1 Describe the industrial manufacture of nitrogen

- By fractional distillation of liquid air. The air is distilled into fractions of varying temperatures within the tower. The tower is coolest at the top. The air separates into fractions as nitrogen separates, first at -190^{0}C followed by argon at -186^{0}C then oxygen at -163^{0}C

12.11.5.2 Explain **the characteristics and importance of nitrogen as a gas**

- It is a non-reactive insoluble gas hence used as refrigerant, food packaging and in the manufacture of ammonia

12.11.5.3 Demonstrate preparations collection and test for ammonia in the laboratory.

- Action of a base on ammonium salt and collected by upward delivery method, turns damp red litmus paper blue.

12.11.5.4 Describe the manufacture of ammonia

- Hydrogen and nitrogen gas are purified to remove impurities such as carbon (iv) oxide, oxygen sulphure (iv) oxide etc that would poison catalyst, hydrogen and nitrogen are mixed in a ratio of 1:3 and compressed to about 350 atm pressure and heated to about 540^{0}C. The hot compressed gases are passed over finely divided iron catalyst mixed with aluminium (iii) oxide promoter. The equation for the reversible reaction is:

  $N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$

  Ammonia is condensed out as liquids.

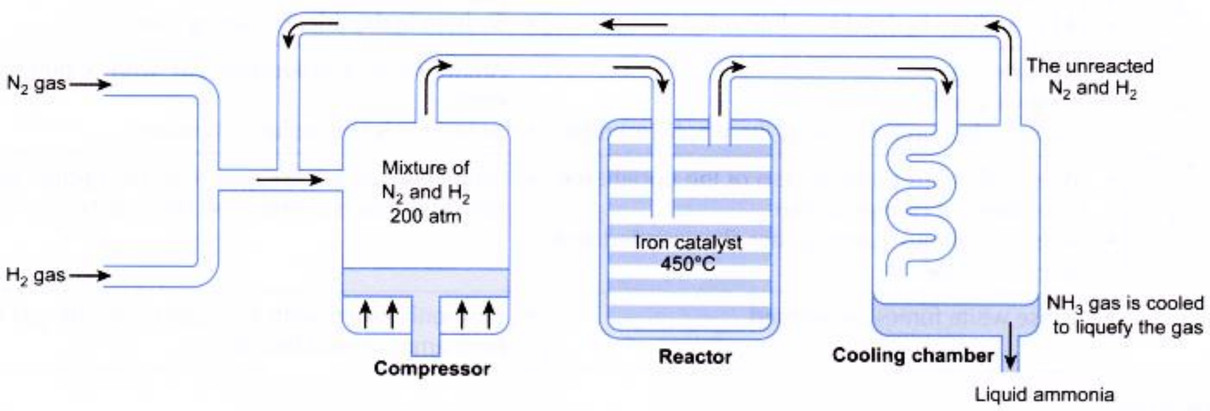


Figure *The manufacture of ammonia through the Haber process*

12.11.5.5 Describe the thermal dissociation of ammonium salts

When ammonium salts decompose they liberate ammonia gas

12.11.5.7 Describe the uses of ammonia

- In the manufacture of explosives
- Liquid ammonia is used as a refrigerant
- In the manufacture of fertilisers

12.11.5.8 Describe the manufacture of nitric acid by the Ostwald process

Step (i): catalytic oxidation of ammonia at 900^{0}C

$4NH_{3\,(g)} + 5O_{2(g)} \xrightarrow{\text{Pt catalyst}} 4NO + 6\,H_2O_{\,(l)}$

This reaction is exothermic and sustains itself once started.

Step (ii): Nitrogen (ii) oxide is cooled and reacted with excess air to form nitrogen (iv) oxide

$2NO_{\,(g)} + O_{2\,(g)} \longrightarrow 2NO_{2\,(g)}$

Step (iii): absorption of nitrogen (iv) oxide in water to form nitric acid

$3NO_{2\,(g)} + H_2O_{\,(g)} \longrightarrow 2HNO_{3\,(aq)} + NO_{(g)}$. Percentage of nitric acid can be concentrated from 60% to 68% by distillation

12.11.5.9 Explain the importance of nitrogenous fertilizers

- Nitrogen for growth,
- phosphorous for root development and
- potassium for seed formation. The three elements are commonly referred to as NPK.
- NOTE**: NPK** stands for Nitrogen, Phosphorous and Potassium.

12.11.5.10 Describe the effects of nitrogenous fertilizer on the environment

- The form acid rain
- Some nitrogen (IV) oxide in air absorbs solar energy and dissociates into nitrogen atoms combines with oxygen in air forming ozone which combines with hydrocarbons producing poisonous compounds.
- If fertilizers are not used correctly, they reach rivers and storms encouraging growth of algae and other marine plants that deprive water of oxygen leading to death of marine animals.

12.11.6. CHLORINE

12.11.6.1 Demonstrate the laboratory preparation, collection and test for chlorine gas.

- By action of hot concentrated Hydrochloric acid on manganese (IV) oxide,
- collected by downward delivery method,
- turns damp blue litmus paper red and then bleaches it

12.11.6.2 Describe the physical and chemical properties of chlorine gas

Physical and chemical properties of chlorine gas such:

- Chlorine has a melting point of -100.98°C, boiling point of -34.6°C, density of 3.214 g/l, specific gravity of 1.56 (-33.6°C), with a valence of 1. Chlorine is a member of the halogen group of elements and directly combines with almost all of the other elements. Chlorine gas is a greenish yellow.

12.11.6.3Describe the uses of chlorine.

- Uses of chlorine for sterilizing water,
- manufacture of PVC, HCl and
- in bleaching agents.
- **Uses:** Chlorine is used in many everyday products. It is used for disinfecting drinking water. Chlorine is used in the production of textiles, paper products, dyes, petroleum products, medicines, insecticides, disinfectants, foods, solvents, plastics, paints, and many other products. The element is used to manufacture chlorates, carbon tetrachloride, chloroform, and in the extraction of bromine. ***Chlorine has been used as a chemical warfare agent***.

12.11.6.4 Describe the industrial manufacture of chlorine

- Manufacture of chlorine: By the electrolysis of brine ($NaCl_{(aq)}$)

12.11.6.5 Demonstrate the method for preparation, collection and test for hydrogen chloride gas

- By action of concentrated sulphuric acid on solid metallic chlorides, collected by downward delivery method, react with ammonia to form white smoke.

12.11.6. 6 Describe the physical and chemical properties of hydrogen chloride gas

- In terms colour, odour, density, solubility and poisonous.

