1. MATTER

COTENT	LEARNING OBJECTIVE (Pupils should be able to)
1.1 The three states of matter	• Define matter • State the basic unit of matter • Discuss the three states of matter • Describe diffusion
1.2 Changes of state	• Describe the changes of state occurring when substances are heated or cooled • Determine the temperature at which these changes occur

1.1 The three states of matter

STATES OF MATTER

Matter is any substance that occupies **space** and has **mass**

There are different substances around us. All these substances are called **matter**.

Matter exists in different forms. These forms are called **states of matter**.

There are 3 states of matter: **solid, liquid** and **gas**.

[Examples of each state]

Solid – salt, wood and glass

Liquid – water, paraffin and oil

Air is a mixture of gases

Gas – hydrogen, oxygen and water vapour (steam)

The table below shows the characteristics of these 3 states of matter

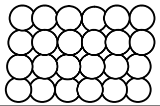
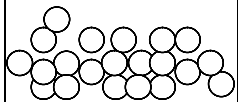
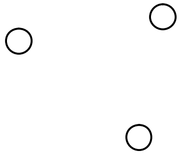
Table: Characteristics of the 3 states of matter

	Solids	Liquids	Gases
Shape	Fixed shape	NO fixed shape. Takes the shape of the container	No fixed shape. Takes the shape of the container
Volume	Fixed volume	Fixed volume	No fixed volume. Takes the volume of the container
Compressibility	Incompressible	Very slightly compressible negligible	Very compressible

KINETIC THEORY

Kinetic Theory is proposed to explain the characteristics of the three states of matter. It states that all matter is made up of extremely small **particles** that are in constant motion. These particles can be atoms, ions or molecules.

Table: Kinetic theory of matter

STATE	SOLID	LIQUID	GAS
Diagram of particles			
Arrangement of particles	Packed **closely**	Packed **loosely**	Spaced **widely**
Movement of particles	**Vibrate** about a fixed position	As well as vibrating, can move rapidly over short distances	**Move** at very high speeds in the space available
Forces between particles	Attractive and repulsive forces counterbalance	Attractive forces are not strong enough to hold particles in a regular pattern	Forces between particles are negligible

DIFFUSION

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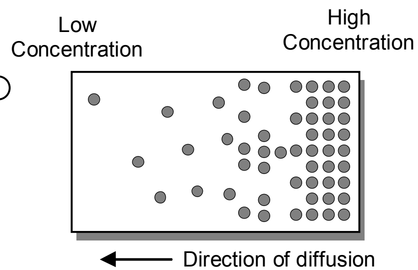
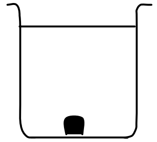
Diffusion is the movement of particles **from** an area of **high** concentration **to** an area of low concentration

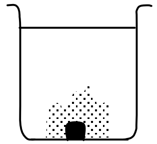


[Experiment]

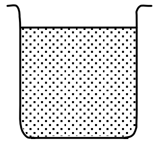
A crystal of copper sulphate(II) is put in a beaker filled with water. Leave the beaker undisturbed and observe carefully.



As the crystal dissolves the colour slowly spreads through the liquid, first covering the bottom.



Eventually the colour distributes itself evenly throughout the liquid.



It is the copper sulphate(II) particles which slowly move from an area of high concentration to an area of low concentration. This is diffusion in liquid.

Three factors which can affect the rate of diffusion

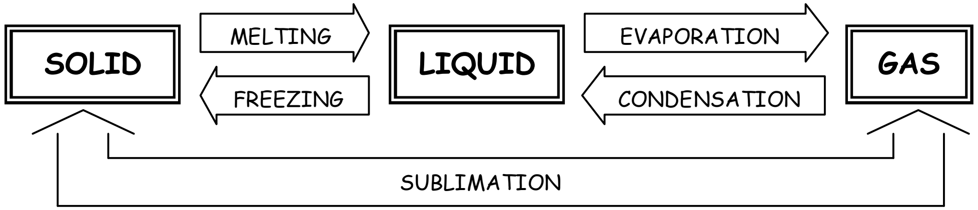
1. The higher the **temperature** is, the faster the diffusion is.
2. The smaller the **size of particle** is, the faster the diffusion is.
3. The larger the **concentration** gradient is, the faster the diffusion is.

1.2 Changes of state

A change of state is a change where one state changes to another

There are some types like **Melting, Evaporation/Boiling, Freezing/Solidification, Condensation** and **Sublimation**.

Physical change can easily reverse and produce NO new substance
E.g. Melting, Evaporation, Condensation
Chemical change can not easily reverse and produce new substances
E.g. Combustion, Decomposition



SUBLIMATION

Sublimation is the direct change of state from a solid to a gas on heating or from a gas to a solid on cooling.

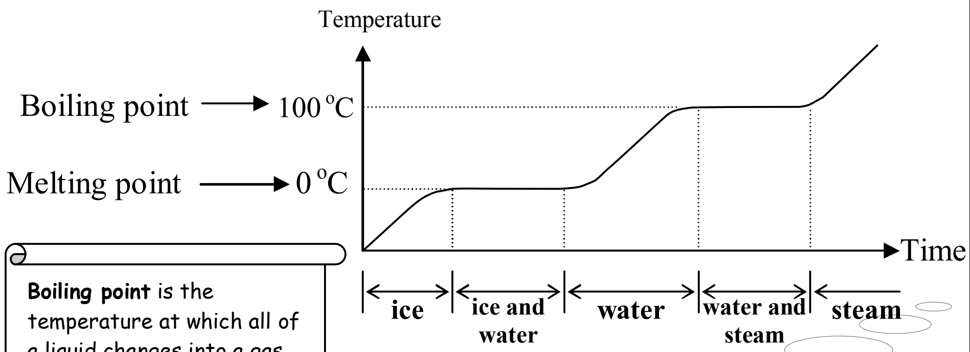
Substances which **sublime** are

- **iodine**, **ammonium chloride** (NH_4Cl),
- **ammonium sulphate** ($(NH_4)_2SO_4$)
- **carbon dioxide** (CO_2, called *dry ice*)

HEATING CURVE

A **heating curve** is a graph showing changes in temperature with time for a substance being heated

The graph below is a heating curve for a substance of water



Boiling point is the temperature at which all of a liquid changes into a gas.
Melting point is the temperature at which a solid completely changes into a liquid.

A **pure substance** has an exact boiling and melting point. **Impurities** raise boiling points and cause lower melting points.

As a substance is heated, it absorbs energy and its temperature rises. Then it changes from a solid to a liquid and finally to a gas.

As you can see in the graph above, there are 2 types of sections; **Slope** and **Flat**

The **flat sections** on the graph indicate the melting and boiling points. Here the **temperature remains the same** over a period of time, as the heat energy is being used **to change the state** of the substance.

Heat energy can be used either to **raise the temperature** of a substance or to **change the state** of it

BOILING AND EVAPORATION

Boiling and evaporation are both physical process that change a liquid into a gas. The liquid absorbs heat energy during these changes in state.

These **must be differentiated** with each other. The table below shows the differences between these 2 processes.

Table: Differences between boiling and evaporation

Boiling	**Evaporation**
Occurs at boiling point	Occurs at any temperature below boiling point
Occurs throughout the liquid	Occurs only at the surface of the liquid
Bubbles observed	No bubbles observed
Occurs rapidly	Occurs slowly

2. EXPERIMENTAL TECHNIQUES

Most of materials we meet in our environment are mixtures. Often, only one substance from a mixture is needed, so it has to be separated from the mixture by physical means.

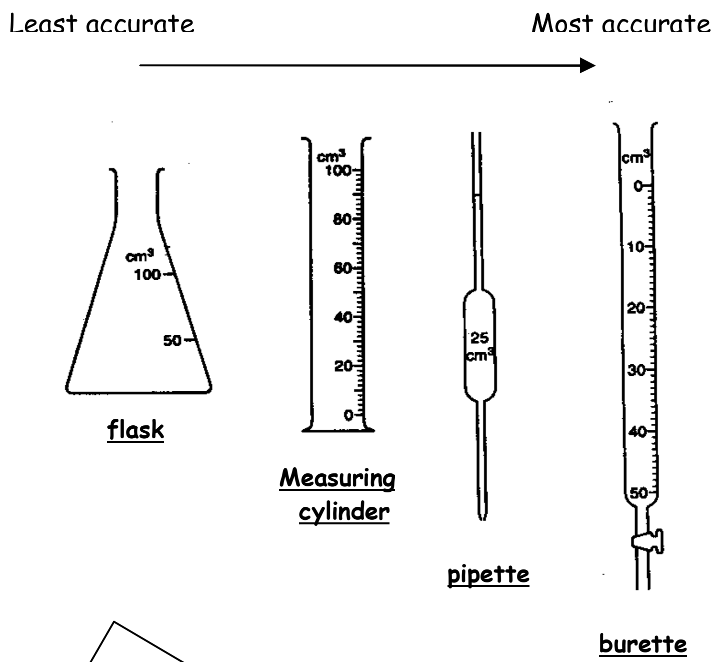
There are many industries in Zambia which produce a variety of products. During the production of any of these products the industry begins with impure raw materials that are often mixtures. The final product has to be extracted from the raw materials by using some of the techniques we are going to learn in this topic.

COTENT	LEARNING OBJECTIVE (Pupils should be able to)
Measurement	• Name and use appropriate apparatus for the measurement of time, temperature, mass and volume, including burettes, pipettes and measuring cylinders • Design arrangements of apparatus, given information about the substances involved
Method of purification	• Describe and use methods of purification by the use of suitable solvent, filtration, crystallisation, distillation. • Suggest suitable purification techniques, given information about the substances involved. • Describe and use paper chromatography and interpret chromatograms. • Identify substances and test their purity by melting point and boiling point determination and by paper chromatography.

2.1 Measurement

VOLUMES OF LIQUID AND GASES

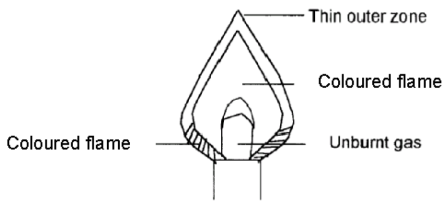
There are some types of apparatus to measure the volumes of liquids. They have differences in accuracy.



A Burette is a long vertical graduated glass tube with a tap at one end which is used to add controlled volumes of liquids with accuracy up to $0.1cm^3$.
A Pipette is a graduated tube which is filled by suction and used for transferring exact volumes ($10cm^3$, $25cm^3$) of liquids.
A Measuring cylinder is used to measure the approximate volume of liquids.
A flask/beaker is used only for estimating volumes of liquids.

HEATING

In experiment, a **Bunsen burner** is usually used for heating. You can identify the temperature of the flame from its colour

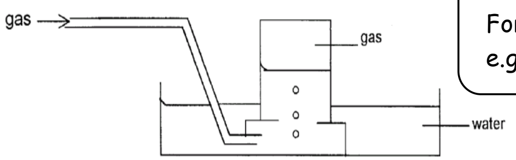


Yellow flame (luminous flame) → Air hole is closed. It **produces pollutant gas** like carbon monoxide.
Blue flame (non-luminous flame) → Air hole is half open. It is **most generally used**.
Blue-green flame → Air hole is completely open. It is used for **strong heating**.

COLLECTION OF GASES

2 factors determine the method used to collect a gas: the **density** of the gas and the **solubility** of the gas in water.

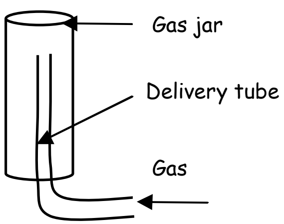
Method of collection
Displacement of water



Type of Gases to be collected

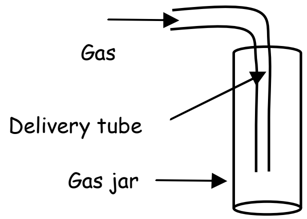
For gases that are **insoluble** in water, e.g. hydrogen, oxygen.

Method of collection	**Method of collection**
Displacement of air –upward delivery	Displacement of air -downward delivery



Type of Gases to be collected

For gases that are **less dense than air**, e.g. hydrogen, ammonia.



Type of Gases to be collected

For gases that are **denser than air**, e.g. hydrogen chloride, carbon dioxide.

Carbon dioxide gas is sometimes collected by displacement of water. Carbon dioxide is sparingly soluble in water to form carbonic acid. Thus, the volume of gas collected will be less than expected because some carbon dioxide will dissolve into the water.

2.2 Method of purification (*Separation techniques)*

There are some techniques to separate mixture. How to determine the method to separate it depends on some physical properties such as solubility, density and so on.

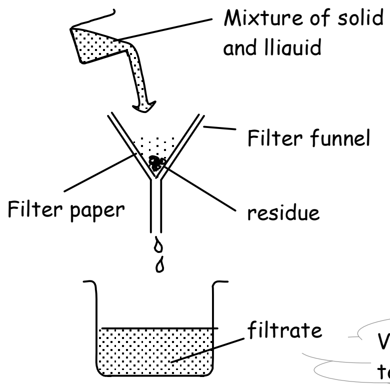
A Solution is a mixture in which the particles of solute and solvent are evenly spread out.
A Solute is the substance which dissolves in the solvent to form a solution.
A Solvent is a liquid that dissolves substances.

[e.g.] Salt water =	Salt	+	Water
(Solution)	(Solute)		(Solvent)

A **mixture** is a material formed by two or more different substances which are physically combined together

FILTRATION

→Used to separate out an **insoluble solid from a liquid**, e.g. separating sand form sand water and water mixture.



Residue is the solid trapped in the filter during filtration.
Filtrate is the clear liquid that passes through the filter during filtration.

Very fine pores in the **filter paper** allow small particles to flow through, but retain the large particles.

CRYSTALLISATION

→ used to separate out a **pure solid from an impure solution**, e.g. separating copper (II) sulphate crystals from impure copper (II) sulphate solution. The impurities will remain dissolved in solution.

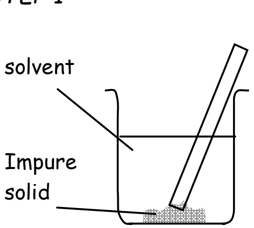
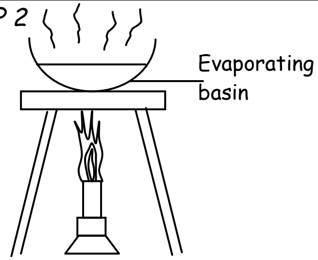
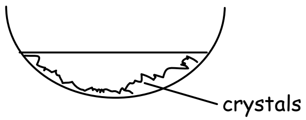
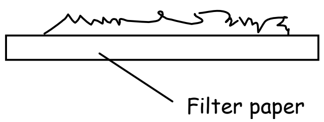
STEP 1  The impure solid is dissolved in a solvent.	*STEP 2*  The solution is heated to evaporate most of the solvent.
STEP 3  The hot solution is allowed to cool. The solid appears as pure crystals.	*STEP 4*  The cold solution is poured off to obtain the crystals. The crystals may be dried by pressing them between sheets of filter paper.

Figure: crystallisation

Crystallisation **must be different** from evaporation. In crystallisation, the solvent is **partially** evaporated, leaving small amount of solution in which the crystals form. Impurities are left behind in the solution when the crystals are filtered off. In evaporation, **all** the solvent is removed. The crystals formed may be impure.

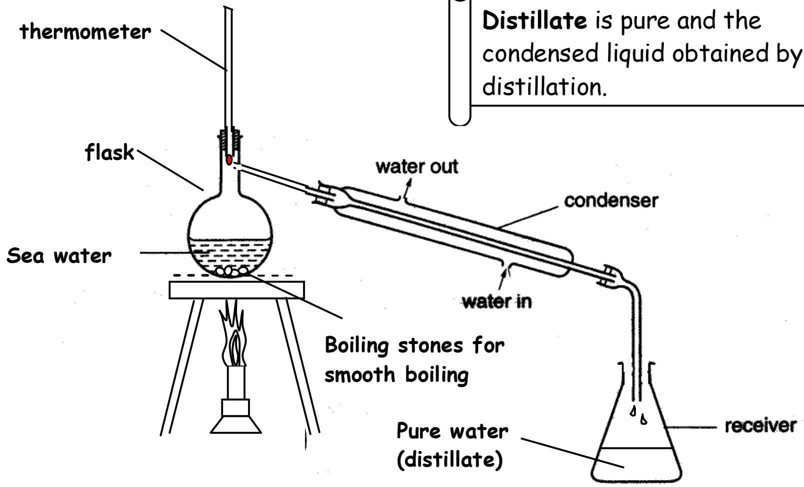
DISTILLATION

Distillation is conducted using **EVAPOTATION** and **CONDENSATION** for separation.

SIMPLE DISTILLATION

→ used to separate a **pure liquid from a solution** containing dissolved solids, e.g. separating pure water from seawater.

Distillate is pure and the condensed liquid obtained by distillation.



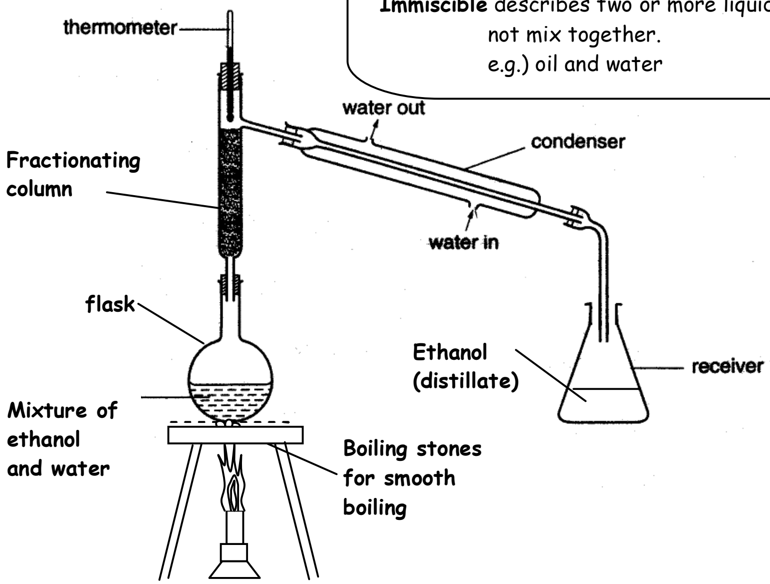
A thermometer is placed at the mouth of the condenser to measure the temperature of the vapour entering it. This temperature is the boiling point of the distillate.

FRACTIONAL DISTILLATION

→ used to separate a **pure liquid from miscible liquids**.
e.g. ethanol from a mixture of ethanol and water.

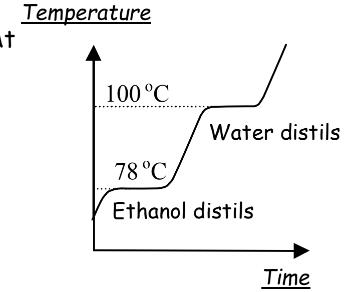
Miscible describes two or more liquids which will diffuse together and form a single phase.
e.g.) alcohol and water

Immiscible describes two or more liquids that will not mix together.
e.g.) oil and water



Fractional distillation separates **according to boiling points**.
The liquid with the **lowest** boiling point will be distilled **first**, followed by the liquid with the next lowest boiling point. As a rough guide, the boiling points of the liquids to be separated should be at least 20° C apart.

◆The graph shows the change in temperature as the mixture of **ethanol** and **water** is being heated in the flask. The temperature will remain at 78 °C when the first ethanol is being collected. When all the ethanol has evaporated over, the temperature will rise again until it reaches 100 °C. At this temperature, water will be collected as the second distillate.



Fractional distillation is also used to separate
1 the components of **crude oil**
2 **fermented liquor** to obtain alcoholic drinks of a higher concentration.
3 components of air

How to separate "Insoluble and Soluble solid".

[EXAPMPLE] Separate a mixture of sand and salt.
-The common procedure is …

1. Dissolve
Place the mixture of sand and salt in a beaker, add water and stir.
→ Salt will dissolve while sand will not dissolve because of their solubilities.

2. Filtration
Pour the liquid of the mixture along a glass rod in to the funnel with a filter paper.
→ Salt water will pass through the filter paper as a filtrate while Sand will be trapped on the paper as a residue

3. Evaporating
Put a little amount of the filtrate into the evaporating dish. Heat the filtrate until all the water is driven off.
→ Salt(solute) will remain in the dish while water(solvent) will go away as steam.

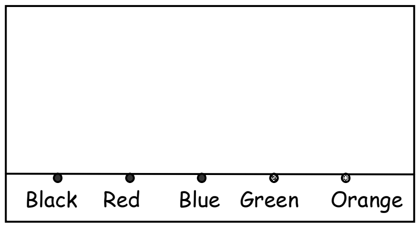
Now,
Salt and sand have been separated

Steps to Separate Insoluble and Soluble Solids:
1. Dissolve 2. Filtration 3. Evaporating

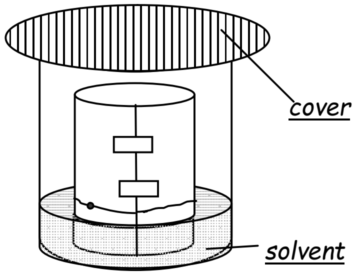
PAPER CHROMATOGRAPHY

→Substances in a mixture are separated according to their **solubility** in the same solvent. The **more soluble** component in the mixture will tend to remain in the solvent and travel further **up** the chromatogram, while the **less soluble** component will separate out **onto** the paper.

Paper chromatography is a method of separating dissolved substance, such as dyes and pigments by spreading them on absorbed paper with a sutable solvent.



Drops of ink on the pencil line



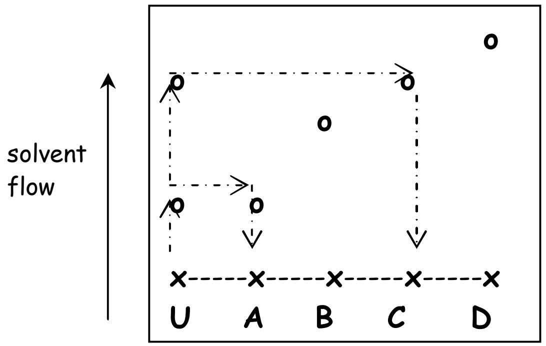
PROCEDURE

1 Use a **pencil** to draw the start line.
2 Use the **black** ink sample to make a small dot on the start line, together with some other coloured ink to use as reference.
3 Fold the paper into a cylinder and place it into a beaker containing the solvent, ensuring that the start line is **above** the solvent level. Cover the beaker while the chromatogram develops.
4 Remove the chromatogram from the beaker just **before** the solvent reaches the top of the paper..

◆ **If the start line is drawn in ink**, the components in the ink will also separate out together with the sample dots when the chromatogram is run.
◆ **If the start line is below the solvent level**, the sample dots will dissolve into the solvent in the beaker instead of travelling up the chromatogram paper.
◆ **The beaker must be covered** when the chromatogram is run to reduce evaporation of the solvent from the beaker and to prevent the solvent from evaporating off the paper as it moves up.

INTERPRETATION OF RESULTS

1. From unknown dye, follow direction of solvent flow until you find a spot
2. Move across the chromatogram until you find a corresponding spot
3. Move in the other direction to identify the known dye



Unknown Dye U = Mixture of Dyes A + C

It is possible for **different substance** to travel **the same distance** in the same solvent in paper chromatography, i.e. have **the same solubility in the same solvent**.
To confirm the identity of a substance, **another round** of chromatography is carried out **using a different solvent**. If the spot still travel the same distance, then the spot must contain the same substance.

SUMMARY of SEPARATION TECHNIQUE

SEPARATION TECHNIQUE	SUBSTANCES TO BE SEPARATED	EXAMPLE
Filtration	• **Insoluble solid** and liquid	*Muddy water*
Crystallisation / Evaporation	• **Solute** (soluble solid) from its solution	*Salt solution*
Distillation	• **Solvent** from its solution	*Salt solution*
Fractional Distillation	• **Miscible liquids** with different boiling points	*Ethanol and water* *Crude oil* *Liquid air*
Decantation / Sedimentation	• Insoluble suspension settles to form sediment	*Mealie-meal and water*
Separating Funnel	• **Immiscible liquids**	*Oil and water*
Floatation	• **Less dense** solid and liquid	*Charcoal dust and water*
Magnetic Separation	• Magnetic materials	*Iron filings and sulphur powder*
Paper Chromatography	• Dissolved substances	*Dyes and pigment of ink*

3. ATOMS, ELEMENTS AND COMPOUNDS

COTENT	LEARNING OBJECTIVE (Pupils should be able to)
Atoms	• State the relative charges and approximate relative masses of protons, neutrons and electrons. • Define proton number (atomic number) and nucleon number (mass number). • Use and interpret such symbols as ${}^{12}_{6}C$ • Use proton (atomic) number and the simple structure of atoms to explain the periodic table, with special reference to the elements of proton (atomic) number 1 to 20. • Define isotopes • Describe the build-up the electrons in 'shells' and explain the significance of valency electrons and the noble gas electronic structures.
Chemical Bonding	• Describe the formation of ionic bonding between metallic and non-metallic elements, e.g. $NaCl$, $CaCl_2$ • Describe the formation of covalent bonds as the sharing of pairs of electrons leading to the noble gas configuration. • Deduce the electron arrangement in other covalent molecules. • Construct 'dots and cross' diagrams to show the valency electrons in covalent molecules. • Describe the differences in volatility, solubility and electrical conductivity between ionic and covalent compounds.
Structures and Properties of materials	• Describe the differences between elements, compounds and mixtures and between metals and non-metals.

In kinetic theory we saw that matter consisted of particles. We looked at how these particles account for the differences in the physical properties of solids, liquids and gases.

Now we shall consider all substances as chemical substances

3.1 ATOMS

Are you sure the meaning of terms like atoms, molecules, elements, compounds and mixtures. Can you distinguish them clearly? Here are definitions for them.

An **ATOM** is the smallest particle of an element which can take part in a chemical reaction and remain unchanged.

A **MOLECULE** is the smallest particle of an elements or a compound which exists independently

An **ELEMENT** is a substance that cannot be broken down into two or more simpler substances by chemical means.

A **COMPOUND** is a substance that consists of two or more elements which are chemically combined in fixed proportions.

A **MIXTURE** is a substance that consists of two or more substances which are not chemically combined

As you have seen they are considered in terms of substances while they are considered in terms of particles.

ATOMS and MOLECULES

In kinetic theory, we saw that matter consists of particles. What are these particles? It is very useful to know about them.

Molecules can be thought of to be 3 types.

1. It consists of those elements in which a single atom forms the molecule.
- These molecules are called monatomic molecule.
[e.g.] helium, neon and argon which are known as the noble gases

2. It consists of atoms of the same element combined together.
- These molecules are called diatomic molecule.
[e.g.] oxygen, hydrogen, nitrogen and chlorine

3. It consists of atoms of different elements combined together.
Here the atoms form molecules of compounds.
[e.g.] carbon dioxide, water and sugar

Now we shall see more about an atom!

STURECTURE OF AN ATOM

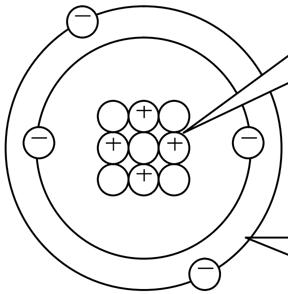
Atoms are made up of three fundamental particles. These are the **proton**, the **neutron** and the **electron**

	Charge	Mass	Position in Atom
PROTON	+	1	nucleus
NEUTRON	0	1	
ELECTRON	-	1/2000	electron shells

Properties of particles found in an atom

- Most of an atom is empty space.
- An atom is electrically neutral.

The protons and neutrons cluster together in the centre, forming the **nucleus**; this is the heavy part of the atom and positive charged.



The electrons circle very fast around the nucleus, at different levels from it along **electron shells**.

These particles are extremely small. If a golf ball were magnified to the size of the Earth, then an atom would be the size of a marble! They have a radius of around 10^{-10} m and a mass of about 10^{-22}g

NUCLIDE NOTATION

One convenient method of writing the names of elements is by applying a shorthand system in which each element is assigned a specific symbol.

$^{A}_{Z}X$

Mass Number (or Nucleon Number) is the total number of protons and neutrons found in the nucleus of an atom.

Atomic Number (or Proton Number) is the number of protons an element has in the nucleus of its atom.

A chemical symbol of an element is a letter or letters derived from the name of the element.

Information about the number of particle in an atom can be found from the **periodic table.**

From Nuclide Notation many information of an atom can be obtained.

- Number of **Neutrons** = **Mass Number – Atomic Number**
- Number of **Electrons** = Number of Protons = **Atomic Number**
  (So that **Atoms** should be **electrically neutral)**

[Example] What are the mass number, proton number and number of neutron in Aluminium, hydrogen and lead?

	Aluminium	Hydrogen	Lead
Form periodic table	$^{27}_{13}Al$	$^{1}_{1}H$	$^{207}_{82}Pb$
Mass number	27	1	207
Proton number	13	1	82
No. of neutrons	27 – 13 = 14	1 - 1 = 0	207 – 82 =125

$^{35.5}_{17}Cl$

NOTICE from the above example that the number of neutrons is obtained from the periodic table using the top number minus the bottom number, however, this is not true for all cases. For example, chlorine in the periodic table is represented as shown left; this does not mean that the chlorine atom contains 35.5-17 = 18.5 neutrons! Chlorine contains 2 isotopes (see later in the chapter) and to account for these 2 isotopes, its mass number is a calculated average value of 35.5

ISOTOPES

Many elements contain atoms that are slightly different from each other.

> **ISOTOPES** are different atoms of the same element which have the **same** number of **protons** but **different** number of **neutrons**.

It is known that 3 isotopes of hydrogen exist, also isotopes of carbon are known.

Element	Isotopes	No. of Protons	No. of Neutrons	No. of Electrons
Hydrogen	$^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$	1 1 1	0 1 2	1 1 1
Carbon	$^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$	6 6 6	6 7 8	6 6 6

Table: Isotopes of some elements

ELECTRONIC SHELLS

Electrons are arranged in electronic shells around nucleus

The number of electrons in each shell is finite as shown below

Shell(from a nucleus)	1st	2nd	3rd	4th .....
Maximum number	2	8	8	....

Table : Maximum No. of electrons that can occupy the shell

The number of electons in each shell is shown by the **elctronic configuration**

ATOM	**Electronic Configuration**
Lithium	2:1
Potassium	2:8:8:1

Table: Electronic Configuration of some atoms

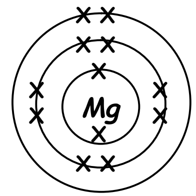
DRAWING of ELECTRONIC STRUCTURE

The number of electrons in each shell can also be shown by drawing as the followings

<u>Rules to draw electronic structure:</u>

[e.g.] $^{24}_{12}Mg$

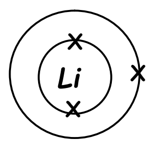
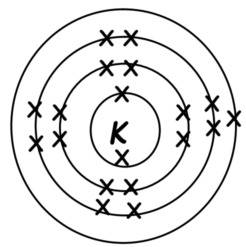
$12e^{-}$ = 2, 8, 2



- Use the symbol of the element to represent the nucleus
- Represent each e^{-} with a cross
- Each new circle represents another electron shell
- Start filling up the shells from the first shell before going on to the next shell
- Group the electrons in pairs for easy counting

> - **Nearly full** shells want to **get extra e^{-}** - gain or share (NON-METALS)
> - **Nearly empty** shells want to **lose extra e^{-}** (METALS)
> - **Valency** = **No. of e^{-}** an atom wants to **lose**, **gain** or **share**

Electrons found in the outer shell are called **Valence electrons (v.e.)**

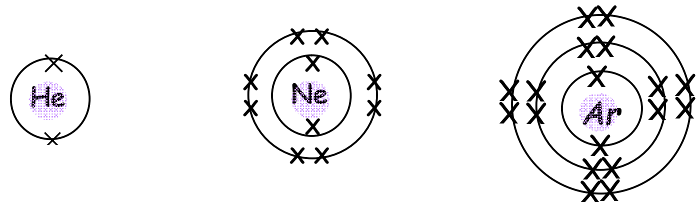
[Example 1] Lithium, Li	[Example 2] Potassium, K
The lithium atom contains 3 electrons arranged in 2 shells.  	The potassium atom contains 19 electrons arranged in 4 shells 

3.2 CHEMICAL BONDING

When atoms combine together in a chemical reaction, we say that **a bond is formed** between the atoms during the reaction. A reactive atom will combine or form bonds with other atoms easily, while an unreactive atom will not.

WHY DO ATOMS FORM BONDS?

→ Atoms of noble gases possess the maximum number of electrons in their outermost shell as shown in the diagram below.



Helium, $^{4}_{2}He$ **Neon, $^{20}_{10}Ne$** **Argon, $^{40}_{18}Ar$**

All the outermost shells (or valence shells) are completely filled. This type of arrangement is very **stable** and highly **unreactive**.

→ Atoms of other elements have incompletely filled outermost shells, as a result, these atoms are **unstable** and therefore, **reactive**.
Most of them tend to become stable like noble gases.

In order to achieve stable electronic structure, they **share**, **gain** or **lose electrons** in their outer electronic shells.

This is why atoms form bonds.

There are 2types of chemical bonds that can be formed between 2 atoms:

1. **Ionic bonds** – valence electrons are transferred from one atom to another
2. **Covalent bonds** – valence electrons are shared

<u>**IONI BONDING (ELECTROVALENT)**</u>

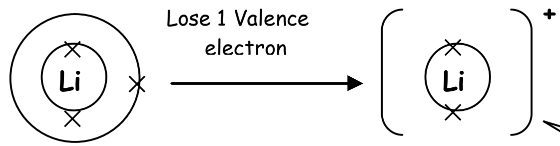
FORMATION OF IONS

Atoms can obtain a full outer shell and become stable when they lose or gain valence electrons. Charged particles called **ions** are formed.

<u>1. FORMATION of POSITIVE IONS</u>

[Example 1] Lithium $^{7}_{3}Li$

Note that **1+** is written simply as **+**. The 1 is not written.

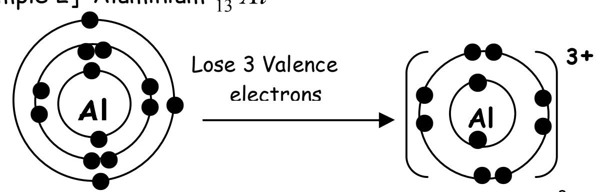


We do not draw empty shells either.

<u>Lithium atom, Li</u>	<u>Lithium ion, Li^{+}</u>
3protons	3protons
3electrons	3-1= 2electorns
7-3 = 4 neutrons	4 neutrons
Overall zero charge, i.e. neutral	**Overall 1+ charge**

The lithium ion now carries a 1+ charge because it has an extra proton. This is represented by enclosing the ion in brackets and writing its charge on the top right hand corner.

[Example 2] Aluminium $^{27}_{13}Al$

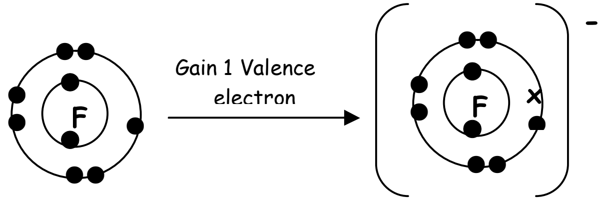


<u>Aluminium atom, Al</u>	<u>Aluminium ion, Al^{3+}</u>
13protons	13protons
13electrons	13-3= 10electorns
27-13 = 14 neutrons	14 neutrons
<u>Overall zero charge, i.e. neutral</u>	**<u>Overall 3+ charge</u>**

In general, when an atom **loses n** valence electrons to form a stable ion, the ion formed will carry **n+** charge.

2. FORMATION of NEGATIVE IONS

[Example 3] Fluorine, $^{19}_{9}F$

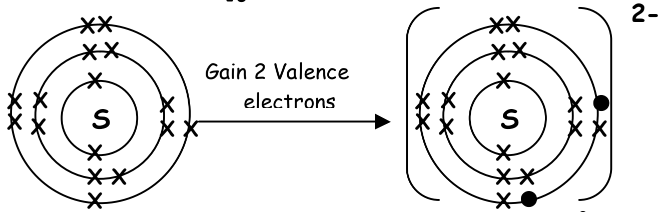


Fluorine atom, F	Fluorine ion, F^-
9protons	9protons
9electrons	9+1= 10electorns
19-9 = 10 neutrons	10 neutrons
Overall zero charge, i.e. neutral	**Overall 1- charge**

The fluoride ion now carries a 1- charge because it has 1 extra electron.

Note: Non-metal name changes to end "*-ide*"
[e.g.] chlorine → chloride   oxygen → oxide,
sulphur → sulphide   Phosphorus → phosphide etc.

[Example 4] Sulphur atom, $^{32}_{16}S$



Sulphur atom, S	Sulphur ion, S^{2-}
16protons	16protons
16electrons	16+2= 18electorns
32-16 = 16 neutrons	16 neutrons
Overall zero charge, i.e. neutral	**Overall 2- charge**

In general, when an atom **gains m** electrons to form a stable ion, the ion will carry **m-** charge.

FORMATION OF IONIC BONDING

- Ionic bonding occurs between **metallic and non-metallic atoms**.
- Valence electrons are **transferred** from the metallic atom to the non-metallic atom so that both atoms achieve a full outer shell and become stable.
- Oppositely charged ions are formed. The metal ion carries positive charge, while the non-metallic ion carries negative charge. These ions **attract** each other with strong **electrostatic forces** to form an ionic bond.

Rules for '*Dot and Cross*' Diagrams for Ionic Bonding:

- Calculate *how many e^- metal* atom wants to *lose*
- Calculate *how many e^- non-metal* atom wants to *gain*
- To find out how many metal and non-metal atoms combine, use:

  Total No. of e^- lost = Total No. of e^- gained

  → No. of Metal Atoms × No. of e^- lost = No. of Non Metal Atoms × No. of e^- gained

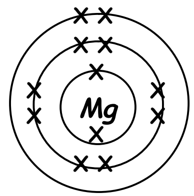
- Draw each atom's e^- shells BEFORE losing or gaining
  - → Metal e^- with a '*cross*' **x**
  - → Non-metal e^- with a '*dot*' **o**
- Re-draw each atoms' shells AFTER losing or gaining
  - → REMEMBER that the *metal e^-'s GAINED by non-metal* are still a '**x**'
  - → Check: Every ion should now have FULL OUTER e^- shells
  - → INDICATE the CHARGE on each ion

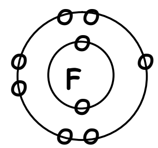
[EXAMPLE]

Magnesium fluoride contains magnesium and fluorine atoms. Draw the dot and cross diagram to represent magnesium fluoride.

ANSWER: *Mg: 12 e^- = 2, 8, 2*   *F: 9 e^- = 2, 7 { Periodic Table }*

BEFORE:

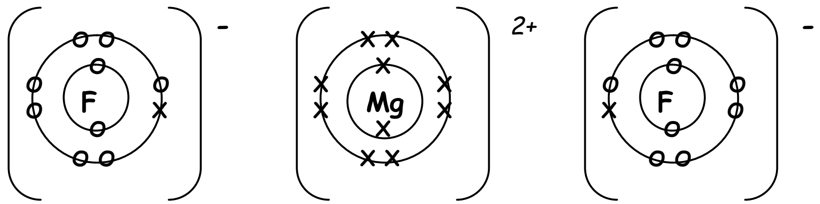




→ lose 2 e^- → Mg^{2+} ion

→ gain 1 e^- → need 2 F atoms (to gain $2e^-$) → $2F^-$ (fluoride) ions

AFTER:



- The formula of the compound formed is written as **MgF_2**.
- Ionic compounds are electrically **neutral**, i.e. once the positive and negative ions combine, these charges '**cancel**' each other out to give a neutral compound.

It is not necessary to draw and show the movement of valence electrons from one atom to another unless the question requires it. Most exam question will ask for the final structure of the compound only.

COVALENT BONDING (MOLECULAR)

FORMATION OF COVALENT BONDING

- Covalent bonding usually takes place between **non-metallic atoms**.
- Valence electrons are **shared** between these atoms.
- The molecules of the compound are held together by **weak intermolecular forces** that are easily broken by heating.

Rules for '*Dot and Cross*' Diagrams for Molecular Bonding:

- For a molecule of an *element*: Represent the e^- of each atom with a 'dot' or a 'cross'
- For a molecule of a *compound*: Represent the e^- of each element with a 'dot' or a 'cross'
- (Usually) Only draw the OUTER e^- shell
- Draw the SHARED e^- FIRST
- Then ADD the REMAINING e^- to the non-shared section for EACH ATOM

→ Check: Each atom should have the correct number of e^- for their shell

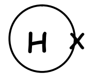
→ Double-check: Every atom should now have FULL OUTER e^- shells

[EXAMPLE]

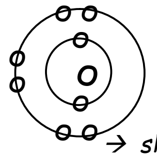
Water is a molecular compound containing hydrogen and oxygen atoms. Draw the dot and cross diagram to represent a molecule of water. Show all the electron shells.

ANSWER: *H: 1 e^- = 1* *O: 8 e^- = 2, 6 { Periodic Table }*

BEFORE:



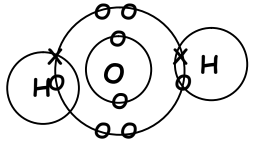
→ shares 1 e^-



→ shares 2 e^-

→ need 2 x H atoms to share with 1 x O atom

AFTER:



Water molecule, H_2O

Reaction: $H_2 + O_2 \rightarrow H_2O$

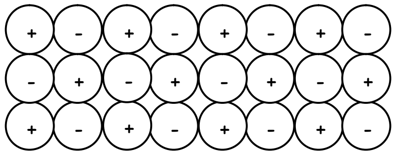
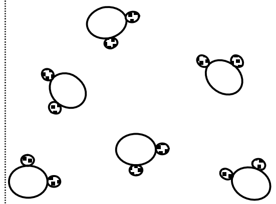
Structural formula: *H – O – H*

When drawing covalent structures, always draw the atom that needs to form the most number of bonds in the centre, then add on the rest of the atoms

DIFFERENCES in PROPERTIES of IONIC and COVALENT BONDING

IONIC (Electrovalent)	COVALENT (Molecular)
• **Ionic** compounds can **conduct electricity** when **molten** or **aqueous** because the **ions** are **free to move**	• **Molecular** compounds **do not conduct electricity** in any form
• **Ionic** compounds have **high MP** and **BP** due to the **strong electrostatic forces** between charged ions	• **Molecular** compounds have **low MP** and **BP** due to **weak inter-molecular forces**
• **Soluble** in water, **insoluble** in organic solvents	• **Soluble** in organic solvents (E.g. ethanol, petrol) **Insoluble** in water

Table below summarises properties

Property	Type of Bonding	
	IONIC (Electrovalent)	**COVALENT (Molecular)**
Conduct Electricity	YES (when AQUEOUS or MOLTEN)	NO
MP and BP	HIGH	LOW
Volatility	NON-VOLATILE	VOLATILE
Usual State (at room temp)	SOLID	GASES OR LIQUIDS
Composition	IONIC LATTICE	MOLECULES
Diagram		
Examples	NaCl, CaO, MgF_2	H_2O, CO_2, O_2

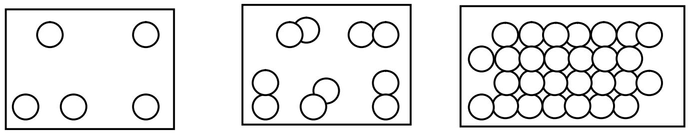
3.3 STRUCTURE AND PROPERTIES OF MATERIALS

ELEMENTS, COMPOUNDS AND MIXTURES

ELEMENTS

Elements are made up of only one kind of atoms. The diagram below shows examples of elements existing as atoms as well as molecules.