Reactions with ammonia and water

12.11.6. 7 Demonstrate the method for preparation of hydrochloric acid.

- By dissolving hydrogen chloride gas in water

12.11.6.8 Describe the reactions of dilute hydrochloric acid.

Such as reaction with alkalis, metals, carbonates, ammonia and silver nitrate

12.11.7 SULPHUR

12.11.7.1 Describe the formation of sulphur dioxide.

- By combustion of sulphur, fossil fuels

12.11.7.2 Demonstrate the laboratory preparation, collection and test for sulphur dioxide

- By action of warm dilute acids on sulphites, collected by downward delivery, turns acidified potassium dichromate (VI) green/decolourises purple potassium manganate (VII).

12.11.7.3 Describe the physical and chemical properties of sulphur dioxide

- The physical and chemical properties of sulphur dioxide.
  - SULFUR DIOXIDE is a colorless gas with a choking or suffocating odor.
  - Boiling point -10°C. Heavier than air.
  - Very toxic by inhalation and may irritate the eyes and mucous membranes. Under prolonged exposure to fire or heat the containers may rupture violently and rocket. Used to manufacture chemicals, in paper pulping, in metal and food processing.
  - Reaction with water, action on indicators and as a reducing agent E.g. turns acidified potassium dichromate (VI) green/decolourises purple potassium manganate (VII).

12.11.7.4 Describe the uses of sulphur dioxide

- **Uses of sulphur dioxide**
  - As food preservative,
  - bleaching wood pulp for paper making,
  - manufacture of sulphuric acid

12.11.7.5 Describe the industrial manufacture of sulphuric acid.

- By Contact Process (catalyst, temperature)

A brief summary of the Contact Process

The Contact Process:

- makes sulphur dioxide;
- convers the sulphur dioxide into sulphur trioxide (the reversible reaction at the heart of the process);
- converts the sulphur trioxide into concentrated sulphuric acid.

Making the sulphur dioxide

This can either be made by burning sulphur in an excess of air:

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

. . . or by heating sulphide ores like pyrite in an excess of air:

$$4FeS_{2(s)} + 11O_{2(g)} \longrightarrow 2Fe_2O_{3(s)} + 8SO_{2(g)}$$

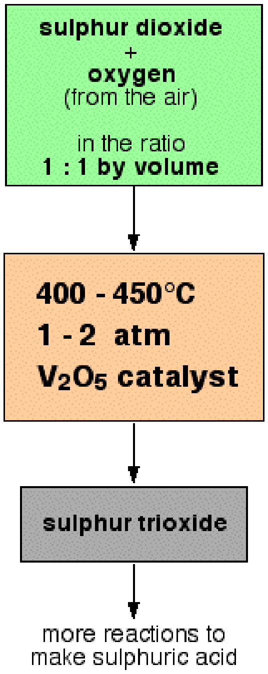
In either case, an excess of air is used so that the sulphur dioxide produced is already mixed with oxygen for the next stage.

Converting the sulphur dioxide into sulphur trioxide

This is a reversible reaction, and the formation of the sulphur trioxide is exothermic.

$$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)} \qquad \Delta H = -196 \text{ kJ mol}^{-1}$$

A flow scheme for this part of the process looks like this:



The reasons for all these conditions will be explored in detail further down the page.

Converting the sulphur trioxide into sulphuric acid

This can't be done by simply adding water to the sulphur trioxide - the reaction is so uncontrollable that it creates a fog of sulphuric acid. Instead, the sulphur trioxide is first dissolved in concentrated sulphuric acid:

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

The product is known as ***fuming sulphuric acid*** or ***oleum***.

This can then be reacted safely with water to produce concentrated sulphuric acid - twice as much as you originally used to make the fuming sulphuric acid.

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

12.11.7.6 Describe the uses of sulphuric acid.

- Used in explosives,
- as drying agent,
- making of soaps,
- fertilisers

12.11.8 CARBON AND CARBONATES

12.11.8.1 Describe allotropes

- When an element exists in several physical forms of the same state, it is said to exhibit allotropy. Each form of this element is an allotrope. Lots of elements exhibit allotropy. Carbon has two very popular allotropes, diamond and graphite. Diamond and graphite are both made of carbon only. However, they look very different and have different physical properties. They are both giant molecular structures.

12.11.8.2 Describe the physical properties of the allotropes of carbon

(i) GRAPHITE

- it is a dark grey shiny solid
- it conducts electricity
- it is a soft material with a slippery feel
- it has a density of 2.25g/cm^3

(ii) DIAMOND

- it is a colourless transparent crystal which sparkles in light
- it does not conduct electricity
- it is a very hard substance
- it has a density of 3.51g/cm^3

12.11.8.3 Describe the formation and properties of carbon monoxide

- **Formation**

- By incomplete combustion of carbon and carbon compounds,
- reduction of carbon dioxide by carbon.

- **Properties of carbon monoxide**

- It is colourless, odourless, density,
- It is insoluble in water,
- It is very poisonous,
- It acts as a reducing agent.

12.11.8.4 Demonstrate the laboratory preparation, collection and the test for carbon dioxide

- By reaction of dilute HCl acids with carbonate or bi-carbonate,
- collection by downward delivery method/above water.
- Test for carbondioxide is that when it is bubbled through lime water the lime water turns milky due to formation of a white precipitate.

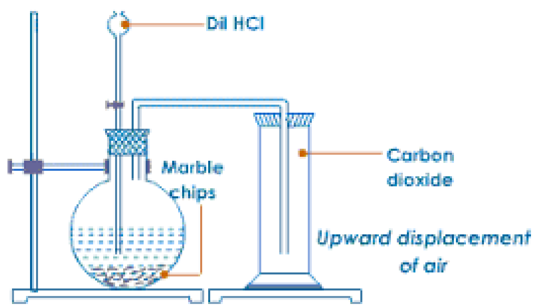


Fig. 11.6 Laboratory Preparation of carbon dioxide

12.11.8.4 Describe the physical and chemical properties of carbon dioxide.

- **Physical properties of carbondioxide**
  - It is a colourless, odourless, tasteless gas
  - It is much heavier than air
  - It is slightly soluble in water, its solubility increases with pressure
  - If carbon dioxide is cooled and pressurized, it turns straight into a solid, which is called dry ice.