(a) (b) (c)



(a) a monatomic gaseous element made up of atoms, e.g. helium
(b) an gaseous element made up of diatomic molecules, e.g. hydrogen, oxygen, nitrogen etc.
(c) a solid element, e.g. iron, copper, etc.

COMPOUNDS

A compound is made up of two or more types of atoms **chemically combined** together. It can not be separated using physical means. Chemical means such as electrolysis are needed

MIXTURES

A mixture is made up of two or more elements or compounds **physically combined** together. The components can be separated easily form one another using physical means such as filtration, a magnet, distillation, etc.

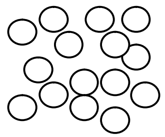
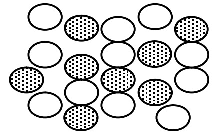
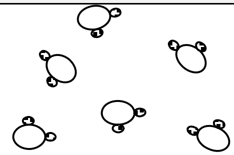
	ELEMENTS	**MIXTURES**	**COMPOUNDS**
Diagram			
Melting and Boling Points	Fixed melting and boiling point	Melts and boils over a range of temperature	Fixed melting and boiling point
Separation	-----	Easily separated using physical means such as distillation	Chemical means such as electrolysis are needed
Examples	• Copper • Iron • Oxygen • Nitrogen • Carbon	• Air • Sea water • Metal alloys • Rock salt	• Carbon dioxide • Water • Common salt • Ethanol

Table: the differences of Elements, Mixtures and Compounds

4. STOICHIOMETRY

COTENT	LEARNING OBJECTIVE (Pupils should be able to)
Formulae and Equations	• State the symbols of the elements and formulae of compounds. • Deduce the formula of a simple compound from the relative numbers of atoms and vice versa. • Determine the formula of an ionic compound from the charge on the ions present and vice versa. • Construct equations with sate symbols, including ionic equations. • Deduce, from experimental results of the identity of the reactants and products, the balanced chemical equation for a chemical reaction.
Stoichiometric calculations	• Define relative atomic mass, Ar. • Define relative molecular mass, Mr. • Perform calculations concerning reading masses reacting masses using simple proportions.

4.1 FORMULAE and EQUATIONS

It is useful to know the names of elements and compounds, how these names can be represented and how chemical changes involving elements and compounds maybe described.

CHEMICAL FORMULAE

One convenient method of writing the names of compounds is by using chemical formulae.

There are some types of chemical formulae. Here we talk about **Molecular formulae.** Later we will see **Empirical formulae** and **Structural formulae**

A **chemical formula** is a way of showing the proportions of elements present in a chemical compound using symbols for the atoms present.

Rules for Chemical Formulae:

It shows the number of **atoms**

- A SMALL number multiplies ONLY the elements or radicals to the LEFT

- A **BIG** number at the FRONT multiplies ALL the elements in the formulae

It shows the number of **molecules**

[EXAMPLE]

- H_2O represents **2** atoms of **hydrogen** and **1** atom of **oxygen** in **1** molecule of **water**
- $2CO_2$ represents **1** atom of **carbon** and **2** atoms of **oxygen** in **2** molecules of **carbon dioxide**
- $4Na_2SO_4$ represents **2** atoms of **sodium**, **1**atom of **sulphur** and **4** atoms of **oxygen** in **4**molecules of **sodium sulphate**.

[EXAMPLE 2]

How many atoms are represented by $2Al_2(SO_4)_3$?

ANS: $\mathbf{2} \times ({}_2 \times Al + {}_3 \times (S + {}_4 \times O)) = 2 \times (2 + 3 \times (1 + 4)) = 34$ *atoms*

CHEMICAL FORMULAE FOR ELEMENTS

1 **Metals** exist **as atoms**. The chemical formula for a metal is its symbol.
[EXAMPLE] Sodium → Na   Magnesium → Mg   Iron → Fe

2 Most **non-metals**, with the exception of the noble gases, exist **as molecules**. Its chemical formula will show both the symbol as well as the number of atoms that make up the molecule.
[EXAMPLE] Hydrogen → H_2,

Where the subscript '2' shows that the molecule is made up of two hydrogen atoms joined together.

Noble gases exist **as atoms**. The chemical formula for a noble gas is thus its symbol.
[EXAMPLE] Helium → He   Neon → Ne.

Table; Chemical formulae for some common elements

Metallic Element	Chemical Formula	Non-Metallic Element	Chemical Formula
Calcium	Ca	Chlorine	Cl_2
Zinc	Zn	Oxygen	O_2
Copper	Cu	Nitrogen	N_2
Lead	Pb	Carbon	C
Manganese	Mn	Sulphur	S
Mercury	Hg	Argon	Ar

CHEMICAL FORMULA for COMPOUNDS

IONIC COMPOUNDS

The formulae of both the positive **ion** and the negative **ion** must be determined before the chemical formula of the ionic compound can be written.

VALENCY

If the Valencies of the elements which take part in the compound are known, writing the chemical formula is simple.

Valency is the combining power of an atom or radical.
In **ionic** compounds it is the same as the **charge** on the ion.
In **covalent** compounds it is equal to the **number of bonds** formed.

+ METALS +		- NON-METALS -		
Element	Symbol of ion	Element	Symbol of ion	VALENCY
Lithium Sodium Potassium	Li^+ Na^+ K^+	Hydrogen Fluorine Chlorine Bromine Iodine	H^- F^- Cl^- Br^- I^-	1
Magnesium Calcium Barium	Mg^{2+} Ba^{2+} Ca^{2+}	Oxygen Sulphur	O^{2-} S^{2-}	2
Aluminium	Al^{3+}	Nitrogen Phosphorous	N^{3-} P^{3-}	3

List of Valency for common ions

for metal = the number of electron in the outermost shell
for non-metal = 8 – the number of electron in the outermost shell

→Some metals can form positive ions **with different charges**, depending on the compound that they are found in.

List of valency for common ions with variable charges

Element	Symbol of ion	Valency	Element	Symbol of ion	Valency
Copper(I)	C^{+}	1	Mercury(I)	Hg^{+}	1
Copper(II)	C^{2+}	2	Mercury(II)	Hg^{2+}	2
Iron(II)	Fe^{2+}	2	Lead(II)	Pb^{2+}	2
Iron(III)	Fe^{3+}	3	Lead(IV)	Pb^{4+}	4
Tin(II)	Sn^{2+}	2	Cobalt(II)	Co^{2+}	2
Tin(III)	Sn^{3+}	3	Cobalt(III)	Co^{3+}	3
Chromium(II)	Cr^{2+}	2	Nickel(II)	Ni^{2+}	2
Chromium(III)	Cr^{3+}	3	Nickel(IV)	Ni^{4+}	4
Manganese(II)	Mn^{2+}	2	Silver(I)	Ag^{+}	1
Manganese(IV)	Mn^{4+}	4	Zinc(II)	Zn^{2+}	2

Sometimes the charges on silver and zinc ions are not represented. Assume then that the silver ion is Ag^{+}, and the zinc ion is Zn^{2+}

RADICALS

Some negative ions exist in groups with an overall charge.

A **radical** is a group of atoms within a compound that maintains its identity throughout a chemical reaction.

List of valency for common ions with variable charges

Radical	Symbol of ion	Valency	Element	Symbol of ion	Valency
Hydroxide	OH^{-}	1	*Manganate(VII)*	MnO_4^{-}	1
Nitrate	NO_3^{-}		Ethanoate	CH_3COO^{-}	
Nitrite	NO_2^{-}		**Ammonium**	NH_4^{+}	
Hydrogen carbonate	HCO_3^{-}		**Carbonate**	CO_3^{2-}	2
Hydrogen sulphate	HSO_4^{-}		**Sulphate**	SO_4^{2-}	
Chlorate	ClO_3^{-}		*Sulphite*	SO_3^{2-}	
			Dichromate(VI)	$Cr_2O_7^{2}$	
			Phosphate	PO_4^{3-}	3

When writing chemical formulae involving radicals, **never take apart** with each element – take it as **a whole group**

How to Deduce Chemical Formulae from Valencies or Ions:

1. **Write the symbols** for the combining elements and radicals

Magnesium oxide → **Mg** (magnesium), **O** (oxygen)

2. **Write the valency** of each on the top right-hand side (leave off the charges for ions)

Mg^{2} O^{2}

3. Re-write the symbols, but **swap the valencies** and write at the bottom right-hand side

Mg^{2} O^{2}

Mg_2 O_2

4. Find the lowest ratio of the two numbers

2:2 → **1:1**

the chemical formula is **MgO**

Ignore writing subscript **1**

[EXAMPLE]

Aluminium sulphate

Al^{3} $(SO_4)^{2}$

(SWAP)

Al_2 $(SO_4)_3$

2:3 lowest ratio

ANS: $Al_2(SO_4)_3$

Calcium Carbonate

Ca^{2} $(CO_3)^{2}$

(SWAP)

Ca_2 $(CO_3)_2$

2:2 → 1:1 lowest ratio

ANS: $CaCO_3$

Make sure;
Put *BRACKETS* around any ***RADICALS*** *before swapping!!!*
Remove the brackets from the radical with subscript 1

VERY VERY IMPORTANT:
You must **LEARN** the **VALENCIES** of the COMMON ELEMENTS, RADICALS and IONS *if you wish to progress any further with your Chemistry!!!!*

How to Deduce the Valency or Ion from the Chemical Formulae:
(WORKING BACKWARDS)

- Write the chemical formula of the compound
  → Include all subscripts even 1
- Re-write the symbols for each element / radical, but swap the subscript numbers and write at the top right-hand side
- MULTIPLY BOTH NUMBERS by the same number to get the CORRECT VALENCY for any known element / radical
- Include the CHARGE of the ION if required
  → 1st ion is + ve and 2nd ion is - ve

[EXAMPLE]

What is the charge of the iron ion in $Fe_2(SO_4)_3$?

Fe_2 $(SO_4)_3$
(SWAP)
Fe^3 $(SO_4)^2$

ANS: The charge on the iron ion is +3

The charge on the ion is the same as the valency

COVALENT COMPOUNDS

Many exceptions exist to the rules for writing the chemical formulae of covalent compounds, making them difficult to remember. Some general rules:

- Many gases are made up of diatomic molecules, i.e. H_2, O_2, N_2 etc
- Group VII elements also exist as diatomic molecules, i.e. Cl_2, Br_2, F_2, etc.

You can use the rule for some covalent compounds like water, carbon dioxide, etc. You can try it.

Note that when **HCl** is in gaseous form, it is called **hydrogen chloride** gas; when it is dissolved in water, it forms a solution called **hydrochloric acid**.

Compound	Formula	Compound	Formula
Carbon monoxide	CO	Nitric acid	HNO_3
Carbon dioxide	CO_2	Sulphuric acid	H_2SO_4
Sulphur dioxide	SO_2	Hydrochloric acid	HCl
Sulphur trioxide	SO_3	Methane	CH_4
Ammonia	NH_3	Ozone	O_3
Hydrogen chloride	HCl	Ethanoic acid	CH_3COOH
Silicon dioxide	SiO_2	Ethanol	C_2H_5OH

List of common covalent compounds

CHEMICAL EQUATIONS

A Chemical equation is a way of summarizing a chemical reaction. Although it can be written in words, an equation is often written using chemical symbols and a chemical formula. When you write chemical equations, to start with, it is advisable to follow these steps:

Steps for making chemical equations
1. **Identify reactants and products** in reaction and **write down the equations in words** using either the information given or your own chemical knowledge. **Reactants** are the chemical elements or compounds that a chemical reaction starts with. **Products** are the chemical elements or compounds that are produced [EXAMPLE] Reactants: **sulphuric acid**, **sodium hydroxide** Products : **sodium sulphate**, **water** **sulphuric acid** + **sodium hydroxide** → **sodium sulphate** + **water**
2. **Put** every substance into the correct **chemical formula**. $\mathbf{H_2SO_4 + NaOH \rightarrow Na_2SO_4 + H_2O}$
3. **Balance the equation**. (Changing the proportions of reactants and products, in such a way that the number of atoms of each element is the same on both sides.) $H_2SO_4 + \mathbf{2}NaOH \rightarrow Na_2SO_4 + \mathbf{2}H_2O$
4. Finally, **put state symbols** in the equation for every reactant and product. $H_2SO_4\mathbf{(aq)} + 2NaOH\mathbf{(aq)} \rightarrow Na_2SO_4\ \mathbf{(aq)} + 2H_2O\mathbf{(l)}$

state symbols
(s) solid / precipitate
(l) liquid / molten
(g) gas / vapour
(aq) aqueous /dissolved in water solution / dilute

Most pupils get injured at the balancing stage! Let's look at it in details!!

STEPS FOR BALANCING THE EQUATION:

- There must be the **same number** of **atoms** on **both sides**, so that all atoms are accounted for and none are lost or gained.
- Only **balance** by putting a number **<u>IN-FRONT</u>** of the formulae where needed i.e. you <u>can not</u> change the chemical formulae of a given substance

1. Find the total number of atoms on each side of the equation

[EXAMPLE]

$Na \quad + \quad O_2 \quad \rightarrow \quad Na_2O$

	Left-side	**Right-side**
Na	1 x 1 = 1	1 x 2 = 2
O	1 x 2 = 2	1 x 1 = 1

2. Find an element that **doesn't balance** and **pencil** in a **number** (**IN-FRONT!!!**) to try and correct it

→ For **odd** and **even** in-balances, try **swapping** the numbers

$\mathbf{2}Na \quad + \quad O_2 \quad \rightarrow \quad Na_2O$

	Left-side	**Right-side**	
Na	**2** x 1 = 2	1 x 2 = 2	OK!
O	1 x 2 = 2	1 x 1 = 1	NO

3. See if this works, it may create another in-balance but pencil in another number (IN-FRONT) and see if this works

$2Na \quad + \quad O_2 \quad \rightarrow \quad \mathbf{2}Na_2O$

	Left-side	**Right-side**	
Na	2 x 1 = 2	**2** x 2 = **4**	NO
O	1 x 2 = 2	**2** x 1 = **2**	OK!

*Balance **just ONE type** of atom at a time!!!*

4. Continue chasing unbalanced elements and the equation will balance eventually

Remember, if the numbers are **not working**, rub them out and **try again**

≩↓

$\mathbf{4}Na \quad + \quad O_2 \quad \rightarrow \quad 2Na_2O$

	Left-side	**Right-side**	
Na	**4** x 1 = **4**	4 x 2 = 4	OK!
O	1 x 2 = 2	2 x 1 = 2	OK!
			→**Congratulation!!**

➢ Always **double-check** your answer that every element is balanced

[EXAMPLE] $\mathbf{3}H_2 \quad + \quad N_2 \quad \rightarrow \quad \mathbf{2}NH_3$

H:	2 x **3** = 6	3 x **2** = 6	(try swap)
N:	2	1 x **2** = 2	OK!

Beginners often find it difficult to balance chemical equations, especially the more complicated ones. A few rules to bear in mind

➢ If an equation cannot be balanced, it may be wrong.
Either the formulae of one or more of the substances involved is /are written wrongly or there may be missing/extra substances in the equation.

➢ **Never change** the chemical formula of compounds when balancing equations. You can only add numbers in front of the chemical formula. For example, $2NaOH$ and Na_2OH has different meanings. $2NaOH$ means you have 2 units of $NaOH$ (=2 Na, 2O and 2H), while Na_2OH means you have 2 Na, 1O and 1 H

The Second dangerous point is putting chemical formulae!!! Let's look at it!

GOING FROM A WORD EQUATION TO A BALANCED EQUATION

IMPORTANT POINTS TO REMEMBER:

- Most non-metals form **molecules**, e.g. H_2, O_2, Cl_2, F_2, N_2 (not C, S or Si)
- Metals do not form molecules!!! e.g. Al, Cu, Mg, Fe
- Better to memorize Formulae of **common compounds** [e.g.] H_2O, CO_2, ammonia NH_3, methane CH_4, hydrochloric acid HCl, sulphuric acid H_2SO_4, nitric acid HNO_3
- **Use the valency** to find the formulae of (unknown) compounds

[EXAMPLE]

magnesium + hydrochloric acid → magnesium chloride + hydrogen

1.Write the FORMULAE for each compound

$Mg + HCl \rightarrow MgCl_2 + H_2$

2.BALANCE the equation

$Mg + \mathbf{2}HCl \rightarrow MgCl_2 + H_2$

	Left-side	**Right-side**	
Mg	1 x 1 = 1	1 x 1 = 1	OK!
H	**2** x 1 = 2	2 x 1 = 2	OK!
Cl	**2** x 1 = 2	1 x 2 = 2	OK!

It's completed!

GOING FROM A "WORDY" QUESTION TO A BALANCED EQUATION:

- Identify ALL the REACTANTS and PRODUCTS from the question (underline each compound referred to in question)
- Write the word equation: *REACTANTS → PRODUCTS*
- Write the symbol equation (remember molecules, metals, common formulae and to use the valency)
- Balance the equation
- Put state symbols

[EXAMPLE]

Sodium chloride *solution is formed by titrating* aqueous sodium hydroxide *with* dilute hydrochloric acid. Water *is also formed.*

Write the balanced chemical equation for the reaction including state symbols.

WORKING OUT:

REACTANTS: sodium hydroxide, hydrochloric acid

PRODUCTS: sodium chloride, water

Every **reactant** on the **left-**hand side
Every **product** on the **right**-hand side

sodium hydroxide + hydrochloric acid → sodium chloride + water

$NaOH + HCl \rightarrow NaCl + H_2O$

	Left-side	**Right-side**	
Na	1 x 1 = 1	1 x 1 = 1	OK!
O	1 x 1 =	1 x 1 = 1	OK!
H	1 x1 + 1 x 1 =2	1 x 2 = 2	OK!
Cl	1 x 1 = 1	1 x 1 = 1	OK!

already balanced!

Note that the **'liquid'** state and the **'aqueous'** state is **not the same**. The 'liquid' state of a substance is pure. For a solid substance, the liquid state is obtained by heating the substance until it melts, while the 'aqueous' state of a substance is obtained by dissolving it in water.

ANS: $NaOH\,(aq) + HCl\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l)$

No score is given if you cannot give a balanced equation, even though the formulae of the compounds in your equation are correct. State symbols are not necessary in your balanced chemical equation unless the question requires it.

IONIC EQUATIONS

Ionic equations are used when a chemical reaction involves the coming together of ions in solution.

IONIC EQUATIONS show only the **changes taking place** in a chemical reaction

Formation of Ions from a chemical formula

- Usually, **only *(aq)*** compounds **split** to **form ions**
- First ion is **+ ve** (metal, H^+ or NH_4^+)

ve: a Valence Electron is an electron found in the outermost electron shell of an atom.

Second ion is – **ve** (non-metals and radicals)

- **Charge** on Ion = **Valency**
- Radicals are unchanged i.e. stay together
- SMALL number denoting number present *moves* to a BIG number **in-front**
- No. of + ve charges = No. of - ve charges

STEPS:

1 Make sure you **have a BALANCED chemical equation**

[EXAMPLE]

Hydrochloric acid + calcium carbonate → calcium chloride + carbon dioxide + water

$$2HCl\,(aq) + CaCO_3\,(aq) \rightarrow CaCl_2\,(aq) + CO_2\,(g) + H_2O\,(l)$$

2 **Split** only **soluble ionic compounds** [***(aq)*** compounds] **into its ions.**

➢ *Insoluble ionic compounds*, *elements* and *covalent compounds* **remain** unchanged.

$$\underline{2H^+}(aq) + \underline{2Cl^-}\,(aq) + \underline{Ca^{2+}}(aq) + \underline{CO_3^{2-}}\,(aq) \rightarrow \underline{Ca^{2+}}\,(aq) + \underline{2Cl^-}\,(aq) + CO_2\,(g) + H_2O\,(l)$$

3 **Cancel out spectator ions**

Spectator ions are the ions that appear in both the left and right side of the equation. electron

$$2H^+(aq) + \cancel{2Cl^-}(aq) + \cancel{Ca^{2+}}(aq) + CO_3^{2-}\,(aq) \rightarrow \cancel{Ca^{2+}}(aq) + \cancel{2Cl^-}\,(aq) + CO_2\,(g) + H_2O\,(l)$$

4 **Rewrite the equation** without spectator ions.

$$2H^+\,(aq) + CO_3^{2-}\,(aq) \rightarrow CO_2\,(g) + H_2O\,(l)$$

This is *IONIC EQUATION!!!*

[EXAMPLE]

Copper metal is displaced from its solution when an iron nail is placed into a solution of blue copper (II) sulphate. The clear solution which remains after the reaction is complete is iron (II) sulphate.
Derive the ionic equation for this reaction.

WORKING OUT:

REACTANTS: Copper (II) sulphate, iron
PRODUCTS : copper, iron (II) sulphate

Copper (II) sulphate + iron → copper + iron (II) sulphate

$\underline{CuSO_4\,(aq)} + Fe(s) \rightarrow Cu\,(s) + \underline{FeSO_4\,(aq)}$ *BALANCED*

$Cu^{2+} + \underline{SO_4^{2-}}\,(aq) + Fe(s) \rightarrow Cu\,(s) + Fe^{2+} + \underline{SO_4^{2-}}\,(aq)$ *SPLIT / CANCEL*

IONIC EQUATION: $Cu^{2+}(aq) + Fe(s) \rightarrow Cu\,(s) + Fe^{2+}(aq)$

Compounds that are sparingly soluble or very sparingly soluble can be considered as insoluble when writing ionic equations involving them.

When you construct ionic equations, the number of each particle and the total charge must be the same on both sides of the equation.

[EXAMPLE]

$$Cl_2 + 2Br^- \rightarrow Cl^- + Br_2$$

is not a balanced ionic equation since the total charge on the LHS is **2-** while the total charge on the RHS is only **1-**.