Chemical properties of carbon dioxide

- Burning magnesium decomposes carbondioxide to produce magnesium oxide and carbon
- It dissolves a little in water and the dissolved gas reacts slightly with the water producing a dilute solution of carbonic acid
- Being an acid gas it reacts and gets absorbed by a solution of an alkali

$$2NaOH_{(aq)} + CO_{2\,(aq)} \longrightarrow Na_2CO_{3(aq)} + H_2O(aq)$$

12.11.7.5 Describe the uses of carbon dioxide

- Used in fire extinguishers
- Dry ice is used as a refrigerant
- For making carbonated drinks
- for photosynthesis

12.11.8.6 Describe the manufacture of lime from limestone

By heating limestone ($CaCO_3$)

$$CaCO_{3(s)} \longrightarrow CaO_{(s)} + CO_{2\,(g)}$$

(Lime)

12.11.8.7 Describe the uses of lime and slaked lime

- In neutralizing acidic soil,
- lime as a drying agent for ammonia

12.11.8.8 Describe the uses of limestone

Used in manufacturing of lime,

- Manufacture of cement
- Manufacture of glass
- Manufacture of iron

12.11.8.9Describe the greenhouse effect

- In the day, heat from the sun (in the form of infrared) passes through the atmosphere heating up the earth. At night, the earth radiates heat, to the outer space. Some atmospheric gases trap heat from the sun, thus preventing the loss of heat. This is called greenhouse effect (which is essential or else the earth will be too cold for life.
- Global warming due to increase of carbon dioxide in the atmosphere.

12.11.9 SILICON

12.11.9.1 Describe the properties of silicon.

The properties of silicon:

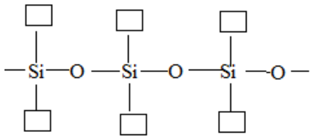
- It's a metalloid

12.11.9.2 Describe the use of silicon.

- Use of silicon:
  - Used in semi-conductors Such as transistors,
  - diodes and
  - capacitors.

12.11.9.3 Describe what silicones are.

- Silicones: macromolecules that exist as oils, waxes or plastics and their structures represented as:



12.11.9.4 Compare the fire resistance of silicone plastics to carbon based macromolecules

- The nature of silicones: With reference to nature of combustion products
  - ❖ Silicones produce silicon dioxide (sand) while organic based macromolecules produce carbon dioxide.

12.11.9.5 Describe the uses of silicon dioxide (sand)

- making glass,
- as fire extinguisher,
- in iron extraction.

12.12 ORGANIC CHEMISTRY

12.12.1.1 Describe an organic compound.

- A compound of carbon other than oxides, carbonates and carbides.

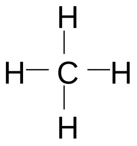
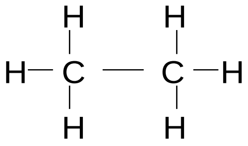
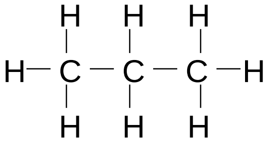
12.12.1.2 Describe a hydrocarbon.

- A binary compound of carbon and hydrogen only.
- The two types are**:** **saturated** and **unsaturated** hydrocarbons.
  - ❖ **Saturated hydrocarbons** contain **single carbon to carbon** covalent bonds .e.g. alkanes.
  - ❖ **Unsaturated hydrocarbons** contain at least **one multiple carbon to carbon** covalent bond. Those with **one double carbon to carbon** covalent bond are **alkenes** while those with **one triple carbon to carbon** covalent bond are **alkynes**.

Alkanes

- Are a group of saturated hydrocarbons with general formula C_nH_{2n+2}.
- All names end with **-ane**.

12.12.1.3 Name the structures of the aliphatic alkanes up to five carbon atoms.

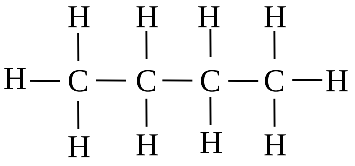
Number of carbon atoms(n)	Name	Molecular formula	Structural formula
n=1	**Methane**	**CH_4**	 $H-\overset{H}{\underset{H}{C}}-H$ CH_4
n =2	**Ethane**	**C_2H_6**	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$ CH_3CH_3
n =3	**Propane**	**C_3H_8**	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$ $CH_3CH_2CH_3$
n =4	**Butane**	**C_4H_{10}**	**complete**
n =5	**Pentane**	**C_5H_{12}**	**complete**

12.12.1.4 Demonstrate the structures of Isomers and their names.

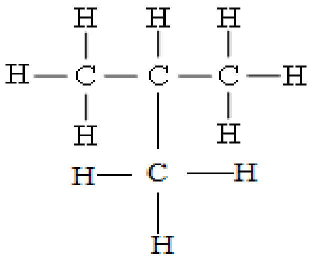
Isomers

- Organic compounds with same molecular formula but with different structures.
- In alkanes isomerism begins at butane (C_4H_{10}).Methane, ethane and propane do not have isomers.

Isomers of butane



$$\begin{array}{ccccccccccc} & & H & & H & & H & & H & & \\ & & \mid & & \mid & & \mid & & \mid & & \\ H & - & C & - & C & - & C & - & C & - & H \\ & & \mid & & \mid & & \mid & & \mid & & \\ & & H & & H & & H & & H & & \end{array}$$