The balanced ionic equation will be

$$Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$$

where the total charge on both sides of the equation is **2-**

4.2 Stoichiometric calculations

RELATIVE MASSES

RELATIVE ATOMIC MASS

Relative atomic mass is the average mass of a large number of atoms of a particular element.

The symbol for relative atomic mass is **Ar**.

All naturally occurring elements are mixture of isotopes and therefore the relative atomic mass of an element takes into account the percentage of various isotopes that may be present.

A_r is simply the **average of the mass numbers for each of the isotopes present** in the element.

A_r = **Sum** for each Isotope { **% Present × Mass Number** }

[EXAMPLE]
Chlorine gas is 75% chlorine-35 atoms and 25% chlorine -37 atoms.

Ar = (75% X35) + (25% X 37) = **35.5**

You can find Ar for each element in Periodic Table.

The relative atomic mass is a ratio and therefore **has no unit**.

There is a clear **distinction** between mass number and relative atomic mass: **the mass number** of an atom is the number of protons and neutrons in the nucleus of the atom. It is ALWAYS a whole number. **The relative atomic mass** of an element is the average mass of its atoms compared to the mass of a Carbon-12 atom.

RELATIVE MOLECULAR MASS

Relative Molecular Mass is the **Sum** of the **A_r** for **all atoms present** in the molecule

The symbol for relative molecular mass is **Mr**.

[EXAMPLE]
Calculate the relative molecular mass of chloroform $CHCl_3$

Mr ($CHCl_3$) = Ar (C) + Ar (H) + Ar (Cl) = 12 + 1 + 3 × 35.5 = **119.5**

[EXAMPLE 2]
Calculate the relative molecular mass of copper (II) sulphate crystals, $CuSO_4\ 5H_2O$

Mr($CuSO_4\ 5H_2O$) = Ar (Cu) + Ar (S) + 9 × Ar (O) +10 × Ar (H)
= 64 + 32 + 9 × 16 + 10 × 1
=**250**

The relative molecular mass is a ratio and therefore **has no unit**.

PERCENTAGE MASS OF AN ELEMENT IN A COMPOUND

The percentage by mass of an element present in a compound is fixed.

This percentage can be calculated using the formula

Formula

$$\%\text{ Mass of element in compound} = \frac{\text{mass of element in compound}}{\text{Mr of compound}} \times 100$$

$$= \frac{\text{No. of atoms} \times \text{Ar of element}}{\text{Mr of comound}} \times 100$$

[Example] Calculate the percentage by mass of oxygen in carbon dioxide, CO_2

$$\%\ O \text{ in } CO_2 = \frac{2 \times 16}{12 + 2 \times 6} \times 100\% \quad = 72.7\%$$

If the calculations are correct, the total percentages of all the elements present in a compound should add up to 100%. Hence in example, % mass of carbon present is calculated simply as 100 – 72.7 = 27.3%

[EXAMPLE 2]
Calculate the percentage by mass of water in sodium carbonate crystals, $Na_2CO_3\ 10H_2O$

$$\%\ H_2O \text{ in } NaCO_3\ 10H_2O = \frac{10 \times (2 \times 1 + 16)}{2 \times 23 + 12 + 3 \times 16 + 10 \times (2 \times 1 + 16)} \times 100\%$$
$$= 62.9\%$$

➡ The percentage by mass can also be used to calculate the mass of an element in a given sample.

Formula

$$\text{Mass of element in sample} = \%\text{ of element in compound} \times \text{sample mass}$$
$$= \frac{\text{Mass of element in compound}}{\text{Mr of comound}} \times \text{sample mass}$$

[EXAMPLE]
Calculate the mass of copper in 32g of copper (II) sulphate.

$$\text{Mass of Cu} = \frac{\text{mass of Cu in } CuSO_4}{\text{Mr of } CuSO_4} \times 32g = \frac{64}{64 + 32 + 4 \times 16} \times 32g = 12.8g$$

5. Periodic Table

COTENT	LEARNING OBJECTIVE (Pupils should be able to)
Periodic Trends	• Describe the Periodic Table as a method of classifying elements and its use to predict properties of elements • Describe change form metallic to non-metallic character across a period • Describe the relationship between Group number, number of valence electrons and metallic or non-metallic character.
Group properties	• Describe lithium, sodium and potassium in Group I as a collection of relatively soft metals showing a trend in melting point and in reaction with water • Predict the properties of other elements in Group I, given data, where appropriate. • Describe chlorine, bromine and iodine in Group VII as a collection of diatomic non-metals showing a trend in colour, state, and in their displacement reactions with group I halide. • Predict the properties of other elements in Group VII, given data, where appropriate. • Describe the noble gases as being inert. • Describe the uses of the noble gases in providing an inert atmosphere

In the last page of this text book PERIODIC TABLE is put as an appendix

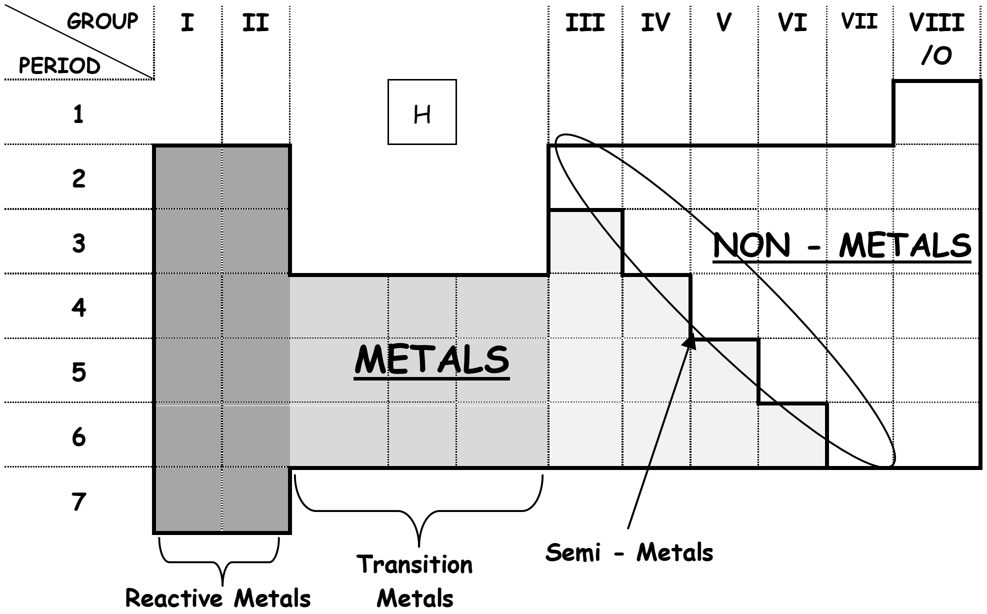
There are 103 elements discovered at present.
Periodic Table can help you classify them and easy to know properties of each element.

5.1 Periodic trends

In Periodic Table, Elements are arranged in order of **increasing proton number** (atomic number)

— A horizontal row in the Periodic Table is known as a **Period**.
There are 7 Periods in the Periodic Table.
— A vertical column in the Periodic Table is known as a **Group**
There are 8 groups in the Periodic Table.

The Periodic table are divided into sections as shown below



METALS AND NON-METALS

A 'zig- zag' diagonal line (staircase line) in the Periodic table divides metallic elements from non-metallic elements.

NON-METALS

Non-metals are elements which do not have the properties of metal and always form the **negative ions** when they react to form ionic compounds.

- Their states are often **gases** at room conditions (N, O, F, Cl, noble gases and etc) or **low melting point solids** (P, S, I and etc).
- They are **Poor** electrical and thermal conductors.

METALS

Metals are a class of chemical elements which always form **positive ions** when they react to form compounds.

- They are often shiny **solid**
- They are **good** conductors of heat and electricity.
- They also form solid oxides that act as bases.

TRANSITION METALS

Transition metals are found in the centre block of Periodic Table.

- They are **hard, strong** metals with **high melting and boiling points**.
- They also have **high density**.

They have partly filled inner electron shells which give them distinctive properties.

They also have the following properties:

- Form IONS in aqueous SOLUTION which are **COLOURED**
  Examples: Copper(II) → Blue
  Iron (II) → pale green
  Iron (III) → reddish brown (when solid)
  yellow (when in solution)
- Form **positively** charged ions with **variable charges**
  Examples: Copper forms either Cu^{+} or Cu^{2+}
  Iron forms either Fe^{2+} or Fe^{3+}

Some uses of transition metals

Many transition metals are often **good catalysts** in industry to speed up reactions.

[Examples] The hydrogenation of oil to make margarine uses a nickel catalyst.
The manufacture of ammonia uses an iron catalyst

Many transition metals are used to **make alloys**

[Example] Steel is made by mixing iron with a small amount of carbon
- an alloy is a mixture of two or more elements

Many transition metals are often useful **engineering materials** as strong and hard metals

SEMI-METALS

Elements near the line (such as Boron and silicon) are called 'semi-metals'. Semi–metals have the characteristics of **both metals and non-metals**.

They are often electrical semiconductors whose physical properties resemble metals but whose chemical properties resemble non-metals.

PERIOD

A period is a horizontal row or elements

- The first 3 rows are called 'short periods'.
- The next 4 rows, which include the transition metals, are called 'long periods'.

Elements in the **same period** have the **same number of electron shells.**

Going **across a period** from left to right, the number of outermost electrons **increases** by one every successive element

=> Elements change from metallic to non-metallic character across a period.

METALS **Reactive → Less reactive**	→	**Semi Metals**	→	NON-METALS **Less reactive → Reactive**

GROUP

A group is a vertical column of elements

Not the same **total** number of electrons

Elements in the same group have the same number of **outer shell electrons** (valence electrons). This means that elements in the same group will have similar chemical properties since they will form ions with the same charge.
They will also form **compounds with similar formulae**.

The group number is the same as the number of outer shell electrons
→ Valency of metals = Group number
Valency of non-metals = 8 – Group number

Going **down a group** form top to bottom, the number of **electron shells increases** by one for every consecutive element

=> Elements become more metallic in character, i.e. they lose valence electrons more easily.

METALS	**NON-METALS**
As you go down a group ↓ **More reactive**	**As you go up a group** ↑ **More reactive**

It becomes easier for an element to lose electrons going **down a group**. With an increase in the number of electron shells, **the attraction** between the positively charged nucleus and the valence electrons are **reduced**. The elements become **more metallic** in character.

The element **hydrogen is unique** because a H atom can form either H, by losing its one valence electron, or H, by gaining one valence electron to complete its outer shell. Forming ions with 1+ charge is characteristic of Group I elements, while forming ions with 1- charge is typical of Group VII elements. This explains why hydrogen is placed by itself in the Periodic Table.

SHEILDING

When we talk about the reactivity of elements, sometimes we use this idea.

Reactivity changes as you move down the Groups due to **shielding**.
This is because **each new e^- shell** is further out from the nucleus and the inner e^- shells **shield** the **outer e^-**'s from the **positive nucleus**.

As **METAL** atoms get **bigger**, the outer e^- is **more easily lost**. This makes METALS **MORE REACTIVE** as you go **DOWN** Groups I and II.

As **NON-METAL** atoms get **bigger**, the extra e^- are **harder to gain**. This makes NON-METALS **LESS REACTIVE** as you go **DOWN** Groups VI and VII.

5.2 Group properties

They all react with water to form *alkalis*, hence their name.

Group I: ALKALI METALS

Elements in Group I are also known as **alkali metals** that are the elements in the first group in the periodic table, which all have a single valence electron.

They are the **most reactive metals** group in the Periodic Table.

PHYSICAL PROPERTIES	CHEMICAL PROPERTIES
• Silvery / white in colour • SOFT and EASY TO CUT with a knife • LOW DENSITIES • relatively LOW MELTING POINTS • Good conductors of heat and electricity • All have 1 e^- in outer shell	• All are VERY REACTIVE • REACT vigorously with COLD WATER to FORM H_2 GAS (Hydrogen gas) • BURN in AIR with COLOURED FLAMES to form OXIDES • Alkali metals REACT with HALOGENS to PRODUCE a NEUTRAL SALT which dissolves to form a colourless solution

The compounds of Group I metals are **all** ionic. Group I metals always form ions **with 1+ charge** in their compounds.

- Metals in general are hard, dense with high melting and boiling points. Group I metals are highly unusual because they are soft, easily cut and have low density and low melting points.

Group I: Alkali Metals – BEHAVIOUR TRENDS

The first 3 elements in the group can float on water

ATOMIC NUMBER	ALKALI METALS		Density (g/cm^3)	Melting point(C°)
3	Li	Lithium	*0.53*	*180*
11	Na	Sodium	*0.97*	*98*
19	K	Potassium	*0.86*	*64*
37	Rb	Rubidium	*1.5*	*39*
55	Cs	Caesium	*1.9*	*29*
87	Fr	Francium	-	-

Some physical properties of G I elements

TRENDS:

-Reactivity increase
-Densities increase
-M.P. and B.P decrease.
-Softer to cut

Francium is the **most reactive** metal in the Periodic Table.

REACTION OF GROUP I ELEMENTS WITH WATER

Group I elements become **more reactive down the group**.

-How to store "Alakli Metal"-
These metals are stored under oil or in vacuum to prevent them from reacting with water and/ or oxygen in the air.

[Example] Reaction of Group I elements **with water**

Li reacts violently
Na reacts very violently, sometimes with an explosion
K reacts explosively

- Group I elements react with cold water to form **metal hydroxides** and **hydrogen gas.**

Group I metal + water → metal hydroxide + Hydrogen gas

$2Na(s) + 2H_2O\,(l) \rightarrow 2NaOH\,(aq) + H_2\,(g)$

The compounds of Group I elements are usually **colourless** unless the compound contain a transition element

The metal hydroxide solutions formed are all **strong alkalis** with pH values more than 7.

Group VII: HALOGENS

Elements in Group VII are also known as **halogens** that are the elements which have seven valence electrons in their outermost shell.
Their ions and compounds are called halides

The name is derived form *Greek* and means "salt-markers."

They are **very reactive non-metals**.

Each molecule is made up of **2 atoms** joined together by a single covalent bond.

PHYSICAL PROPERTIES	CHEMICAL PROPERTIES
• Non-metals which form DIATOMIC MOLECULES with COLOURED VAPOURS • POISONOUS (use fume cupboard) • Poor conductors of heat and electricity • All have one less e^- in outer shell → form 1- ions	Elements can either be ionic or covalent. • REACT with METALS to form IONIC compounds (neutral SALTS) • REACT with another NON-METALS to form MOLECULAR compounds

Group VII: Halogens – BEHAVIOUR TRENDS

Fluorine and chlorine are gases; **bromine** is a **liquid** while iodine is a solid.

ATOMIC NUMBER	HALOGENS		Melting point(C°)	Boiling point(C°)	State (at r.t.p)	Colour	**TRENDS:**
9	F_2	Fluorine	*-220*	*-188*	*Gas*	*Pale yellow*	*-Reactivity decreases*
17	Cl_2	Chlorine	*-101*	*-35*	*Gas*	*Yellow green*	*-M.P. and B.P increase.*
35	Br_2	Bromine	*-7*	*59*	*Liquid*	*Reddish brown*	*- State Gas to Solid*
53	I_2	Iodine	*114*	*184*	*Solid*	*black*	*- Colour*
85	At_2	Astatine	-	-	*Solid*	-	*darker*

Physical properties of G VII elements

Astatine is the **least reactive** element in Group VII.

The element Iodine has many colours, depending on what physical state it is in. It is **purple in gaseous** state, **black in solid** state, and forms a **reddish brown solution** when dissolved in water.

States are under **Room Conditions**

DISPLACEMENT REACTIONS

Group VII elements become **less reactive down the group**.

More reactive halogens will **DISPLACE less reactive** halogens from their aqueous salt solutions

[Example1]

$$Cl_2\,(g) + 2NaI\,(aq) \rightarrow 2NaCl\,(aq) + I_2\,(aq)$$

Chlorine, being the more reactive halogen, will displace iodine from aqueous sodium iodide. The reddish brown solution formed contains sodium chloride and aqueous iodine.

[Example2]

$$F_2\,(g) + 2NaCl\,(aq) \rightarrow 2NaF\,(aq) + Cl_2\,(aq)$$

Fluorine, being more reactive than chlorine, will displace it from aqueous sodium chloride. The yellowish solution formed contains sodium fluoride and chlorine.

[Example3]

$$Br_2\,(l) + 2KI\,(aq) \rightarrow 2KBr\,(aq) + I_2\,(aq)$$

Bromine, being more reactive than iodine, will displace it form aqueous potassium iodide. The reddish brown solution formed contains potassium bromide and iodine.

[Example4]

There is **no reaction** between iodine (I) and sodium chloride (NaCl). Iodine is less reactive than chlorine and does not displace chlorine from sodium chloride.

Displacement reactions are redox reactions.

USES OF HALOGENS

- Small amounts of **fluorine** is added to tap water and **toothpaste** to prevent tooth decay
- **Chlorine** is used to treat tap water and swimming pools **to kill harmful germs and bacteria**
- **Iodine** is used as an **antiseptic**; Small amounts of iodine are needed in our bodies to prevent goitre (swelling of thyroid gland)
- **Silver halides** are used on **black and white photographic film**

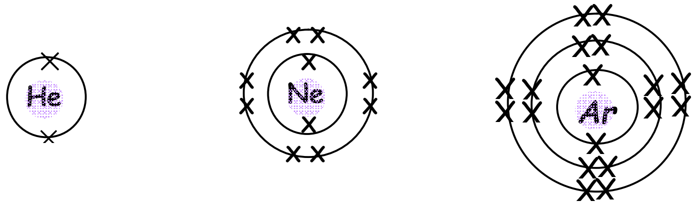
  $2AgBr$ + LIGHT → Br_2 + $2Ag$ (silver metal deposit)

Group VIII/O: NOBLE GASES

Group VIII elements are also known as **noble gases** or **inert gases** that are extremely inert.

ELECTRONIC STRUCTURE OF GROUP VIII ELEMENTS

Group VIII elements are **the least reactive elements** in the Periodic Table because all their outer shells are completely filled.



Helium, $^{4}_{2}He$ **Neon, $^{20}_{10}Ne$** **Argon, $^{40}_{18}Ar$**

The full electronic structure of the first 3 noble gases

NOBLE GASES all have the following PROPERTIES:

All are **colourless**
All are **gases** that consist of single atoms.
All are **monatomic** gases with **Low Melting and Boiling points**.
All are stable, **hardly combine** with other atoms (INERT)

Group O: Noble Gases – BEHAVIOUR TRENDS

ATOMIC NUMBER	NOBLE GASES		Density (g/cm^3)	Boiling point(C°)
2	He	Helium	*0.14*	*-269*
10	Ne	Neon	*0.67*	*-246*
18	Ar	Argon	*1.38*	*-186*
36	Kr	Krypton	*2.89*	*-157*
54	Xe	Xenon	*4.56*	*-108*
86	Rn	Radon	*7.70*	*-62*

Some physical properties of G VIII elements

TRENDS:

-Reactivity inert
-Densities increase
- B.P increase.

Group O: Noble Gases – USES

Noble gases occupy 1% of the atmosphere. Of all the noble gases, Argon is the most abundant in air.

- **Helium** is used in **balloons and airships** because it is less dense than air (the second lightest gas and not flammable like hydrogen)
- **Neon** is used in **advertising signs** because it glows red when electricity is discharged through it.
- **Argon** is used to fill **filament lamps** (light bulbs). It prevents the filament inside the bulb from burning out.
- **Krypton** and **Xenon** are used in **lamps in lighthouses**, **stroboscopic lamps**, and **photographic flash units.**

→ *All these uses are because noble gases are **CHEMICALLY INERT!!***

The terms '**unreactive' and 'inert' must be distinguished** from each other.
A substance can be unreactive, but given the correct condition, it will undergo reaction to form new substance. If a substance is said to be inert, then it is stable and will not take part in reaction no matter what conditions are provided.

6. CHEMICAL REACTIONS

COTENT	LEARNING OBJECTIVE (Pupils should be able to)
Rate of reactions	• Describe the effect of concentration, pressure, particle size and temperature on the speeds of reactions and explain these effects in terms of collisions between reacting particles. • Interpret data obtained from experiments concerned with speed of reaction.
Redox reaction	• Define oxidation and reduction (redox) in terms of oxygen gain/loss
Energy changes	• Define exothermic and endothermic reactions • Describe bond breaking as an endothermic process and bond forming as an exothermic process

6.1 Rate of Reaction

Rate of Reaction is the Speed at which the chemical reaction proceeds.
The speed of a chemical reaction refers to how fast reactants are used up or how fast products are formed in a reaction.

- Different chemical reactions have different speeds.

EXAMPLES

- Reaction of potassium metal with water -> very fast
- Resting of an iron nail in the presence of air and water -> slow, takes a few days.
- Gold reacting with oxygen in the air -> no reaction, speed of reaction of gold with oxygen is zero.

MEASURING THE RATE OF REACTION

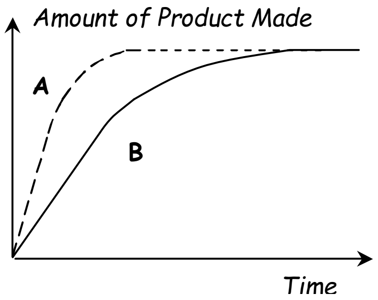
The speed of a reaction is defined as

$$\text{speed of reaction} = \frac{\text{change in amount of reactant or product}}{\text{time}}$$

We can measure the speed of reaction **by observing** either how quickly the **reactants are used up** or how quickly the **products are forming**.
Common methods are shown below.

1. **Change in mass** (usually gas given off)
   - Any reaction that produces a gas can be carried out on a MASS BALANCE and the mass disappearing is easily measured
2. **Volume of gas** produced
   - Uses a GAS SYRINGE to measure the volume of gas produced
3. Precipitation
   - Observe a MARKER through a solution which becomes CLOUDY as the product precipitates and measures the time taken for the marker to DISAPPEAR

ANALYSING GRAPHS FOR RATE OF REACTION



- Graph is **steepest** at **START**
  → **REACTION** is **FASTEST** at start
- Graph then starts to **level out**
  → **REACTIONS SLOWS** as more REACTANTS get USED UP
- When graph is **flat** (level)
  → **REACTION** has **FINISHED**
- Reaction ***A*** is **faster** than Reaction **B**

The greater the gradient is, the faster the reaction is.

STEPS: How to Plot Graphs

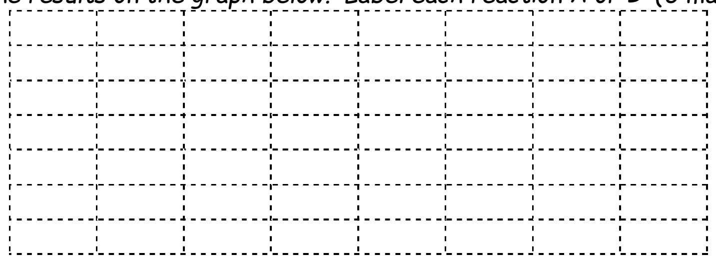
- Use the **data range** to choose a suitable **SCALE** for each axis.
  → Note that the scale markings should be **evenly spaced**!
- **PLOT** each **data point** carefully
- Draw a **LINE OF BEST FIT**
- **LABEL** each axis (include **units**)
- *Give the graph a* **TITLE** e.g. Graph of Loss of Mass with Time

[EXAMPLE]

Calcium carbonate when heated undergoes thermal decomposition to form calcium oxide and carbon dioxide. The loss of mass during the reaction was measured for two different reactions.

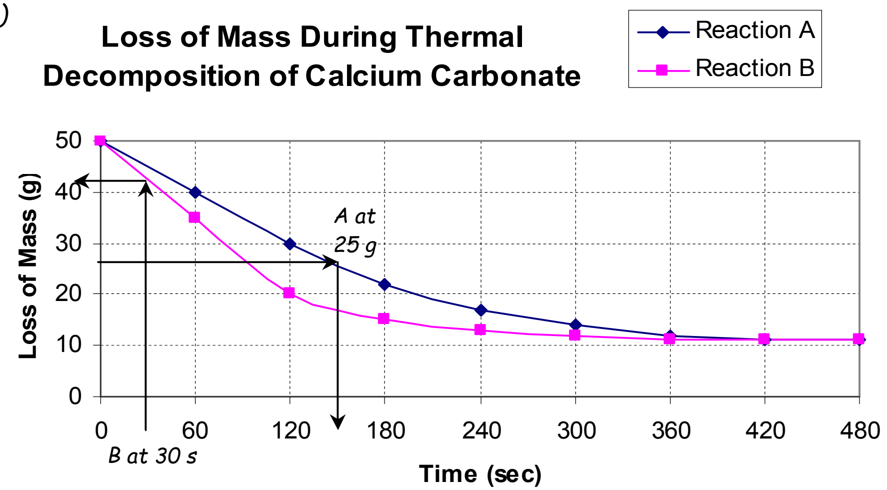
Time (sec)	0	60	120	180	240	300	360	420	480
A Loss in Mass (g)	50	40	30	22	17	14	12	11	11
B Loss in Mass (g)	50	35	20	15	13	12	11	11	11

(a) Plot the results on the graph below. Label each reaction A or B (3 marks)



(b) Which of the two reactions was the fastest? Suggest the reason for this difference. (2 marks)
(c)(i) What was the mass lost for Reaction B after 30 seconds? (2 mark)
(ii) After what time did Reaction A lose 25 g of mass? (1 mark)
(d). Did Reaction A start with less calcium carbonate, more calcium carbonate or the same amount of calcium carbonate as Reaction B? (1 mark)

ANSWER:
(a)



(b) Reaction B was faster (as it has the steeper curve). The reason for this is that Reaction B was heated more strongly than Reaction A.
(c) (i) Mass Lost = Original Mass – Mass at 30 s
= 50 g – 42 g = 8 g { 2 marks!! use graph & calculation}
(ii) 150 sec { use graph }
(d) Reaction A and Reaction B started with the same amount of calcium carbonate because they both ended up with exactly the same amount of mass lost during the reaction.

The reaction is complete once the gradient of the curve becomes zero. In the above example, the reaction is completed in 420 seconds.

It is incorrect to say that since the reaction is completed in 420 seconds, the reaction is half completed at 420/2 = 210 seconds. This is because the rate of reaction changes with time – it is faster at the beginning, becomes slower as the reaction proceeds and finally stops. To determine the time when the reaction is half completed, we need to look at how long it takes for half the amount of reactant to be used.

FACTORS AFFECTING THE SPEED OF A REACTION

A reaction is caused by the **collision** of particles of substances.
The factors below affect the collision and therefore affect speed of reaction.

1. **Concentration** of reactant – more concentrated reactants, faster reactions
2. **Pressure** of reactant (gaseous reactions only) – higher pressure, faster reactions
3. **Temperature** – higher temperature, faster reactions
4. **Particle size** of reactant – smaller particle reactants, faster reactions
5. Use of a **Catalyst**

I.e. INCREASE in SURFACE AREA

CATALYSTS

A **CATALYST** is a substance which INCREASES the speed of reaction without being changed or used up in the reaction

- Catalysts **work best** with a **big surface area** [*e.g.*] *powder, pellets or gauze*
- Catalysts are specific to certain reactions
- Enzymes are biological catalysts
- Catalysts are used to REDUCE COSTS in Industrial Reactions

→ Catalysts **LOWER** the **ACTIVATION ENERGY**

Activation energy is the minimum amount of energy required to start a chemical reaction. This energy is used in the breaking of chemical bonds

EXAMPLES

- *IRON CATALYST is used to produce AMMONIA in the HABER PROCESS.*
- *A PLATINUM CATALYST is used in the production of nitric acid and in the CATALYTIC CONVERTER in a car engine.*
- *The catalytic converter is used in the engine of the car to promote combustion of the fuel to reduce pollution from unburnt exhaust gases.*

6.2 Redox Reactions

Redox reactions are reactions that involve both oxidation and reduction

Oxidation is a chemical reaction involving the gain of oxygen.
Reduction is a chemical reaction involving the loss of oxygen.

Oxidation and reduction reactions occur simultaneously.
If one reactant is oxidised, then the other reactant must be reduced.

LOSS or GAIN OF OXYGEN

When a substance gains oxygen during a chemical reaction, it is oxidised. If it loses oxygen the substance is reduced.

EXAMPLES

$$H_2(g) + CuO(s) \rightarrow Cu(s) + H_2O(g)$$

H_2 is oxidised to H_2O because it has gained oxygen.
CuO is said to be reduced to Cu because it has lost oxygen.

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

CO is oxidised to CO_2 because it has gained oxygen.
Fe_2O_3 is reduced to Fe because it has lost oxygen.

The definition using oxygen is the easiest to use. However, the use is limited to reactions involving oxygen atoms.

Actually there are other definitions for redox reactions which use gain or loss of hydrogen/electron, or oxidation number.

The most versatile definition by far is the one using oxidation numbers although we do not talk about those definitions here.

DEFINING OXIDISING and REDUCING AGENTS

Oxidising agents are substances that help oxidation take place. In the process, they become reduced. Similarly, reducing agents are substances that help reduction take place. In the process, they become oxidised.

EXAMPLE $CuS + 4H_2O_2 \rightarrow CuSO_4 + 4H_2O$

CuS is oxidised to $CuSO_4$ as it has gained oxygen, and H_2O_2 is reduced to H_2O because it has lost oxygen.

Since H_2O_2 causes CuS to become oxidised (by losing oxygen to it), it is the oxidising agent. On the other hand, CuS is the reducing agent since it causes H_2O_2 to become reduced (by removing oxygen from it)

6.3 Energy changes

When chemical reactions take place, energy is either taken in or given out from the surroundings in the form of heat and/or light. We describe reactions as either exothermic or endothermic, depending on whether energy is absorbed or given out.

EXOTHERMIC REACTIONS

An exothermic reaction is a chemical reaction during which heat is given out, causing a temperature rise in the surroundings.

EXAMPLES

✧ When sodium carbonate is dissolved in a beaker of water, the temperature o the solution rises from 28C° to 40C°

✧ When methane is burnt, heat energy is evolved and the temperature of the surroundings rises.

✧ When acids react with alkalis, neutralisation takes place wit the evolution of heat. The temperature of the solution formed rises.

After all the bonds in the reactants are broken, the atoms will form new bonds to give the products of the reaction; Heat energy will be released when these new bonds are formed.
Hence, **bond forming is exothermic**.

ENDOTHERMIC REACTIONS

An endothermic reaction is a chemical reaction during which heat is taken in, causing a temperature drop in the surroundings.

EXAMPLES

✧ When ammonium chloride is dissolved in a beaker of water, the temperature of the solution drops from 28C° to 22C°

✧ Heat energy must be supplied during the thermal decomposition of calcium carbonate.

✧ Light energy must be absorbed before photosynthesis by plants can take place.

In a chemical reaction, bonds between reactants must be broken so that the atoms can rearrange themselves to form products. Heat energy must be taken in by the reactants for bond breaking.
Hence, **bond breaking must be endothermic.**

EXOTHERMIC REACTIONS	ENDOTHERMIC REACTIONS
• **HEAT** is **GIVEN OUT**	• **HEAT** is **TAKEN IN**
• **Temperature RISES**	• **Temperature DECREASES** (or reaction requires heating)
• **Energy released** when forming bonds is **GREATER** than the **energy absorbed** when breaking bonds	• **Energy absorbed** when breaking bonds is **GREATER** than the **energy released** when forming bonds
[e.g.] Combustion, Freezing and condensing, Neutralisation reactions, Haber process, Reduction of iron (III) in the blast furnace, Adding concentrated H_2SO_4 to water, Adding water to anhydrous $CuSO_4$	*[e.g.] Decomposition of limestone, Decomposition of halide crystals by light, Melting and boiling, Photosynthesis, Dissolving certain salts (KCl or NH_4NO_3)*

Summary of exothermic and endothermic

7. ACIDS, BASES AND SALTS

COTENT	LEARNING OBJECTIVE (Pupils should be able to)
Acid, Base and Alkali	• Describe the meaning of the terms acid and alkali in terms of the ions they contain or produce in aqueous solution. • Describe the characteristic properties of acids as in their reactions • Describe the characteristic properties of bases as in their reactions with acids and with ammonium salts and their effects on indicator paper. • Describe neutrality and relative acidity and alkalinity • Describe the formation of hydrogen and product of the reaction between; reactive metals and water/metals and acids • Classify oxides as either acidic, basic, or amphoteric related to metallic or non-metallic character. • Describe and explain the importance of controlling acidity in soil
Preparation of Salt	• Describe the preparation, separation and purification of salts. • Suggest a methods of preparing a given salt from suitable starting materials, given appropriate information
Identification test	• Describe the use of aqueous sodium hydroxide and aqueous ammonia to identify the aqueous cations. • Describe tests to identify the anions. • Describe tests to identify the gases. • Describe the identification of hydrogen using a lighted splint (water being formed) • Describe the identification of oxygen using a glowing splint. • Describe the identification of carbon dioxide using lime water.

For human beings air and water are two of the commonest, indeed, the most important chemical substances in the world. There are however, other classes of chemical materials which are not only common but are also very important in our everyday lives. These classes are the acids, bases and salts which are the subject matter of this chapter.

Acids are chemical compounds which produce hydrated hydrogen ions H^+ (aq) when in aqueous solution.

Bases are chemical compounds that react with acids to form a salt and water

Alkalies are water-soluble bases which produce hydrated hydroxide ions OH^- (aq) when in aqueous solution.

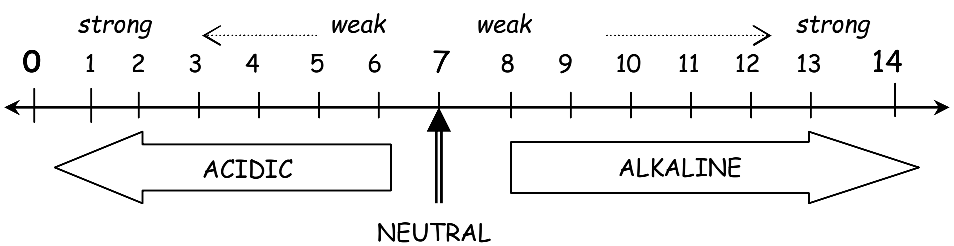
Salts are chemical compounds formed when the hydrogen of an acid is partially or wholly replaced by a metal or other positive ion (E.g. ammonium ion).

7.1 Acid, Base and Alkali

PH SCALE

The pH scale shows the strength of an acid or alkali in an aqueous solution.
It is a measure of the concentration of H^+ ions present in the solution.

pH is an abbreviation for "*potential of hydrogen*"



The pH scale ranges from 0 to 14

- A pH value of less than 7 indicates that the solution is acidic.
- A pH value of more than 7 indicates that the solution is alkaline.
- A pH value of 7 indicates that the solution is neutral.
  It is neither acidic nor alkaline.
  EXAMPLES Pure water, saltwater, and various organic liquids

INDICATOR

We can check whether a solution is acidic or alkaline by indicators.
An acid-base indicator changes colour, reversibly. Some indicators with change in colour are shown below.

INDICATOR	Colour in:	
	ACID	ALKALI
Litmus paper	Red	Blue
Phenol phthalein	Colourless	Pink or red
Methyl orange	Red	Yellow
Bromothymol blue	Yellow	Blue

UNIVERSAL INDICATOR

Universal indicator is a mixture of several indicators and turns a range of colours corresponding to different pH values.

← *Acidic* **Neutral** *Alkaline* →

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
	red			orange		yellow		Green	Green-blue		blue			violet	

Colour change in Universal indicator with pH

The range of colours in different solutions of pH for the universal indicator approximates the rainbow colours – red, orange, yellow, green, blue, indigo and violet. Taking pH 7(neutral) to be green, colours to the left of green the rainbow indicate acidic solution, while colours to the right indicate alkaline solutions.

ACIDS

Acids are substance that will dissolve in water and undergo ionization to form hydrogen ions. The table below shows some common acids found in the laboratory and the ions they contain.

Name of acid	Ions present	Salt formed
Hydrochloric acid, **HCl**	H^+ , Cl^-	- chloride, -Cl
Sulphuric acid, **H_2SO_4**	H^+, SO_4^{2-}	-sulphate, $-SO_4$
Nitric acid, **HNO_3**	H^+, NO_3^-	- nitrate, $-NO_3$
Ethanoic acid, **CH_3COOH**	H^+, $CH_3\,COO^-$	- ethanoate, $-CH_3\,COO$

Common Acids

Acids have acidic properties only when they are dissolved in water.

Note that HCl in gaseous form is called **hydrogen chloride**. If it is dissolved in water , it will undergo ionisation to form a solution called **hydrochloric acid**.

BASICITY of an acid = NUMBER OF H^+ IONS produced when aqueous
[e.g.] Basicity of H_2SO_4 = 2 H^+ ions = 2

TYPES OF ACIDS

1. MINERAL acids
   -These are STRONG acids and they IONISE COMPLETELY
   [e.g.] *hydrochloric*, *sulphuric* and *nitric* acids

2. ORGANIC acids
   -These are WEAK acids and they only PARTIALLY IONISE
   [e.g.] *carbonic*, *acetic* (*vinegar*) and *citric* acids

PHYSICAL PROPERTIES

- Acids taste **sour.**
  [e.g.] vinegar and lemon (They contain ethanoic acid and citric acid respectively)
- Acids turn **litmus paper red**
- Acids have **pH values less than 7**

REACTION OF ACIDS

There are 3 common reactions of acid

1. Acid + **base** → salt + $\mathbf{H_2O}$
2. Acid + **metal** → salt + $\mathbf{H_2}$
3. Acid + **carbonate** → salt + $\mathbf{H_2O}$ + $\mathbf{CO_2}$

REACTION OF ACID WITH BASE

Acids react with bases to give **salts** and **water** only

This reaction is also called **neutralisation**

[EXAMPLE 1] Acid with Metal oxide

Hydrochloric acid + zinc oxide → zinc chloride + water

$2HCl\ (aq) + ZnO\ (s) \rightarrow ZnCl_2\ (aq) + H_2O\ (aq)$

[EXAMPLE 2] Acid with Metal Hydroxide

Hydrochloric acid + sodium hydroxide → sodium chloride + water

$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (aq)$

REACTION OF ACID WITH METAL

Acid react with metal above hydrogen in the reactivity series to give **salts** and **hydrogen gas**.

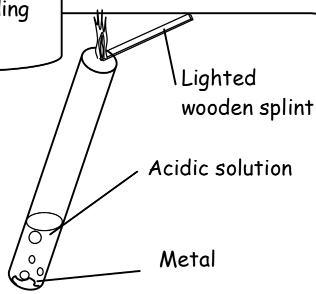
[EXAMPLE]

Hydrochloric acid + magnesium → magnesium chloride + hydrogen gas.

$2HCl\ (aq) + Mg\ (s) \rightarrow MgCl_2\ (aq) + H_2\ (g)$

The metal reactivity series lists metals according to their reactivity. SEE the topic 'METAL'.

We can test for hydrogen gas using a lighted splint. The flame extinguishes with a 'pop' sound.



REACTION OF ACIDS WITH CARBONATES

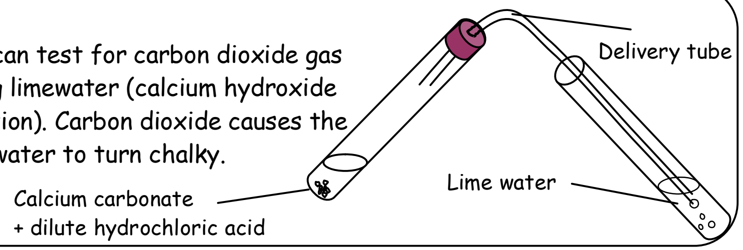
Acids react with carbonates to give **salts**, **carbon dioxide gas** and **water**.

[EXAMPLE]

calcium carbonate + hydrochloric acid → calcium chloride + carbon dioxide + water

$CaCO_3\ (s) + 2HCl\ (aq) \rightarrow CaCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$

We can test for carbon dioxide gas using limewater (calcium hydroxide solution). Carbon dioxide causes the limewater to turn chalky.



BASES AND ALKALIS

A base is a substance that reacts with an acid to **give a salt and water only**. This reaction is called neutralisation.

Neutralisation is the chemical reaction between a base and an acid to form a salt and water

Acid	+	Base	→	Salt	+	Water
HCl (s)	+	NaOH (aq)	→	NaCl (aq)	+	H_2O (l)

SOLUBLE BASES AND INSOLUBLE BASES

- Many bases are insoluble in water.
  Bases that **can dissolve in water** form solutions called **alkalis**
- Bases are usually **metal oxides** or **metal hydroxides**.

Table below lists some common bases and alkalis.

Insoluble bases		Soluble bases	
Name	Formula	Name	Formula
Magnesium oxide	MgO	Sodium oxide	Na_2O
Copper(II) oxide	CuO	Calcium hydroxide	$Ca(OH)_2$
Lead(II) hydroxide	$Pb(OH)_2$	Ammonium hydroxide	NH_4OH

Common bases and alkalis

SIMPLE CLASSIFICATION ON BASES

Bases

Oxides | Hydroxides

Soluble | Insoluble | Soluble | Insoluble

→ **Alkalis** | **Basic oxides** | **Alkalis** | **Basic hydroxide**

It can be seen from this that all alkalis are bases not all bases are alkalis.

Solubility depends on the combination of ions involved. We will see that later.

→ Hydroxide ions, OH^- are produced when bases dissolve in water to form alkalis.

[EXAMPLE] Sodium hydroxide, NaOH

$NaOH\ (aq) \rightarrow Na^+ (aq) + OH^- (aq)$

The ability of alkalis to neutralise acids is due to the presence of these hydroxide ions.

PHYSICAL PROPERTIES OF ALKALIS

- Alkalis feel **slippery**
- Edible alkalis have **a bitter taste**.
- Alkalis turn **litmus paper blue**
- Alkalis have **pH values greater than 7**

REACTION OF BASES

There are 2 common reactions of bases

1. Base + **acid** → salt + $\mathbf{H_2O}$
2. Base(alkali) + **ammonium salt** → salt + **ammonium gas** + $\mathbf{H_2O}$

REACTION OF BASES WITH AMMONIUM SALTS

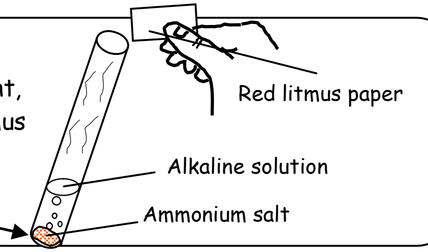
Alkalis react with ammonium salts to produce **salts**, **ammonia gas** and **water**.

[EXAMPLE]

sodium hydroxide + ammonium chloride → sodium chloride + ammonia + water

$NaOH\ (aq) + NH_4Cl\ (aq) \rightarrow NaCl\ (aq) + NH_3\ (g) + H_2O\ (l)$

The ammonia gas evolved is pungent, colourless and turns damp red litmus paper blue



SIGNIFICANCE OF pH MESUREMENTS

Apart form enabling us to determine whether substances are acidic or alkaline, pH values have very important significance and implications in industry, agriculture, pharmacy and medicine.

CONTROL OF pH IN AGRICULTURE

Most plants need a soil pH of 6.5 to 7.5 to grow well. If the ground is too acid, slaked lime (solid calcium hydroxide) can be added to neutralise the acid. This process is called liming the soil.

→ Slaked lime is chosen because

1. It is cheap and easily available.
2. Slaked lime is sparingly soluble in water. Once the aid is neutralised, the excess base will remain as solid in the soil. It will not dissolve in water to make the soil too alkaline.

Aqueous ammonia and aqueous sodium hydroxide are alkalis that can also neutralise; however, slaked lime has an advantage over them.
→ The person spraying the solution (e.g. sodium hydroxide solution) will not know when enough alkali has been added to neutralise the acid if the products of neutralisation appears as a colourless solution. Excess alkali will cause the ground to become alkaline.