$CH_3CH_2CH_2CH_3$

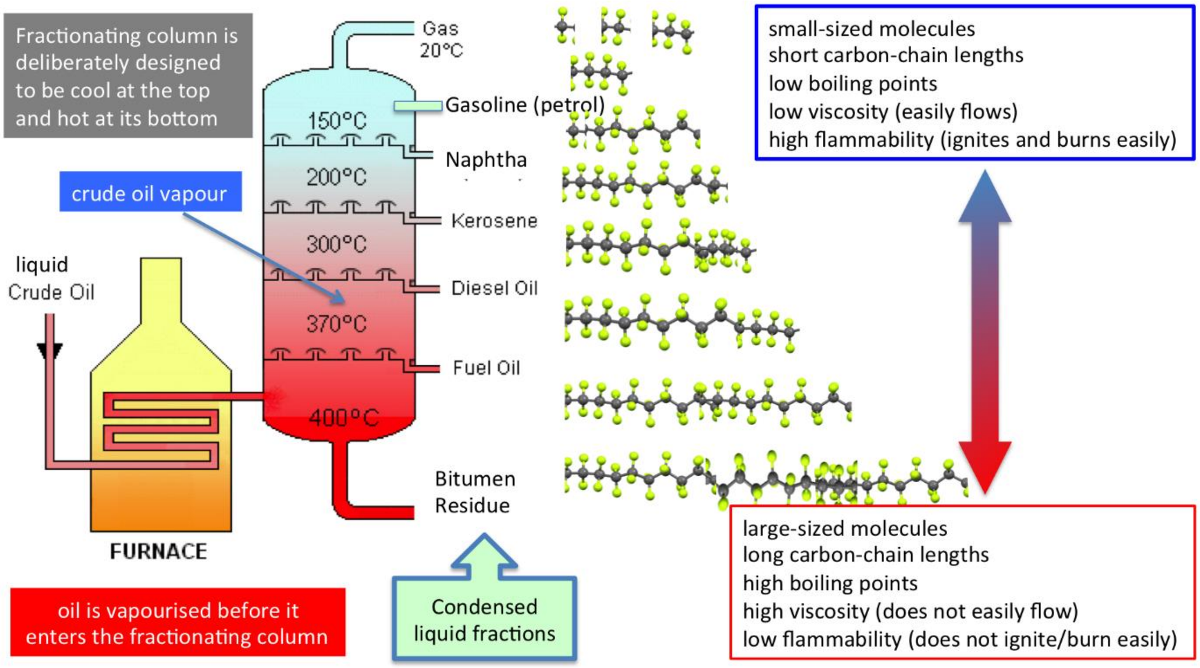
Butane

2-Methyl propane

12.12.1.5. Describe fractional distillation of petroleum (crude oil)

- Crude oil contains a mixture of hydrocarbon most of them being alkanes. Some of the compounds of crude oil can be separated by fractional distillation.
- Solids such as sand are separated by filtration. The crude oil layer is pumped into furnace at about 400^0C to vaporize it.
- The hot vapour is pumped into a fractionating column near the bottom. The tower contains perforated horizontal plates at different temperature ranges that are the same as boiling points.

Diagram of fractionating tower



12.12.1.6 Describe the uses of the fractions of crude oil.

- In order of increasing boiling points(also molecular mass) the fractions separated and their uses are in the table below:

Fraction	Length of carbon chain	Uses
Petroleum gases	C_1-C_4	For heating in gas stoves and Bunsen burners.
Petrol	C_4-C_{12}	As solvent; fuel in petrol engines.
Kerosene	C_{11}-C_{15}	Fuel in airplanes; lighting in lamps.
Diesel	C_{15}-C_{19}	Fuel in diesel engines; production of petrol.
Lubricating oil	C_{20}-C_{30}	Lubricants in machines; making polishes and waxes.
Fuel oil	C_{30}-C_{40}	Fuel in ships and in blast furnace.
Bitumen	C_{50} and above	Surfacing roads and production of other smaller fractions.

12.12.1.7 Describe the chemical properties of alkanes.

- Combustion
  - ❖ Burn in excess air (oxygen) with a pale blue flame forming carbon dioxide and water (complete combustion).

    $CH_{4\,(g)} + 20_{2\,(g)} \rightarrow CO_{2\,(g)} + 2H_2O_{(l)}$
  - ❖ In limited (insufficient) air burn to produce carbon monoxide, water (incomplete combustion).

    $CH_{4\,(g)} + O_{2\,(g)} \rightarrow CO_{(g)} + H_2O_{(g)}$

- In the presence of sunlight or UV light, alkanes react with halogen forming series of products depending on the amount of halogen.

$CH_4 + Cl_{2\,(g)} \rightarrow CH_3Cl_{\,(g)} + HCl_{\,(g)}$

- Cracking is the decomposition of long chain hydrocarbons into short chain alkanes and alkenes.
  - ❖ Two types of cracking are thermal cracking and catalytic cracking.
  - ❖ Whenever an alkane is cracked an alkene is always produced.

$$C_{10}H_{22} \rightarrow C_5H_{22} + C_5H_{10}$$

12.12.1.8 Account the apparent unreactivity of alkanes compared to other organic compounds.

- Alkanes are unreactive compared to other organic compounds such as alkenes because they are saturated hence do not have an active site for chemical attack.

12.12.1.9 Illustrate the unsaturation.

- **Unsaturation** is due to **the double carbon to carbon covalent** bond (as in alkenes) or **the triple carbon to carbon covalent** bond (as in alkynes).
- Unsaturated compounds, therefore, quickly decolourises the red colour of bromine in tetrachloromethane.

12.12.1.10. Describe and name the structures of the alkenes up to 5 carbon atoms.

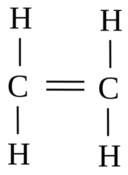
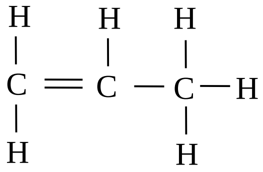
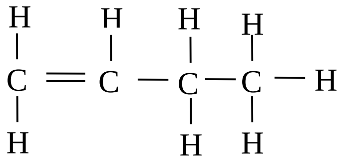
- Alkenes are unsaturated hydrocarbons with general molecular formula C_nH_{2n}.
- Alkenes are named by replacing the suffix –**ane** in an alkane with the same number of carbon atoms with –**ene**.e.g.

Alk**ane** → alk**ene**

Eth**ane (n=2)** → eth**ene (n=2)**

$C_2H_6 \rightarrow C_2H_4$

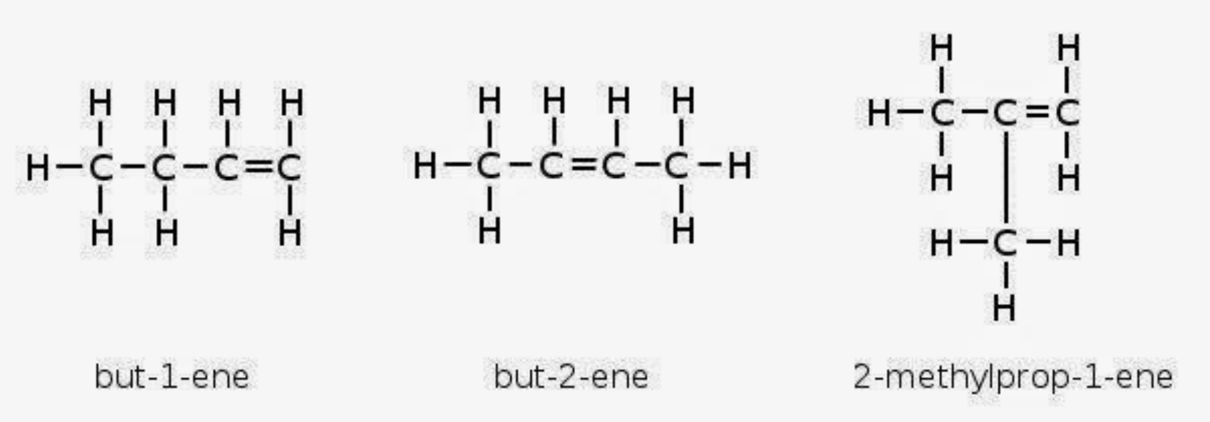
Names and structures of the first five alkenes