OXIDES

Oxides are formed **when substances burn in oxygen gas**. Oxides have acidic, basic, amphoteric or neutral character, depending on which type of oxide they belong to.

Type of Oxide		Examples
1.	**ACIDIC (Non-metallic)**	Carbon dioxide CO_2, Sulphur dioxide SO_2, Nitrogen dioxide NO_2
2.	**BASIC (Metallic)**	Magnesium oxide **MgO**, Calcium oxide **CaO**, Sodium oxide Na_2O
3.	**AMPHOTERIC**	Zinc oxide **ZnO**, Aluminium oxide Al_2O_3, Lead(II) oxide **PbO**, Tin oxide and
4.	**NEUTRAL**	Water H_2O, Carbon monoxide **CO**, Nitrogen monoxide **NO**

PROPERTIES OF DIFFERENT TYPES OF OXIDES

ACIDIC OXIDES

Acidic oxides are usually **oxides of non-metals**.
- They form acids (*H^+ ions*) when dissolved in water

[EXAMPLE] Carbon dioxide + water → carbonic acid
$CO_2 + H_2O \rightarrow H_2CO_3$

Natural rain has a pH slightly lower 7. Carbon dioxide in the air will dissolve n rainwater to produce a weakly acidic solution of carbonic acid.

BASIC OXIDES

Basic oxides are **oxides of metals**.
-They react with acids to produce salt and water only.

[EXAMPLE]
Copper (II) oxide + sulphuric acid → copper (II) sulphate + water
$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$ (aq)

Here, neutralisation takes place.
Basic oxides that dissolve in water form solutions called alkalis

AMPHOTERIC OXIDES

Some oxides of metals known as amphoteric oxides **behave as acidic or basic oxides.**
- When they react with acids, they behave as basic oxides;
- When they react with alkalis, they behave as acidic oxides.

[EXAMPLE] Zinc oxide reacts with an acid and a base for neutralisation.

1. Sulphuric acid + zinc oxide → zinc sulphate + water
$\underset{\text{[acid]}}{H_2SO_4} + \underset{\text{[base]}}{\underline{ZnO}} \rightarrow ZnSO_4 + H_2O$
- In this case, zinc oxide is acting as a base

2. Zinc oxide + sodium hydroxide → sodium zincate + water
$\underset{\text{[acid]}}{\underline{ZnO}} + \underset{\text{[base]}}{2NaOH} \rightarrow Na_2ZnO_2 + H_2O$
- In this case zinc oxide is acting as an acid.

NEUTRAL OXIDES

Neutral oxides do not dissolve in water to form acids nor do they react with bases to form salts. ***NEITHER acidic nor basic propertie***

[EXAMPLES] Water, Carbon monoxide, Nitrogen monoxide, Sulphur monoxide, etc

7.2 Salt Preparation

As you have seen, the reaction of an acid results in the products of a salt.

Now we shall see how to prepare the salt required.

Salts are chemical compounds formed when the hydrogen of an acid is partially or wholly replaced by a metal or other positive ion (E.g. ammonium ion).

Compounds in which the H^+**ions** in an **acid** have been **replaced** by **ammonium ions**, $NH_4^+$are called ammonium salts.

SELECTION OF METHOD

SOLUBILITY RULES

The method chosen to prepare a salt depends on its solubility.

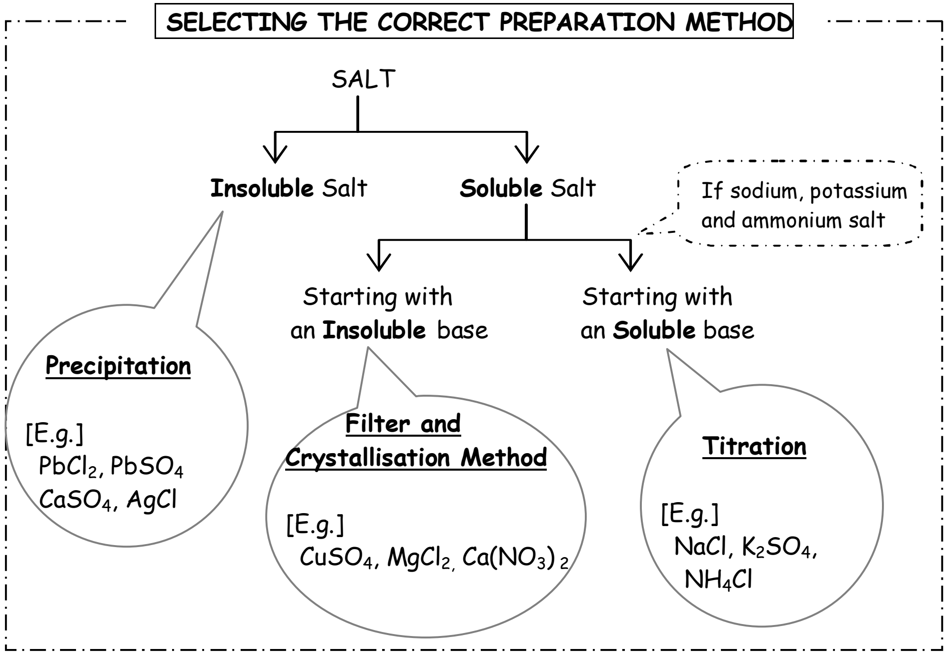
→ The solubility depends on the combination of positive and negative ions.

SOLUBLE	INSOLUBLE
• ALL **nitrates** • ALL **chlorides** *EXCEPT FOR: silver chloride and lead (II) chloride* • ALL **sulphates** *EXCEPT FOR: calcium sulphate, barium sulphate and lead (II) sulphate* Note that All **sodium**, **potassium** (even other Group I elements) **Ammonium** and **Nitrate** compounds are soluble.	• ALL **carbonates** *EXCEPT FOR: sodium carbonate, potassium carbonate and ammonium carbonate* • ALL **sulphides** *EXCEPT FOR: sodium sulphide, potassium sulphide and ammonium sulphide* • ALL **oxides** *EXCEPT FOR: sodium oxide, potassium oxide (Group I) and ammonium oxide* ○ALL **hydroxides** *EXCEPT FOR: sodium hydroxide, potassium hydroxide (Group I), ammonium hydroxide and calcium hydroxide*

Table: Solubility of compounds

➡ The selection of salt preparation method is summarised below.

- **Precipitation** is carried out if on insoluble salt is required.
- If a soluble salt is needed, it is prepared by **Filter and crystallisation method** or by **Titration**.



[EXAMPLES] *Name the correct method to prepare*

(a) *potassium chloride*

KCl → soluble salt →the base can contain K → soluble base

→ *use TITRATION*

(b) *zinc sulphate*

$ZnSO_4$ →soluble salt→ the base can not contain Na, K or NH_4 →insoluble base

→ *use FILTER AND CRYSTALLISATION METHOD*

(c) *silver chloride*

AgCl → insoluble salt → *use PRECIPITATION*

PREPARETION OF INSOLUBLE SALT

Insoluble salts are prepared by **mixing solutions** containing their **positive** and **negative ions** using the method of PRECIPITATION.

-The reactant solutions are chosen so that on **exchanging ions** the **unwanted** product is **still soluble** but the given **insoluble salt** will form as a **precipitate**.

SOLUBLE SALT + SOLUBLE SALT → SOLUBLE SALT + INSOLUBLE SALT (precipitate)

Precipitate is an insoluble solid formed when a chemical reaction occurs between two dissolved ionic substances

To find the Reactant Solutions:

Choose 2 starting solutions.

- One must contain the **positive ion** of the insoluble salt required
- The other must contain h the **negative ion** of the insoluble salt required.

[EXAMPLE] To make barium sulphate (*$BaSO_4$*),

REACTANT IONS:		*AFTER MIXING:*
Ba^{2+} NO_3^-	→	*$NaNO_3$ (soluble)*
Na^+ ***SO_4^{2-}***		*$BaSO_4$*

We can use barium nitrate and sodium sulphate.

The easiest salt solution containing the **positive ions** is the **METAL NITRATE** (as all nitrates are soluble).

For a solution containing the **negative ions**, can use the **SODIUM SALT** (as all sodium salts are soluble)

[EXAMPLE 2] To make silver chloride (*AgCl*),

REACTANT IONS:		*AFTER MIXING:*
Ag^+ NO_3^-	→	*$NaNO_3$ (soluble)*
Na^+ ***Cl^-***		*AgCl*

We can use silver nitrate and sodium chloride

PROCUDURE OF PRECIPITATION

- **Dissolve** each reactant separately in water
- **Mix** chemically equivalent quantities of the **reactant solutions**
- **Filter** the solution and **wash** the **precipitate** in warm distilled water
- **Dry** the solid salt that was produced in an **oven** (105°C)

Refer to the topic of 'separation technique'

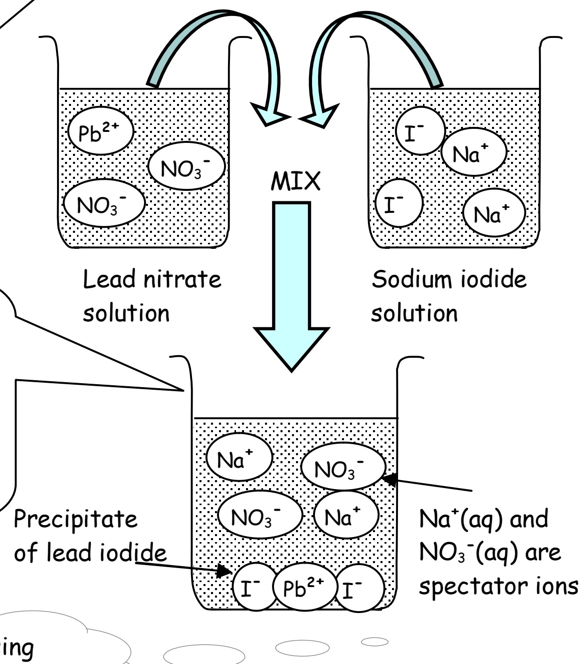
[E.g.] To prepare Lead(II) iodide, PbI_2

$$Pb(NO_3)_2(aq) + 2NaI(aq) \rightarrow PbI_2(s) + 2NaNO_3(aq)$$

The ionic equation is

$$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$$

The reactants involved in a precipitation reaction must be in solution form because the ions must be able to move and interact with one another when the reactants are mixed together



When the ions in the insoluble salt encounter each other, they will attract each other to form a solid that will sink to the bottom of the container and be collected as the **precipitate.**

Other salts prepared using precipitations include silver chloride, lead chloride and etc.

PREPARETION OF SOLUBLE SALT

As you have seen, we have two methods to prepare soluble salts depending on the solubility of the base that would be a starting material.

SOLUBLE SALTS are prepared using two methods:

1. Filter and crystallisation method
   Neutralising an ACID with EXCESS INSOLUBLE REACTANT
2. Titration
   Neutralising an ACID with the EXACT AMOUNT of ALKALI

FILTER AND CRYSTALLIDATION METHOD

This method is used for preparation of soluble salts when a suitable insoluble starting material can be found.

→The acid reacts with an EXCESS of insoluble reactant that can be:

1. METAL
2. BASE (INSOLUBLE)
3. CARBONATE

Therefore to prepare a given salt, we need to **choose** the correct **acid** and a **suitable insoluble reactant** (METAL, OXIDE, HYDROXIDE or CARBONATE).

STEPS

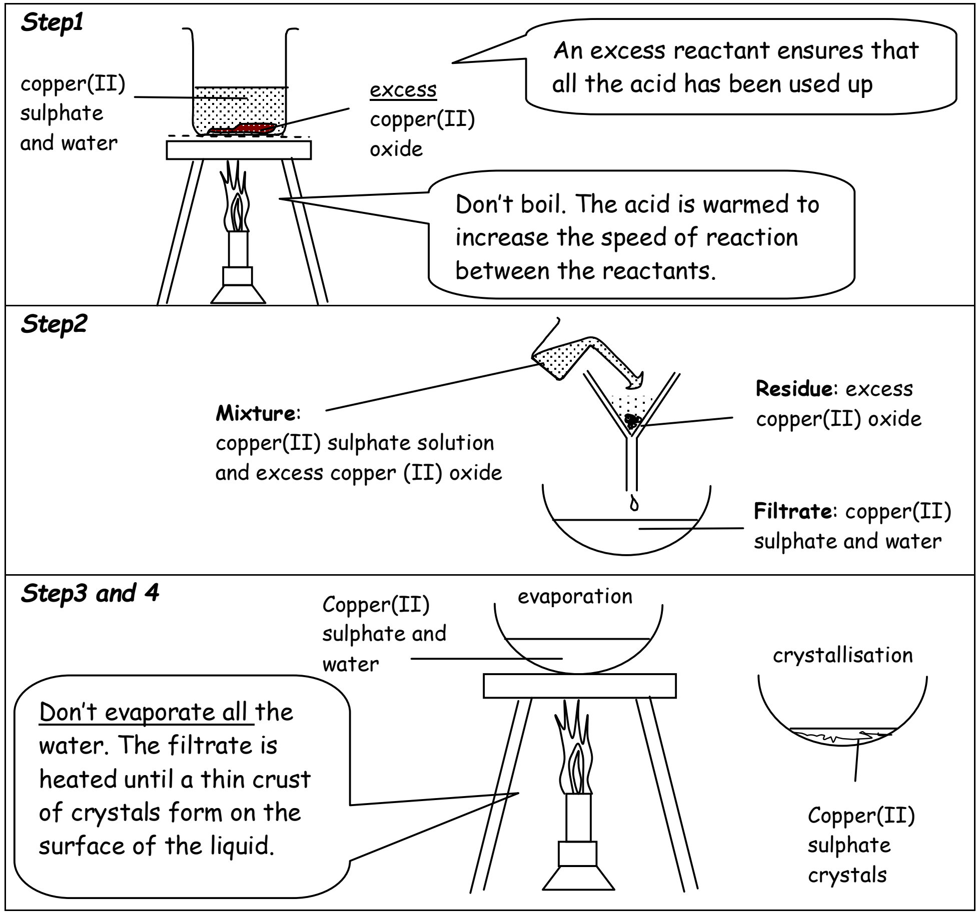
1. **Neutralise** the **acid** with and **excess** of the **insoluble reactant**
2. **Filter** off any **unreacted** reagent
3. **Evaporate** the solution to the **crystallisation point**
4. **Cool** to produce crystals of the salt
5. Filter, wash and dry the crystals before collection.

[EXAMPLE] The preparation of copper (II) sulphate

Starting materials; copper (II) oxide and dilute sulphuric acid.

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

This reaction is used below to illustrate the procedure.



Preparing a soluble salt by filter and crystallisation method.

If a metal **carbonate** is used to prepare a salt using this method, there will be bubbles of **carbon dioxide gas** as the metal carbonate is added to the acid in step1. When there is no more bubble, all the acid has been used up and we may proceed to next step.

TITRATION

The **soluble** salts of **ammonium** and **Group I** metals (sodium, potassium and lithium) are prepared using the TITRATION METHOD.

This is because all their compounds are soluble (including the metals themselves) and **very reactive**. The Group I metals are so reactive resulting in too violent reaction that we CAN NOT USE EXCESS reactant.

This method is used when it is not possible to find a suitable insoluble starting material like a metal, a metal oxide or a carbonate that can be easily filtered off at the end of the reaction.

→ **TITRATION** means using the **EXACT** quantities of reactants for the reaction.

INDICATOR

In a titration, an indicator is needed to show the endpoint of one reactant needed to exactly neutralise a given volume of the other reactant.
A common indicator used in the laboratory is the screened methyl orange.

Acidic solution	**END POINT** (neutral)	**Alkaline** solution
RED	GREY or COLOURLESS	GREEN

Colour change in methyl orange

STEPS OF TITRATION

To prepare a given salt, the most common procedure is to react the **alkali solution** with the **dilute acid** using a burette. Indicator is used to determine when the exact amount of reactant has been added.

[EXAMPLE]

The preparation of sodium nitrate is used to illustrate the procedure.

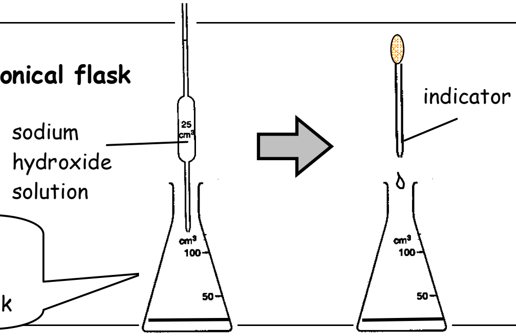
Starting materials: aqueous sodium hydroxide and dilute nitric acid

$$NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

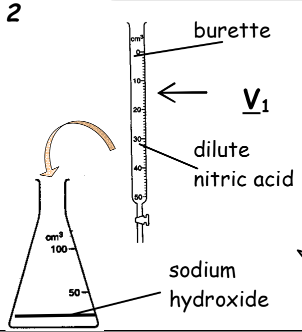
Step 1

- Place the soluble **dilute acid** in a **conical flask**
- Add a few drops of **indicator** (*e.g. methyl orange*)

Using pipette, measure 25.0cm of sodium hydroxide into conical flask



Step 2



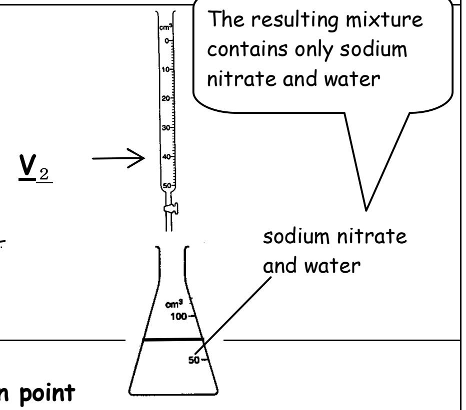
- From a **burette**, slowly add the **alkali solution**. Ensure that the solution is **mixed well**.
- When the indicator begins to change colour, the reaction should be slowed to a drip.

At this point, just enough acid is added to neutralise the alkali, all the alkali has reacted.

Step 3

- Once the **colour change** is **complete**, the reaction is complete (**END POINT**). The burette should be turned off.

From the titration result, we can know the exact volume of nitric acid needed to react with 25.0cm of sodium hydroxide. Volume of nitric acid, $V_a = V_2 - V_1$



Step 4

- **Evaporate** the solution to **crystallisation point**
- **Cool** to produce crystals of the salt
- Filter, wash and dry the crystals

-In a strict titration, second titration should be carried out. The salt solution obtained in the first titration is thrown away because it is affected by the indicator. Second one is done without it. The exact volume of acid to be added is obtained from the first.

7.3 Identification Tests

IDENTIFY SALT SOLUTIONS

To identify any salt solutions, we can take the following steps;

1. IDENTIFY / TEST for the METAL cation present
2. IDENTIFY / TEST for the SALT anion present

→ SALT SOLUTION = {METAL} + {SALT}
Test1 *Test2*

Cations are **positively** charged ions
Anions are **negatively** charged ions

TEST 1 :Identification of METAL CATIONS

When testing for a cation using either aqueous **sodium hydroxide** or **aqueous ammonia**, two observations will help identify the **cation** present:

1. the **colour** of the precipitate formed on adding a few drops of chemical regent;
2. the **solubility** of the precipitate **in excess** chemical regent.

- Table below summarises the test for cations.

Name of METAL	Cation present	Effect of NaOH solution		Effect of NH_4OH solution	
		Colour of Precipitate	IN EXCESS	Colour of Precipitate	IN EXCESS
Calcium	Ca^{2+}	white	insoluble	white	insoluble
Magnesium	Mg^{2+}	white	insoluble	white	insoluble
Iron (II)	Fe^{2+}	green	insoluble	green	insoluble
Iron (III)	Fe^{3+}	brown	insoluble	brown	insoluble
Copper(II)	Cu^{2+}	blue	insoluble	blue	dark blue sol^n
Zinc	Zn^{2+}	white	colourless sol^n	white	colourless sol^n
Lead (II)	Pb^{2+}	white	colourless sol^n	white	insoluble
Aluminium	Al^{3+}	white	colourless sol^n	white	insoluble

Table: Test for Cations

'sol' means **'SOLUTION'**

→ The cations react with the hydroxide ions present in aqueous sodium hydroxide or aqueous ammonia **to form insoluble hydroxides**. These insoluble hydroxides appear as precipitates.

[EXAMPLE]

$$\underset{}{Fe^{2+}(aq)} + \underset{\text{From NaOH or } NH_3}{\underline{2OH^-(aq)}} \rightarrow \underset{\text{Green precipitate}}{\underline{Fe(OH)_2}}$$

-Some of these precipitate dissolve in **excess** aqueous **sodium hydroxide** to form soluble complex salts.

These appear as colourless solution. This occurs for **amphoteric** metal hydroxides (*Al^{3+}, Zn^{2+} and Pb^{2+}*) which react with the alkalis. Again in an **excess** of **ammonium solution**, **Zn** and **Cu** redissolve to form soluble complex salts. These appear as colourless solution or dark blue solution.

FLOW CHART

From the previous table

- copper (II), iron(II) and iron(III) ions are easily identified by the characteristic colour of their precipitations.
- Aluminium, lead (II) and zinc ions all give the same observations when aqueous sodium hydroxide is used. However, only zinc ions will give a white precipitate soluble in excess aqueous ammonia; aluminium and lead ions do not.

✓ Solution containing **Al^{3+} ,Pb^{2+}**or **Zn^{2+}**
↓ - Add a few drops of aqueous ammonia and shake
✓ White precipitate formed
↓ - Add excess aqueous ammonia
✓In excess aqueous ammonia, precipitate is

INSOLUBE - **Al^{3+}or Pb^{2+}** present   SOLUBLE - **Zn^{2+}**combined

- To distinguish between aluminium and lead(II) ions, dilute hydrochloric acid or aqueous potassium iodide can be used:

Al^{3+} (aq)	+	$3Cl^-$(aq)	→	$AlCl_3$
		From hydrochloric acid		Colourless solution
Pb^{2+} (aq)	+	$2Cl^-$(aq)	→	$PbCl_2$
		From hydrochloric acid		White precipitate

-Similar results will be obtained if aqueous potassium iodide is used. Aluminium ions will give a colourless solution of aluminium iodide while lead(II) ions will give a yellow precipitate of lead(II) iodide.