Number of carbon atoms	Name	Molecular formula	Structural formula
n=2	Ethene	C_2H_4	 CH_2CH_2
n=3	Propene	C_3H_6	 CH_2CHCH_3
n=4	Butene	C_4H_8	 $CH_2CHCH_2CH_3$
n=5	Pentene	C_5H_{10}	COMPLETE
n=6	Hexene	C_6H_{12}	COMPLETE

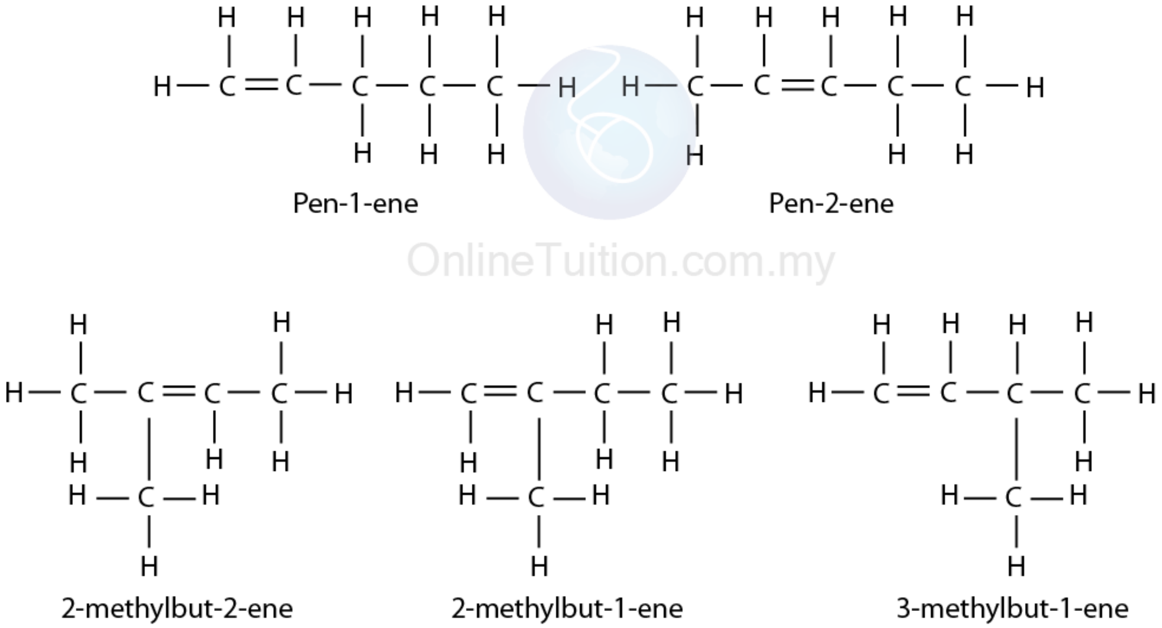
12.12.1.11 Demonstrate the Isomers of alkenes

- In alkenes isomerism starts at Butene. Ethene and Propene do not exhibit isomerism.
- Isomerism considered is positional isomerism i.e. only changing the position of the double carbon to carbon bond.

Isomers of Butene



Isomers of Pentene



12.12.1.12. Describe the chemical properties of alkenes.

Reactions of alkenes include combustion and addition reactions.

- Combustion(Reaction with air or oxygen)
  - ❖ Burn in excess air with luminous flame forming carbon dioxide and water.

$C_2H_4 \, _{(g)} + 3O_2 \, _{(g)} \rightarrow 2CO2 \, _{(g)} + 2H_2O \, _{(l)}$

- ❖ In limited air produce carbon monoxide and water

  $C_2H_4 + 2O_2 \rightarrow 2\ CO + 2\ H_2O$

- **Addition reaction:** is a reaction in which two or more molecules combine to yield a single molecule (or product).This is due to the double carbon to carbon bond they have.
- Alkenes undergo additional reactions with:
  - ❖ Hydrogen (Hydrogenation) to produce a corresponding alkane.

$$C_2H_4\ _{(g)} + H_2\ _{(g)} \rightarrow C_2H_6\ _{(g)}$$

Ethene ethane

  - ❖ Chlorine and bromine (Halogenation) to produce di-chloroalkane and di-bromoalkane respectively. Reaction with Bromine is used as a test for unsaturated compounds. Unsaturated compounds quickly decolourises the red brown colour of bromine.

$$C_2H_4\ _{(g)} + Cl_2\ _{(g)} \rightarrow CH_2ClCH_2Cl\ _{(l)} \text{ (1, 2-dichloroethane)}$$

$$C_2H_4\ _{(g)} + Br_2\ _{(g)} \rightarrow CH\ BrCH_2Br\ _{(l)} \text{ (1, 2-dibromoethane)}$$

Red-brown colourless

  - ❖ Hydrogen chloride (hydrohalogenation) to produce an alkyl halide (or halo alkane).

$$C_2H_4\ _{(g)} + HCl\ _{(g)} \rightarrow CH_3CH_2Cl \quad \text{(ethyl chloride or chloroethane).}$$

  - ❖ Steam (hydration): in the presence of phosphoric acid produce an alcohol (alkanol).

$$C_2H_4\ _{(g)} + H_2O\ _{(g)} \rightarrow C_2H_5OH\ _{(l)} \text{ (ethanol)}$$

  - ❖ Polymerization alkene reacts with itself to produce a polyalkene.

$$\ldots + C_2H_4\ _{(g)} + C_2H_4\ _{(g)} + \ldots \rightarrow \ldots\ CH_2CH_2CH_2CH_2 \ldots\ldots \text{ (Polyethene)}$$

12.12.1.13. Illustrate the differences and similarities between saturated and unsaturated hydrocarbons.

- **Similarities**
  - ❖ Both contain carbon and hydrogen only.
  - ❖ Both undergo combustion reaction.

- **Differences**

Saturated hydrocarbons	**Unsaturated hydrocarbons**
Contain single C-C bonds in their structures.	Contain a double C=C or a triple C=C bond .
Do not rapidly decolourise red brown bromine.	Decolourise red brown bromine rapidly.
Do not decolourise acidified purple potassium manganate (VII).	Decolourise acidified purple potassium manganate (VII).

12.12.1.14 Describe the chemical tests for unsaturated hydrocarbons.

- Unsaturated hydrocarbons (alkenes and alkynes) are tested using bromine solution (or bromine water).
    - ❖ Unsaturated hydrocarbons decolourises bromine solution rapidly.

12.12.1.15 Describe the uses of alkenes.

- Ethene is used in manufacturing of polythene which is used for making plastic bags, water pipes etc.
- Ethene in low concentration is used to ripen fruits such as bananas.
- Propene is used in manufacture of polypropene which is used to make ropes.

12.12.2.1 Describe the chemical composition of an alcohol.

- **Alcohol** is an organic compound with hydroxyl group (OH) as a functional group.
- They have a general formula $C_nH_{2n+1}OH$.
- Names are derived from corresponding alkane by replacing ending **e** by **ol**. e.g.

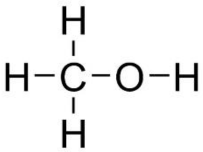
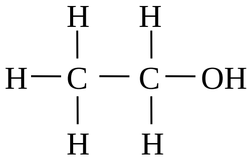
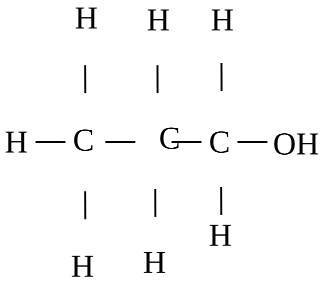
Alkan**e** → alkan**ol**

Methan**e (n=1)** → methan**ol (n=1)**

CH_4 → CH_3OH

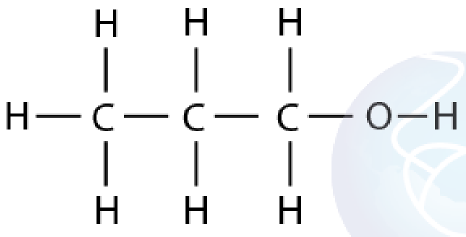
12.12.2.2 Names and structures of primary alcohols up to five carbon atoms.

Names and structures of the first five alcohols.

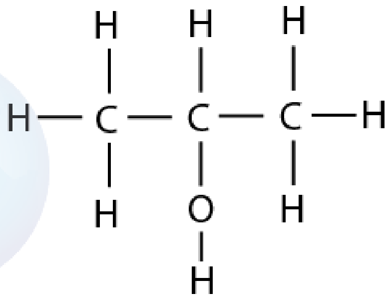
NUMBER OF CARBON ATOMS	NAME	MOLECULAR FORMULA	STRUCTURAL FORMULA
N=1	Methanol	CH_3OH	 CH_3OH
N=2	Ethanol	C_2H_5OH	 CH_3CH_2OH
N=3	Propanol	C_3H_7OH	 $CH_3CH_2CH_2OH$
N=4	Butanol		complete

12.12.2.3 Demonstrate Isomerism in alcohols.

- **Isomers of propanol**

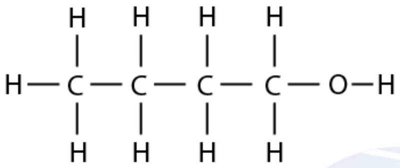


Propan-1-ol

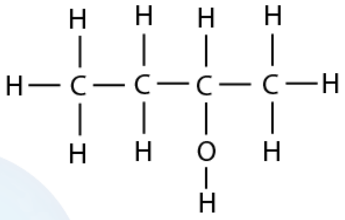


Propan-2-ol

- **Isomers of Butanol**



Butan-1-ol



Butan-2-ol

Exercise

- **Draw and name the Isomers of pentanol**

12.12.2.4. Describe the formation of alcohols.

- Hydration of an alkene: Alkene is absorbed in 98% sulphuric acid at 80°C and 25 atm pressure to form an alkyl hydrogen sulphate which is hydrolyzed by boiling water to form an alcohol.

$C_2H_4 + H_2SO_4 \rightarrow C_2H_5HSO_4$ (ethyl hydrogen sulphate).

$C_2H_5HSO_4 + H_2O \rightarrow C_2H_5OH + H_2SO_{4\,(aq)}$

- **Fermentation of carbohydrates**: Various carbohydrates such as maize, potatoes and barley are pressure cooked to release starch grain. The mixture is then treated for an hour at 60^0C with malt which supplies enzymes diastase that hydrolyze it to maltose.

$(C_5H_{10}O_5)_n + nH_2O \rightarrow C_{12}H_{22}O_{11}$

Yeast is added at room temperature. Its enzyme maltase hydrolyses maltose to glucose.

$C_{12}H_{22}O_{11} + H_2O \rightarrow 2C_6H_{12}O_6$

Another enzyme in yeast, zymase, converts glucose to ethanol.

$$C_6H_{12}O_{6\ (g)} \rightarrow 2\ CO_{2\ (g)} + 2C_2H_5OH_{\ (l)}$$

12.12.2.5. Describe the chemical properties of alcohols.

- Combustion: Burn in excess oxygen to form water and carbon dioxide.

$$C_2H_5OH\ (l) + 3O_2\ (g) \rightarrow 2\ CO_2\ (g) + 3H_2O\ (l)$$

- Reaction with sodium metal to produce hydrogen gas and an alkoxide.

$$C_nH_{2n}{+}1OH + 2Na \rightarrow 2C_nH_{2n}{+}1ONa + H_2\ (g)$$

Alkanol sodium alkoxide

- Dehydration: Alcohols are dehydrated by excess concentrated sulphuric acid at 170^0C to form alkene and water.

$$C_2H_5OH \rightarrow C_2H_4 + H_2O$$

- Reaction with alkanoic acid (esterification): When alkanol is boiled with alkanoic acid in presence of concentrated sulphuric, an ester is formed.

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

Ethanoic acid + ethanol → ethylethanoate (ester).

- Oxidation of alcohols : Alcohols are oxidized in excess of oxidizing agent such as acidified potassium manganate (VII) or potassium dichromate (VI) in two distinct stages first to alkanal and then to an alkanoic acid with no loss of carbon.

12.12.2.6 Describe the uses of alcohols.

- Alcohols, for example, ethanol is used:
  - ❖ As an organic solvent.
  - ❖ As a fuel.
  - ❖ As an antiseptic.
  - ❖ As a constituent in alcoholic beverages.
  - ❖ As a thermometric liquid.