TEST 2: Identification of SALT ANIONS

- the table below summarises the tests for anions.

ANION PRESENT	Formula	TEST and Result
Carbonate *also* *Hydrogen Carbonate*	CO_3^{2-} HCO_3^-	• **Add** hydrochloric **ACID** ⇨ **Carbon dioxide** is produced ⇨ Turns **limewater milky**
Chloride	Cl^-	• Acidify by adding dilute *nitric acid* • Add **silver** nitrate solution ⇨ **White precipitate** forms (*AgCl*)
Sulphate	SO_4^{2-}	• Acidify by adding dilute *hydrochloric acid* • Add **barium** chloride solution ⇨ **White precipitate** forms ($BaSO_4$)
Iodine	I^-	• Acidify by adding dilute *nitric acid* • Add **lead(II)** nitrate solution ⇨ **Yellow precipitate** forms (PbI_2)

Table Tests of Anions

When recording the observations after conducting tests for the carbonate and the nitrate ion, remember to include the smell and colour of the gas, the chemical test result for the gas as well as the name of the gas. Simply copying from the data sheet provided as 'carbon dioxide produced' or 'ammonia produced' is insufficient and will lead to a loss of marks.

[EXAMPLE]

Which solution will form a brown precipitate if sodium hydroxide is added and a white precipitate if silver nitrate is added?

Test 1: NaOH → brown precipitate → Fe^{3+} → iron (III) [cation]
Test 2: $AgNO_3$ → white precipitate → Cl^- → chloride [anion]

Salt solution = Iron (III) chloride

IDENTIFICATION OF GASES

◆ Carbon dioxide, sulphur dioxide and chlorine are all acidic gases and will turn moist blue litmus paper red. Hence, the blue litmus paper test is not a conclusive test; it only indicates the presence of an acidic gas. It is necessary to conduct confirmatory tests in order to conclude the presence of a particular gas.

◆ Ammonia, chlorine and sulphur dioxide have characteristic smell and are thus easily identified.

When testing for hydrogen gas, hold the lighted splint at the mouth of the test tube.

When testing for oxygen gas, insert the glowing splint into the test tube

Table below summarises the test for gases.

GAS	FORMULA	TEST and RESULT
Hydrogen	H_2	• Burns with a **'POP'** sound
Oxygen	O_2	• **Relights** a **glowing** splint
Carbon Dioxide	CO_2	• Turns **limewater milky**
Chlorine	Cl_2	• Turns damp blue litmus red then **bleaches** litmus paper → *Yellowish-green colour* → *Choking smell*
Ammonia	NH_3	• Turns damp **red litmus BLUE** → *Pungent smell*
Hydrogen Chloride	*HCl*	• Turns damp **blue litmus RED** → *Choking smell*
Sulphur Dioxide	SO_2	• Turns damp **blue litmus RED** → *Choking smell*

Table: Test for gases

FLAME TESTS

- React a small quantity of compound with a couple of drops of concentrated hydrochloric acid
- Dip a clean piece of platinum wire into the mixture and put into the flame of a Bunsen burner
- The flame colour in each reaction is shown below.

METAL PRESENT	CATION	FLAME COLOUR
Sodium	Na^+	Orange-yellow
Potassium	K^+	Lilac-pink
Calcium	Ca^{2+}	Brick red
Barium	Ba^{2+}	Pale green
Copper (II)	Cu^{2+}	Green
Lead (II)	Pb^{2+}	Blue

Table: Test for flames

REACTION SCHEMES

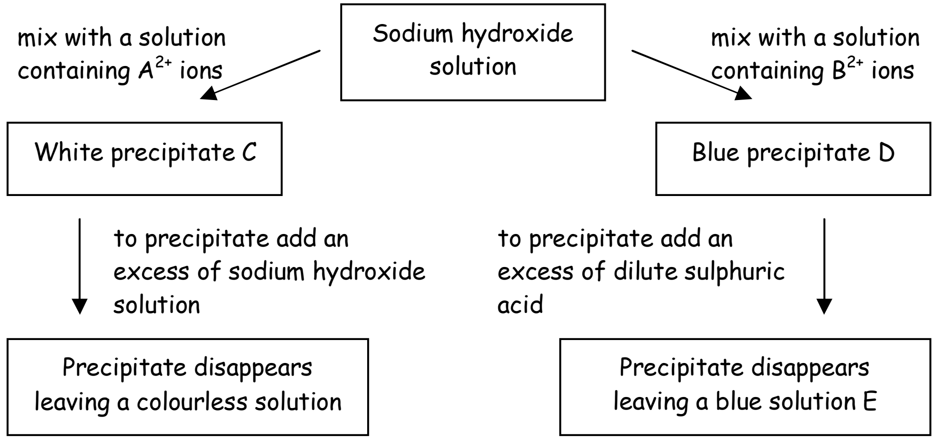
When solving Reaction Schemes it is essential that you know all the reactions typical for acids and alkalis and ammonium salts.

1. acid + **metal** → salt + $\mathbf{H_2}$
2. acid + **alkali / base** → salt + $\mathbf{H_2O}$
3. acid + **carbonate** → salt + $\mathbf{H_2O}$ + $\mathbf{CO_2}$
4. **ammonium solution** + acid → **ammonium salt** + H_2O
5. alkali + **ammonium salt** → salt + $\mathbf{NH_3}$ + $\mathbf{H_2O}$

You may also need to know the identification tests for the various cations, anions and gases.

[EXAMPLE]

The diagram below shows some properties and reactions of the ions, A^{2+} and B^{2+}, and the substances C, D and E. *{ 1997 Paper 3 Section B }*



(a) Suggest identities for the ions, A^{2+} and B^{2+}, and substances C, D and E

ANSWER:

A^{2+}: add NaOH → white precipitate. → colourless sol^n in excess → 2+ charge → zinc ion, Zn^{2+} (or Pb^{2+} ion but NOT Al^{3+} ion!!)

B^{2+}: add NaOH → blue precipitate. → copper ion, Cu^{2+}

C: $Zn(OH)_2$ (or $Pb(OH)_2$) zinc hydroxide (or lead (ii) hydroxide}

D: $Cu(OH)_2$ copper (ii) hydroxide

E: $CuSO_4$ copper (ii) sulphate

*(b) Write a chemical equation for any **one** of the reactions shown in the diagram*

ANSWER:

$2NaOH + Zn^{2+} \rightarrow Zn(OH)_2 + 2Na^+$ or

$Zn(OH)_2 + 2NaOH \rightarrow Na_2Zn(OH)_4$ or

$2NaOH + Cu^{2+} \rightarrow Cu(OH)_2 + 2Na^+$ or

$Cu(OH)_2 + H_2SO_4 \rightarrow CuSO_4 + 2H_2O$ { easiest answer!! }

(c) Explain how the sodium hydroxide solution can be used to distinguish between a solution containing an iron (II) compound and a solution containing an iron (III) compound.

ANSWER:

When sodium hydroxide solution is added to a solution containing an iron (II) compound, a green precipitate forms. When sodium hydroxide solution is added to a solution containing an iron (III) compound, a red-brown precipitate forms.

8. METALS

COTENT	LEARNING OBJECTIVE (Pupils should be able to)
Properties of metals	• Describe the general physical properties of metals • Explain why metals are often used in the form of alloys. • Identify representations of metals and alloys from diagrams of structures
Reactivity series	• Place in order of reactivity calcium, copper, hydrogen , iron, magnesium, potassium, sodium and zinc by reference to the reactions if any of the metals with water(or steam) • Account for the apparent unreactivity of aluminium in terms of the presence of an oxide layer which adheres to the metal. • Deduce an order of reactivity from a given set of experimental results.
Extraction of metals	• Describe the ease in obtaining metals from their ores by relation the elements of the reactivity series. • Describe the effect of aluminium on human beings.
Iron	• Describe the essential reactions in the extraction of iron from haematite Describe methods of rust prevention • Describe the idea of changing the properties of iron by the controlled use of additives to form alloys called steels • State the uses of mild steel and stainless steel.
Copper	• Describe the extraction and purification of copper from its ore. • State the uses of copper related to its properties
Aluminium	• State the uses of aluminium in air craft and food containers • State the uses of zinc for galvanising and for making brass (with copper)

In Periodic table we saw the most of the elements are metals. The non-metals confined to the top right-hand corner of the periodic table. Of over 100 elements which we know, only 21 are non-metals.
Now we shall investigate metals.

8.1 Properties of metal

PHYSICAL PROPERTIES and USE

In terms of appearance, Non-metals are different with each other while metals are all alike. Only copper and gold are coloured; all are shiny.
Let's see physical properties of typical metals comparing with those of non-metals.

METALS are…	NON-METALS are…
• Conduct electricity and heat	• Do not conduct electricity and heat
• Shiny	• Dull
• Malleable and ductile	• Brittle
• High MP and BP → usually solid	• Low MP and BP → usually liquid or gas
• High densities	• Low densities
• Magnetic materials	• Non-magnetic materials
• Strong and Tough	

Except for **Cu** and **Al**

Except for **Mercury** that is liquid at r.t.p. and **Gallium** that melts below 30℃

Malleable means 'be easily made **into sheet** without breaking.
Ductile means 'be easily made **into wire** without breaking.

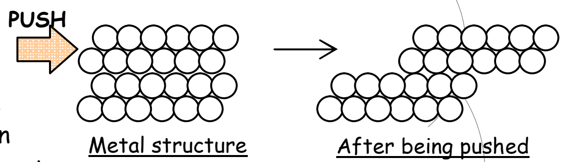
→ Due to the properties above, metals are used in many ways

METAL	USES	USEFUL PROPERTY
Copper	• Electric cables	• Excellent conductor of electricity
Aluminium	• Soft drink cans	• Does not corrode
Tin	• Coat 'tin' cans used for food e.g. jam	• Non-poisonous
Gold / Silver	• Jewellery	• Malleable and very unreactive

ALLOYS

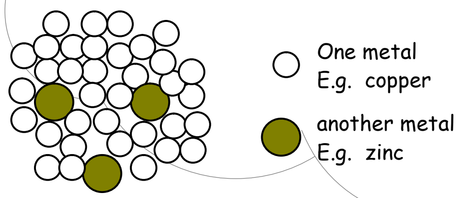
Pure metals are usually too soft and weak for most uses. To improve the strength and hardness of pure metal, we do the following treatment.

In pure metals, the atoms are arranged orderly in layers. When a force is applied to the metal, the layers of metal atoms can slide one over another



→ To improve the strength and hardness of pure metals, atoms of another element can be added, usually in small amounts.

These atoms prevent the atoms of the metal from sliding over on another, making the metal stronger and harder and less likely to have its shape distorted. The final product is an **alloy** of the metal



An **alloy** is a mixture of two or more elements that are usually metals except for carbon in steel.

ADVANTAGES OF ALLOYING

1. Stronger and harder than the pure metals.
2. Improved metal appearance
3. Increased resistance to corrosion

- Some examples of alloys are below

ALLOY	MIXTURE OF	USES	USEFUL PROPERTY
Mild steel	Iron and Carbon	• car bodies	• hard and strong
Stainless steel	Iron, Chromium and Nickel	• cutlery • surgical instrument	• corrosion resistant
Brass	Copper and Zinc	• Screw	• corrosion resistant
Bronze	Copper and Tin	• ornament	• good appearance
Duralumin	Aluminium and Magnesium	• aircraft and bicycle frames	• strong and lightweight
Solder	Lead and Tin	• welding metals	• low MP
Pewter	Tin and Lead	• ornament	• good appearance

8.2 Reactivity Series

The reactivity series is a list of metals placed in order of their reactivity, as determined by their reaction with water and dilute acid.

REACTION OF METALS WITH WATER

Some metals react with cold water while others react only with steam.

- Table 1 lists the reaction of some metals with water

Table1 ; reaction of metals with water

Metal	Observation/Equation
Potassium (K)	Reacts **very violently**. Enough heat is produced to ignite the hydrogen gas produced. The hydrogen burns with a blue flame. $2K\,(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$
Sodium (Na)	Reacts **violently**. The hydrogen gas produced may catch fire. $2Na\,(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
Calcium (Ca)	Reacts **readily**. Hydrogen gas and calcium hydroxide solution are formed. $Ca\,(s) + 2H_2O(l) \rightarrow Ca(OH)_2\,(aq) + H_2(g)$
Magnesium (Mg)	Reacts **very slowly** with cold water. A test tube of hydrogen gas is produced only after a few days. $Mg\,(s) + 2H_2O(l) \rightarrow Mg(OH)_2\,(aq) + H_2(g)$
Zinc (Zn), Iron (Fe), Lead(Pb), Copper(Cu), Silver(Ag)	**Do not react** with cold water

Table2 ; reaction of metals with steam

Metal	Observation/Equation
Magnesium (K)	The hot magnesium **reacts violently** with steam to form magnesium oxide (a white powder) and hydrogen gas. A bright white glow is produced during the reaction. $Mg(s) + H_2O(l) \rightarrow MgO(s) + H_2(g)$
Zinc (Zn)	Hot zinc **reacts** with steam to produce zinc oxide and hydrogen gas. Zinc oxide is yellow when hot and white when cold. $Zn(s) + H_2O(l) \rightarrow ZnO(s) + H_2(g)$
Iron (Fe)	Red hot iron **reacts slowly** with steam to form hydrogen gas and tri-iron tetraoxide.
Lead(Pb), Copper(Cu), Silver(Ag)	**Do not react** with steam

→ FROM THE OBSERVATIONS of the reactions of metals with water

1. When metals react with water or steam, metal hydroxides or metal oxide and hydrogen gas are formed.

Metal + water → Metal hydroxide + Hydrogen gas
Metal + steam → Metal oxide + Hydrogen gas

Note that magnesium reacts with both water and steam.
When it reacts with water, the product is magnesium hydroxide;
when it reacts with steam. The product is magnesium oxide.

2. The more vigorous the reaction, the more reactive the metal.

◆Potassium, sodium, calcium are reactive metals.
◆Magnesium, zinc and iron are fairly reactive metals.
◆Lead, copper and silver are unreactive metals.

REACTION OF METALS WITH DILUTE HYDROCHLORIC ACID

The reaction of metals with dilute acid is also taken into account for the reactivity series (The dilute acid is hydrochloric acid here).

Table3 ; reaction of metals with dilute hydrochloric acid

Metal	Observation/Equation
Potassium(K), Sodium(Na)	**Explosive** reaction. Reaction is not usually carried out because it is too dangerous to do in a laboratory.
Calcium (Ca)	Reacts **vigorously** to give hydrogen gas and calcium chloride. $Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
Magnesium (Mg)	Reacts **rapidly** to give hydrogen gas and magnesium chloride. $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
Zinc (Zn)	Reacts **moderately fast** to give hydrogen gas and zinc chloride. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
Iron (Fe)	Reacts **slowly** to give hydrogen gas and iron(II) chloride. $Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
Copper(Cu), Silver(Ag)	**Do not react** with cold water

If a piece of aluminium foil is reacted with hot dilute hydrochloric acid, the initial rate of reaction will be very slow as the acid reacts with the layer of aluminium oxide on the surface of the foil. Once the oxide layer is removed, the reaction will speed up as **aluminium is a reactive metal**.

➡ We can draw the reactivity series of metals from the reactions of metals with water and dilute acid as shown below.

THE METAL REACTIVITY SERIES

VERY REACTIVE		Potassium	K
		Sodium	Na
		Calcium	Ca
FAIRLY REACTIVE		Magnesium	Mg
		Aluminium	Al
NOT VERY REACTIVE		Zinc	Zn
		Iron	Fe
		Tin	Sn
		Lead	Pb
		HYDROGEN	H_2
NOT AT ALL REACTIVE		Copper	Cu
		Silver	Ag
		Platinum	Pt
		Gold	Au

The more reactive metal has the higher tendency to lose valence electrons and form positive ions.

Hydrogen is included in the series although it is a non-metal. It serves as **a reference point** in the series

- Metals above hydrogen will react with dilute acids to give hydrogen gas.
- Metals below hydrogen will not react with dilute acids.

Important Note: **Aluminium** is placed higher in the reactivity series although it shows no observable reaction with dilute hydrochloric acid. **It appears less reactive** due to the **protective layer** of aluminium oxide (Al_2O_3) that keeps the metal inside.

The metal reactivity series may differ from book to book, depending on how many metals are included in it.

Generally

Group I metals will be located at the top of the series, since they are the most reactive metals in the Periodic Table.

Group II metals, **Group III** metals and finally, the **transition metals** will follow them.

DISPLACEMENT REACTIONS

Displacement reactions from solutions can be predicted using the reactivity series.

More reactive metals will **displace** a **less reactive** metal from its compound or solution (a **colour change** of the solution is often **observed**)

→A metal higher in the series will displace a metal lower in the series.

[EXAMPLE 1] Iron + copper (II) nitrate solution

$$Fe\ (s) + Cu(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$$

A brown metallic deposit of coper metal will form as the solution turns from blue to pale green due to the formation of iron (II) ions.

[EXAMPLE 2] Iron + zinc (II) sulphate solution

Iron is lower than zinc in the reactivity series. Since it is less reactive than zinc, no displacement reaction will take place.

[EXAMPLE 3] Copper + silver nitrate solution

Since copper is above silver in the reactivity series, copper will displace silver from silver nitrate solution.

$$Cu\ (s) + 2AgNO_3\ (aq) \rightarrow Cu(NO_3)_2\ (aq) + 2Ag(s)$$

A layer of silver will form on the copper metal. The solution will also turn from colourless to blue due to the formation of copper(II) ions.

Displacement reactions also take place for **Group VII** elements. The more reactive halogen will displace the less reactive halogen from a solution containing its ions.

8.3 Extraction of metals

The method of extraction of a metal from its compounds is determined by its position in the metal reactivity series. The more reactive the metal, the harder it is to extract the metal from its compounds

-There are 2 methods for extracting metals from their ores:

1. Reduction of the molten metal compound
2. Electrolysis of the molten metal compound

Extraction of metals and the reactivity series

Most reactive

Reactivity decreases

Least reactive

Potassium	K
Sodium	Na
Calcium	Ca
Magnesium	Mg
Aluminium	Al
Zinc	Zn
Iron	Fe
Tin	Sn
Lead	Pb
Copper	Cu
Silver	Ag
Platinum	Pt
Gold	Au

Metals **more reactive** are extracted by **electrolysis** of their molten salts or molten ores

Metals **less reactive** are extracted by **reducing** the ore with **carbon** or carbon monoxide

The least reactive metals e.g. silver or gold even occur native i.e. unreactive

Electrolysis involves the use of large amounts of electricity and is a very expensive process compared to reduction using carbon. It is only used to extract very reactive metals because their compounds are too stable to be reduced using carbon.

When carbon is used to extract a metal from its metal oxide, a redox reaction takes place. Carbon is said to be the reducing agent as it reduces the metal oxide to the metal by removing oxygen from it

EXAMPLE: Copper (II) oxide and carbon $\underbrace{2CuO}_{\text{Black}} + C \rightarrow CO_2 + \underbrace{2Cu}_{\text{reddish brown}}$

SCRAP METALS AND RECYCLING

Metal ores are finite and limited and expensive to mine. It is essential that we recycle those scrap metals that are still useful. Iron, steel, copper and aluminium are the most easily recycled metals.

ADVANTAGES of RECYCLING

- Saves energy and reduces greenhouse gas emissions (CO_2)
- Preserves non-renewable material
- Reduces land degradation, air and water pollution through mining
- Reduces the amount of land fill required for disposal of scrap metal

Recycling is sometimes not feasible because of the costs involved. Transportation, sorting through waste and cleaning the scrap metal, etc. may cost more than extracting the metal from its ores. This is true for some cheaper metals.

8.4 Iron

EXTRACTION OF IRON

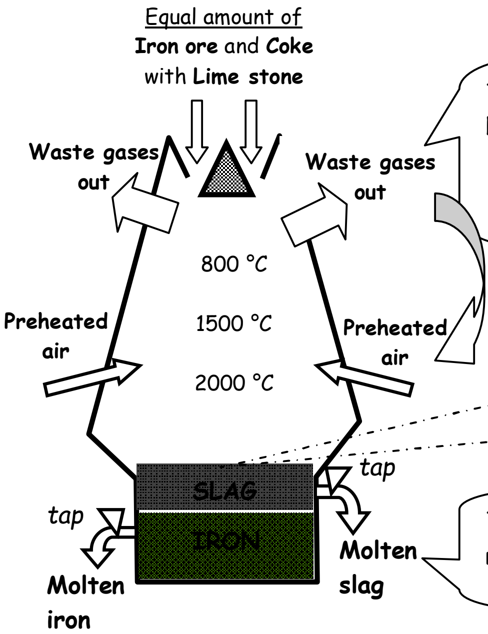
Iron is extracted from its ore **haematite**, Fe_2O_3, by reduction using carbon in a BLAST FURNACE

RAW MATERIALS

- Haematite (Fe_2O_3 containing iron(III))
- Coke (carbon)
- Limestone ($CaCO_3$)

BLAST FURNACE

-The diagram below shows the blast furnace



The **waste hot gases** are **used to heat** the **incoming hot air**.
(The toxic gas carbon monoxide recombines with oxygen to form the greenhouse gas carbon dioxide.)

The **molten slag** is less dense than the molten iron, so that it can **floats** on iron and be gained from the top tap.

The **molten slag** can be used for **road building** or to make fertiliser

The raw iron obtained in this process is known as **cast iron** or **pig iron**.
It can be purified further by bubbling oxygen gas through it to burn away impurities. The purified iron is then used to make alloys such as steel.

PROCESS

Let us see the process in terms of the reactions that take place in the furnace.