12.12.3.1 Draw and name structures of carboxylic acids up to five carbon atoms.

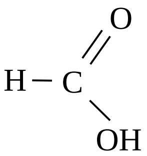
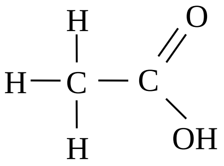
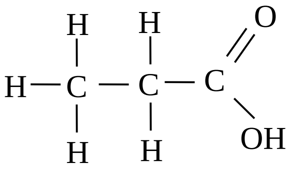
- **Carboxylic acids** have a general formula $\mathbf{C_nH_{2n+1}COOH}$, where the carboxyl group **-COOH** is the functional group. Are also called organic acids or alkanoic acids.
- Named by replacing **e** in an alkane with the same number of carbon atoms by **oic acid.**

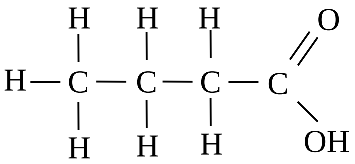
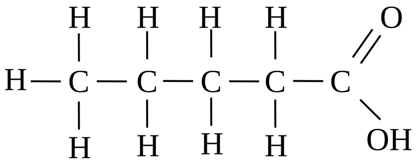
  Alkan**e** → alkan**oic acid**

  Methan**e** → methan**oic acid**

  CH_4 → HCOOH

First five carboxylic acids

NUMBER. OF CARBON ATOM	NAME	MOLECULAR FORMULA	STRUCTURAL FORMULA
N=0	Methanoic acid	HCOOH	 HCOOH
N=1	Ethanoic acid	CH_3COOH	 $CH_3\,COOH$
N=2	Propanoic acid	$C_2H_5\,COOH$	 $CH_3CH_2\,COOH$

N=3	Butanoic acid	C_3H_7COOH	 $CH_3CH_2CH_2COOH$
N=4	Pentanoic acid	C_4H_9COOH	 $CH_3CH_2CH_2CH_2COOH$

12.12.3.2 Describe the formation carboxylic acid.

- Carboxylic acids are formed by oxidation of alcohols in excess of oxidizing agent.
- By hydrolysis of esters.

12.12.3.3. Demonstrate the chemical properties of carboxylic acid.

- They react with a base forming a salt and water only.
- Liberate hydrogen with active metals.
- React with carbonate and hydrogen carbonate to form salt, water and carbon dioxide.
- Esterification: React with alcohols (in the presence of conc.H_2SO_4) to form an ester and water.

12.12.3.4. Describe the uses of carboxylic acids.

Uses carboxylic acids include:

- Preservative and tenderizer for food.
- For flavouring food stuff.
- As a solvent for many organic chemicals especially in textile, paint and colour making industry.

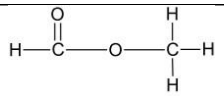
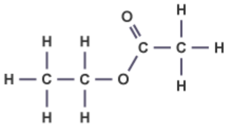
12.12.3.5 Describe and name the structure of ester up to five carbon atoms.

- Ester are organic compounds with general formula RCOOR where R are alkyl radicals derived from carboxylic acid and alcohol used respectively.

- Formed by reacting an alcohol and an organic acid in the presence of concentrated sulphuric acid. A reaction called esterification.
- Named by writing the radical name from the alcohol (alkyl) and the radical name from the alkanoic acid (alkanoate).i.e. **alkyl-alkanoate.** For example, an ester formed between ethanol and ethanoic acid is:

  Alk**anol** → alk**yl** radical　　　　alkano**ic acid** → alkano**ate**
  Eth**anol** → eth**yl**　　　　ethano**ic acid** → ethano**ate**
  hence name is: **ethylethanoate.**

Names and structures of esters up to five carbon atoms

Number of carbon atoms	Name	Molecular formula	Structural formula
n=1	**Methyl methanoate**	$HCOOCH_3$	
n=2	**Ethylethanoate**	$CH_3COOC_2H_5$	
n=3	**Propylpropanoate**	$C_2H_5COOC_3H_7$	**complete**
n=4	**Butyl butanoate**	$C_3H_7COOC_4H_9$	**complete**
n=5	**Pentylpentanoate**	$C_4H_9COOC_5H_{11}$	**complete**

12.12.4.2 Describe the chemical properties of ester.

- Combustion: undergo complete combustion in excess air forming carbon dioxide and water.
- Hydrolysis: are hydrolysed to the respective alcohol and organic acid which formed it.

12.12.4.3 Describe the uses of esters.

- Related to their physical property of having **a nice pleasant fruity smell (or scent or aroma, esters are:**
  - ❖ used in making perfumes.
  - ❖ used in food flavourants.
  - ❖ added Varnishes.

12.12.5.1 Describe homologous series.

- It is a collection of organic compounds belonging to the same family with the same general formula (consider alkanes, alkenes, alcohols, acids, esters).
- Members of a homologous series are called homologues.

12.12.5.2 Describe the general characteristics of the members of a homologous series.

- Members of each homologous series:
  - ❖ use the same general molecular formula.
  - ❖ show similar chemical properties.
  - ❖ adjacent members differ by a –CH_2 group.
  - ❖ can be prepared by similar methods.
  - ❖ show a gradual change in physical properties.i.e. state at r.t.p, melting point, boiling point, density and solubility of members gradually changes **as molecular mass changes.**

Summary of the four homologous series.

Homologous series	Functional group	General formula	suffix
Alkane	-C-C-	C_nH_{2n+2}	-ane
Alkene	-C=C-	C_nH_{2n}	-ene
Alkanol (alcohol)	-OH	$C_nH_{2n+1}OH$	-anol
Alkanoic acid (carboxylic acids)	-COOH	$C_nH_{2n+1}COOH$	-anoic acid

12.12.6.1 Describe macromolecules.

- A giant (large) molecule formed by joining many small molecules (monomers).
- Two types are synthetic and natural macromolecules.

12.12.6.2 Describe synthetic macromolecules.

- These are human made (**man- made**) giant molecules (polymers).
- Examples include polythene, polystyrene, polyvinyl chloride (PVC) to mention but a few .These are made through **addition polymerization**.

12.12.6.3 Describe the formation of poly-alkenes.

- By addition polymerization e.g. polythene, polyvinyl chloride, polypropene, polystyrene etc.
- Named by simply adding poly in front of the monomer used.

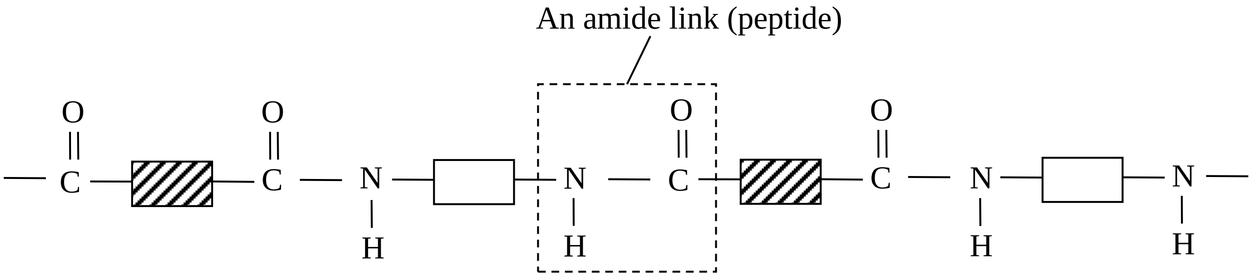
12.12.6.4 Classify plastics.

- **Thermoplastics**: These get soft runny when they are heated, but become hard when they are cooled.
- **Thermal set plastics**: These become permanently hard once they are heated.

12.12.6.5 Describe the formation of nylon and terylene.

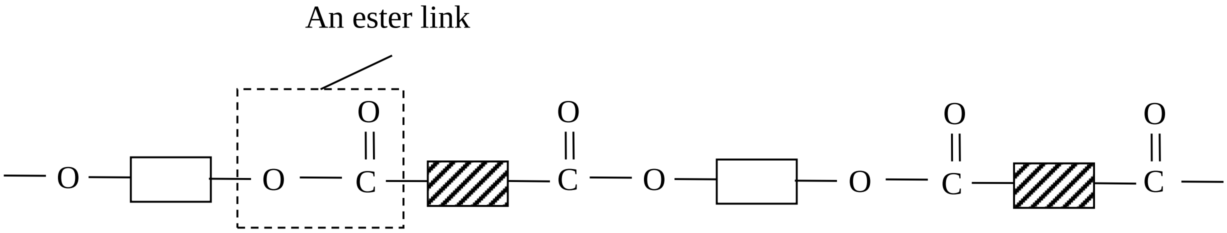
- Nylon is formed by condensation polymerization, this is where monomers combine to form long chain molecules (polymers) with loss of a small molecule like water, ammonia or hydrogen chloride leading to formation of amide linkages, -NHCO. Hence Nylon is a polyamide.
- Formed from diamines and dioic acids.

Structure of nylon



- Uses of nylon include:
  - ❖ Making clothes (fabrics).
  - ❖ Making ropes.
  - ❖ Making nets.
- Terylene formed from diols and dioic acids through condensation polymerization leading to the formation of ester linkages, -COOH. Hence Terylene is polyester.

Structure of Terylene



- Uses of terylene include:
    - Making clothes (fabrics).
    - Making ball bearings.
    - Making sails.
    - Making conveyor belts.

12.12.6.6 Differentiate between the structure of Nylon and Terylene.

- **Nylon** is a **polyamide** and **Terylene** is a **polyester**.

12.12.6.7 Describe the typical uses of plastics and synthetic fibres.

- Plastics are used as carrier bags, buckets, pipes, and apparatus.
- Nylon and terylene in making clothes, tents, strings, ropes, carpets, rubber reinforcement, sewing thread.

12.12.6.8 Describe the biodegradability of synthetic fibres.

- Biodegradable synthetic fibres are fibres that can be broken down by microorganisms (can rot).
- Non- biodegradable are synthetic fibres that cannot be broken down by microorganisms (cannot rot) e.g. plastics, Nylon and Terylene.
- Non-biodegradable polymers cause:
    - ❖ Litter problems when thrown about in cities and countryside.
    - ❖ Air pollution when burnt.

12.12.6.9 Describe natural macromolecules.

- Natural macromolecules are molecules that occur naturally (are made by God).Examples are:
    - ❖ Cellulose materials whose monomer is glucose e.g. wood, cotton. Used in making clothes and paper respectively.

- Natural rubber whose name is 2- methyl buta-1, 3 – di-ene (Isoprene) used for making tyres.
- Protein whose monomer units are amino acids. Used for body building.
- Fats and oils whose monomer units are fatty acids and glycerol, used in soap making, body insulation and energy source.
- Starch whose monomer is glucose used as energy source.

12.12.6.10 Describe the composition of carbohydrates.

- Carbohydrates contain carbon, hydrogen and oxygen in the form $C_xH_{2y}O_y$ where x is a multiple of six.

12.12.6.11 Identify the linkages in starch, proteins and fats.

- Starch is made from glucose molecules as monomer units. The glucose join together forming oxygen bonds (glycosidic bonds) through condensation polymerization.
- Proteins are made from amino acids as monomer units. The amino acids join together forming amide (or peptide) linkages through condensation polymerization.
- Fats are made from one glycerol and three fatty acids molecules. The four molecules join together forming ester linkages through esterification (not condensation polymerization).

12.12.6.12 Relate the linkages in synthetic and natural polymers.

Similarities:

- **Terylene** has the same as **Fats** which are **ester linkages, -COO.**
- **Nylon** has the same linkages as **Proteins** which are **amide linkages, -NHCO.**

Differences:

- The lengths of the hydrocarbon chain denoted by the boxes are different in both the natural and the synthetic macromolecules.
- Natural macromolecules are part of our diet (are edible) while synthetic macromolecules are not edible.
- Natural macromolecules are biodegradable (rot) while synthetic macromolecules are non-biodegradable (do not rot)

12.12.6.13Describe the hydrolysis of fats.

- **Hydrolysis** is a chemical reaction between water (or steam) with another reagent. Therefore, hydrolysis of fats involves reaction between fats and water. It produces Glycerol and the fatty acids. Occurs in the presence of an alkali such as Sodium hydroxide, hence known as alkali hydrolysis of fats.
- **Saponification** is reaction between **fats** and an **alkali** to produce **soap** and **Glycerol**. Hence Soap is a salt of fatty acids.

12.12.6.14 Identify the products of the hydrolysis of starch and proteins.

- Hydrolysis of starch produces simple sugars (monosaccharide). Hydrolysis occurs in the presence of an acid hence known as acid hydrolysis of starch. There are simple sugars among them are glucose and galactose. These are soluble hence are separated and identified using chromatography.

  Hydrolysis of proteins produces amino acids. There are over twenty amino acids, among them are glycine, alenine, cytocine to mention but a few. These are separated and identified by