COKE (carbon)

Carbon burns in air to form carbon dioxide:

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

Carbon dioxide combines with more coke to form carbon monoxide:

$CO_2(g) + C(s) \rightarrow 2\underline{\mathbf{CO}}(g)$

LIMESTONE ($CaCO_3$)

Limestone decomposes to form calcium oxide and carbon dioxide:

$CaCO_3(s) \rightarrow \underline{\mathbf{CaO}}(s) + CO_2(g)$

IRON ORE (HAEMATITE)

-Iron ore (Haematite) contains IRON (III) OXIDE (Fe_2O_3) and IMPURITIES (e.g. sand SiO_2)

REDUCING IRON ORE TO IRON

[redacted]

REMOVING IMPURITIES

[redacted]

STEEL

The iron that is formed in a blast furnace is not hard enough to be used industrially as the metal is too soft, so that the iron is alloyed into steel.

Steel is an alloy made by mixing iron with carbon or other metals.

There are many types of steel depending on the type and amount of additives to it.
There are 2 kinds of steel; carbon steels and alloy steels. The chart below shows that.

STEEL

Carbon steels contain mainly iron and carbon

Mild steel is softer and more easily shaped. It is used to make car bodies and machinery.

High carbon steel is strong but brittle. It is used to make knives, hammers, chisels, saws and other cutting and boring tools.

Alloy steels contain iron and carbon and a transition element such as manganese, nickel, chromium, tungsten or vanadium.

Stainless steel is an example of an alloy steel.
It does not rust, is extremely durable and resistant to corrosion even upon heating. It is used to build chemical plants, in making of cutlery and surgical instruments.

RUSTING

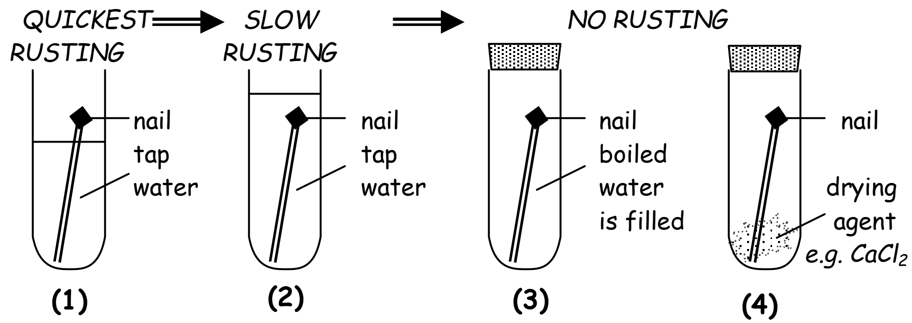
Rusting is the corrosion of iron or steel to form hydrated iron (III) oxide $Fe_2O_3 . nH_2O$

FORMATION OF RUST

Rusting is a redox reaction

For rusting to occur, both **AIR** (oxygen) and **WATER** must be present.

[EXPERIMENT]
Iron nails are put in various test tubes. Let's see formation of rusting in each of them.



RUST FORMS **NO RUST FORMS**

From the conditions for rusting (THE PRESENCE OF AIR AND WATER), you can tell the results above.

- In the test tube (1), enough amounts of air and water are there.
- In the test tube (2), it looks like that only water is there. But small amount of air can exist in water, so that rusting can occur.
- In the test tube (3), unlike the tube (2), there is no air although water is filled. It is because air that used to exist in water has been removed by boiling.
- In the test tube (4), there is no water with drying agent.

→ In the test tubes that satisfy the condition for rusting, rusting occur.

→ Unlike aluminium which reacts with oxygen in the air to form a protective layer on the metal surface, rust is brittle and flaky. The irons underneath will eventually rust and flake away

The overall reaction that takes place in rusting is given by the equation

$$4Fe(s) + 3O_2(g) + 2nH_2O(l) \rightarrow 2Fe_2O_3 . nH_2O(s)$$

- This is an oxidation reaction that takes place slowly. In this process, iron is first oxidised to iron(II) ions before the iron are further oxidised to iron(III) ions.

PREVENTING THE FORMATION OF RUST

There are 2 main ways of preventing rusting of iron or steel.

BARRIER PROTECTION

1. Coat the iron/steel object with a layer of substance present air and/or water from reaching the metal.
   [EXAMPLES] painting, oil or greasing
2. coat the iron/steel object with a less reactive metal or with plastic
   [EXAMPLE] steel food cans coated with tin (tin-plating)

SACRIFICIAL PROTECTION

Coat the iron/steel object with a more reactive metal. The more reactive metal will corrode in place of iron.

[EXAMPLE] galvanizing

Galvanizing is a method of protecting a metal (e.g. iron or steel) from corrosion by covering it with a thin layer of ZINC through dipping or electroplating.

In this method, a reactive metal is used for a coating metal, but not all reactive metals are suitable. For example, magnesium is not used as a coating on an iron or steel object because it will react with the oxygen in the air to form magnesium oxide. Magnesium oxide flakes easily and will come off the surface, exposing more magnesium for reaction. In this way, a magnesium coating will wear out very quickly. Hence magnesium is not suitable for the coating metal.

8.5 Copper

EXTRACTION OF COPPER

Copper is an **unreactive** metal so it can be **extracted** from its ore, by **heating with carbon**

COPPER ORES

- **CUPRITE**, Cu_2O (by heating with carbon)
- **MALACHITE**, $CuCO_3$-$Cu(OH)_2$ (by decomposing on heating)

PROCESSING COPPER ORES INDUSTRIALLY

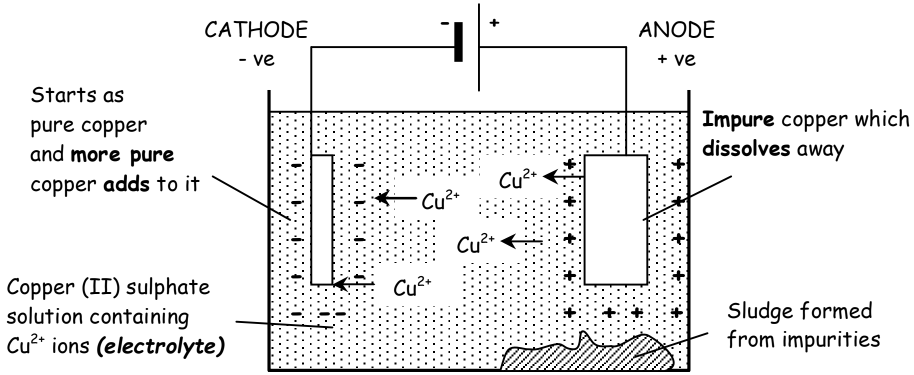
The two main ways to process copper ores industrially are:

- FLOTATION, roasting and SMELTING
- LEACHING with dilute sulphuric acid *(more commonly used in Zambia)* then using ELECTROLYSIS or adding SCRAP IRON

PURIFYING EXTRACTED COPPER INDUSTRIALLY

- **Very pure** copper is needed for **electrical conductors**
- **ELECTROLYSIS** is used to produce **VERY PURE COPPER**

ELECTROLYSIS to PURIFY COPPER



Pure copper is **deposited** on the pure **cathode (-)**

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

Copper **dissolves** from the impure **anode (+)**

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

STEPS:

+ The **ANODE** (**positive** electrode) is made from **impure** copper
  At this electrode, the copper atoms give up e^- to **form Cu^{2+} ions** which **dissolve** in the solution
  - These Cu^{2+} ions are then ATTRACTED to the negative electrode
– The **CATHODE** (**negative** electrode) starts as a thin piece of **very pure copper**
  At this electrode, Cu^{2+} ions gain e^- to **form Cu atoms** which **deposit** on the cathode which increases in size
  - The **impurities** in the anode fall to the bottom as a **sludge** as the anode **dissolves** away

USES OF PURE COPPER

Copper is used to make **electrical wiring** and **heat exchangers** because it is an excellent conductor of electricity and heat

COPPER ALLOYS

- **Brass** is an alloy of **copper** and **zinc** and is used to make **musical instruments** and **bimetallic strips**
- **Bronze** is an alloy of **copper** and **tin** and is used to make **trophies**
- Both of these alloys are **non-corroding**

8.6 Aluminium

Aluminium and its alloys have the following properties.

- It has low density
- It has good electrical and heat conductivity.
- It is resistant to corrosion.
- It is a relatively strong metal.

Electrolysis is **expensive** as it uses **lots of electricity**

EXTRACTION OF ALUMINIUM

Aluminium is a **reactive** metal so must be **extracted** from its ore by **electrolysis**

ALUMINIUM ORE

- **bauxite**, Al_2O_3

→ A **molten state** is needed for electrolysis. This can be very expensive. Al_2O_3 has a **very high melting point** over **2000°C**. Instead the Al_2O_3 is **dissolved** in **molten cryolite** (a less common ore of aluminium). This only requires a **temperature** of about **900°C**, which is much cheaper.

USES OF ALUMINIUM AND ALLOYS

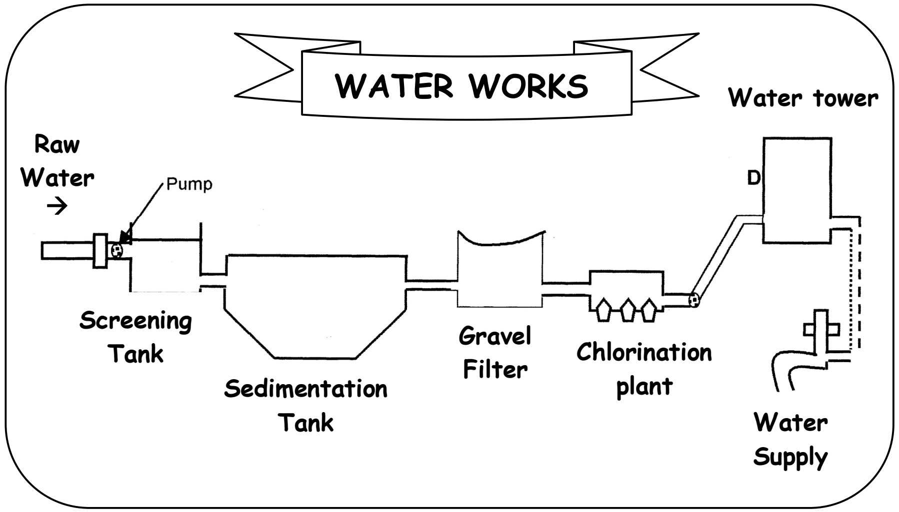
- **Overhead electrical cables** are made of aluminium as it is **lightweight** and a **good conductor** of electricity
- **Cooking utensils** and **food containers** are made of aluminium as it **does not corrode** (due to its protective oxide layer) and is a **good conductor** of heat
- **Aircraft** and **bicycle frames** are made from aluminium **alloy(duralumin)** as they are **strong** and **lightweight**

SUMMARY Reactivity Series and Reactions of Metals

METAL	REACTION WITH WATER	REACTION WITH DILUTE ACID	REACTION WITH AIR	ACTION OF HEAT ON CARBONATE
K	React with cold water	Violent reaction with dilute acids	Burns very easily with a bright flame	No reaction
Na				
Ca				
Mg	React with steam	React fairly well with dilute acids with decreasing ease		Decompose to form oxide and CO_2
Al			Burn slowly to form oxide	
Zn				
Fe	Reacts reversibly with steam		React slowly with air when heated	
Pb	No reaction with water or steam	May react slowly if warmed		
Cu		No reaction with dilute acids, may react with concentrated acids		
Ag		No reaction	No reaction	Decomposes to form Ag, O_2 and CO_2

9. NON-METALS

COTENT	LEARNING OBJECTIVE (Pupils should be able to)
Water	• Explain the effects of water pollution • Suggest ways of reducing water pollution • Describe in outline the purification of water supply in terms of filtration and chlorination. • State uses of water in industry and the home.
Air	• Describe the volume composition of clean air • Name common pollutants • State the sources of each of the following pollutants
Common Non-metals	• Explain the use of hydrogen in a manufacture of ammonia and of margarine and as fuel in rockets • Name the uses, oxygen tents in hospitals, and with acetylene (a hydrocarbon) in welding. • Describe the need for nitrogen, phosphorus and potassium compounds in plant life. • Describe the essential conditions for the manufacture of ammonia by the Haber process • Name the uses of ammonia in the manufacture of fertilisers such as ammonium sulphate and nitrate. • Discuss the effect of chemical fertilizers on the soil.



9.1 WATER

Water is the most abundant liquid on earth – it covers 70% of the earth's surface. Water is used at home for drinking, cooking, cleaning and washing. Now let us see this important liquid.

WATER PURIFICATION

WATER TREATMENT

Treatment of drinking water is carried out at the waterworks.

- Three main stages Sedimentation, Filtration and Chlorination are involved

3 STAGES IN WATER TREATMENT

RAW WATER

Raw water is first screened to remove large solid impurities.

↓

1. SEDIMENTATION

- Alum *(a coagulating agent)* is used to make solid particles stick together.
- The solid clumps sink to the bottom in the sedimentation tank and are removed.

↓

Lime (calcium oxide) is added to reduce acidity.
Activated carbon is added to remove foul odour and taste

2. FILTRATION

- The water is filtered to remove any remaining solid particles

↓

3. CHLORINATION

- Chlorine is added to kill germs and bacteria present in the water.

↓

CLEAR WATER

The water supplied to our homes does not have to be pure but safe to drink. It is called **potable** (drinkable).

Important Note: Chlorine is a **highly poisonous** substance. It is important to use the right quantity even when using '*Chlorine*' for chlorinating water at home.

DESALINATION

Desalination is the process of removing dissolved salts from seawater. The sea thus provides a ready source of drinking water.

-Two methods of desalination are commonly used.

1. Distillation; seawater is evaporated and the pure water vapour formed is condensed, e.g. solar distillation
2. Reverse osmosis; Pure water is extracted from seawater using a semi-permeable membrane under high pressure.

WATER POLLUTANTS

Water from rivers and lakes contains dissolved mineral salts, organic matter as well as some pollutants.

POLLUTANT	HARMFUL EFFECTS	SOURCE OF POLLUTANT
ACIDS	Aquatic life cannot survive in low pH water. Low pH water also causes poor growth of vegetation.	Acid rain
NITRATES AND PHOSPHATES	Causes eutrophication – excessive growth of vegetation which uses up dissolved O_2. This causes the fish to die. After the vegetation decays, the water becomes stagnant	Excess fertilisers washed off from crops
HEAVY METALS	Poisonous to mankind	Waste from industries involved in mining and processing metals
SEWAGE	Health problems such as infections. Can also cause eutrophication.	Untreated household waste and excretion from animals
OIL	Kills aquatic life as oxygen can non longer pass through and dissolve in water	Ships with oil spills

INDUSTRIAL USES OF WATER

Water is used in many different ways by industries.

INDUSTRIAL USES	EXAMPLES
As an essential ingredient for a product	• Beer making • Whisky production
Coolant	• Producing electricity in a coal or oil fired power station
Source of energy	• Hydroelectricity
Raw material in manufacturing process	• Paper manufacture
Polar solvent	• Dissolving ionic compounds

9.2 AIR

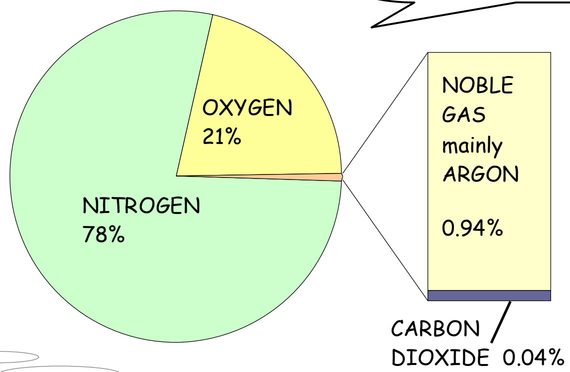
We human beings can not survive without air. Earth is surrounded by the atmosphere that contains air. Let's see this important gas.

COMPOSITION OF AIR

Clean, dry air is a mixture of gases.

Air also contains water vapour in variable amounts, depending on the humidity of the surroundings

% COMPOSITION OF AIR



The percentages of the gases that make up air will vary slightly form place to place, depending on local conditions.

AIR POLLUTION

Air is said to be polluted when it contains chemicals in high enough concentrations to harm living things or damage non-living things.

COMMON AIR POLLUTANTS

Common air pollutants include

- Sulphur dioxide SO_2
- Nitrogen oxides NO and NO_2
- Carbon monoxide CO
- Methane CH_4
- Lead compounds

SOURCES OF AIR POLLUTANTS

Now let's look at these sources of air pollutants respectively.

SULPHUR DIOXIDE SO_2

SOURCE:

- **Burning of fossil fuels** containing sulphur and sulphur compounds
  e.g. coal, natural gas or petroleum
  - **power stations**
  - **car exhaust**

EFFECTS:

- Sulphur dioxide irritates the eyes and causes **breathing difficulties**
- Produces **acid rain**

MINIMISING MEASURES:
→ treat exhaust gases with wet calcium hydroxide to remove SO_2

NITROGEN OXIDES NO and NO_2

SOURCE:

- At **high temperatures**, the N_2 and O_2 **in air combine** to form nitrogen oxides
  - **exhaust** from **car** engines
  - **power stations** and factories
  - **naturally** occur from **bush fires** and **lightning**

EFFECTS:

- Produces **acid rain**

MINIMISING MEASURES:
→ **Cars** use a **catalytic converter**

METHANE CH_4

SOURCE:

- Bacterial decay of **vegetable matter**, **animal dung** and **rubbish** buried in landfills

EFFECTS:

- It can combine with oxides of nitrogen in the presence of sunlight to form **photochemical smog.**
- It is also a green house gas that can cause **global warming.**

Global warming is the gradual change in world climate caused by the greenhouse effect.
It is thought that it may cause changes in weather patterns, causing droughts and storms, and even melting the polar ice caps to bring about severe flooding

CARBON MONOXIDE CO

SOURCE:

- Carbon monoxide is produced during the **INCOMPLETE combustion** of **carbon containing compounds** *i.e. insufficient O_2*
  - areas with high concentration of vehicles due to **car exhaust**
  - faulty gas appliances
  - areas where combustion takes place with poor ventilation
    e.g. using a brazier inside

EFFECTS:

- Carbon monoxide poisoning causes **suffocation**
  because the haemoglobin in our blood reacts more readily with CO than O_2

MINIMISING MEASURES:
→ Use more air during combustion
→ **Cars** use a **catalytic converter**

LEAD COMPOUNDS

SOURCE:

- **Lead compounds** are **added** to some **fuels** to make the car engines run properly

EFFECTS:

- Lead can cause **brain damage** and is especially harmful to young children

MINIMISING MEASURES:
→ Use **lead-free petrol**

ACID RAIN

.It is a long time since the effect of acid rain was known commonly. The acid rain causes serious damages to the environment over wide areas in the world. It is about time for us to face this global problem earnestly.

SOURCE:

- **Sulphur dioxide** in the air reacts with oxygen and water to **form sulphuric acid** which **dissolves** in rain clouds to form **acid rain** with a pH of 4 even down to 2
- **Nitrogen oxides** also form **nitric acid** with air to form **acid rain**

EFFECTS:

- **Corrodes metal structures** e.g. bridges and cars
- **Corrodes limestone buildings** as cement contains carbonate that readily react with the acids
- **Endangers aquatic life** as fish and plants can not survive in acidic water
- Causes the **soil** to become **acidic**, causing plants to die more readily

MINIMISING MEASURES:

→ **Remove** the **acidic gases**, NO_2 and SO_2, at the **source** (*see above points*)

→ **Reduce** water and **soil acidity** using **slaked lime**, $Ca(OH)_2$

➡ You know, there are many global problems. Apart form acid rain, another common problem on the Earth is global warming caused by the green house effect

- **The green house effect** is the trapping of heat energy in the atmosphere because of the effects of greenhouse gases. The infrared radiation (heat energy) is given off from the earth's surface as it is warmed up by the Sun

Green houses gases are gases in the atmosphere which absorb infra-red radiation, causing an increase in air temperature.
The most important is **carbon dioxide**, which is increased by burning fossil fuels and by deforestation, which reduces the amount of carbon dioxide removed by photosynthesis. Another is **methane**, a by-product of rice farming and cattle-rearing.

9.3 Common non-metallic elements

HYDROGEN

Hydrogen is the first and lightest element in the periodic table. It is the most abundant element in the universe (it is present in water and in all organic compounds) and is the main constituent of stars.

PREPARATION

⇨ In industry: Steam reforming method which is the reaction of methane and steam is applied.

$$CH_4(g) + H_2O(g) \rightarrow CO_2(g) + 4H_2(g)$$

⇨ In a laboratory: There are 3 simple ways

✧ Reaction of reactive metal and water

✧ Reaction of metal with acids

✧ Electrolysis acidified water

$$2H^+ + 2e^- \rightarrow H_2 \text{ (collects at anode)}$$

Check the topic of 'acid, base and salt'

IDENTIFICATION

When a lighted splint is held at the mouth of a test tube containing hydrogen gas, the gas burns explosively, making a **"pop" sound**.

USES

- Hydrogen is used in the **Haber Process** to produce ammonia
- **Hydrogenation** is used to change **vegetable oils** into **margarine**.
  -Vegetable oil is unsaturated and the hydrogen breaks the double carbon bonds to form saturated margarine.
- **Rockets** burn liquid hydrogen as a fuel with liquid oxygen to form water. This is a very **lightweight fuel**.

Hydrogen is so flammable that there would be a risk of explosion and it would be hard to liquefy for storage in modern fuel. But it may have a use as a common non-polluting fuel for road vehicles in the near future.

OXYGEN

Oxygen is the most important gas in the air. It is a colourless, odourless gas.

PREPARATION

⇨ In industry: Fractional distillation of liquid air is applied

⇨ In a laboratory: Thermal decomposition with catalyst of manganese dioxide.

✧ Heating Hydrogen peroxide onto manganese dioxide powder

$2H_2O_2 \rightarrow 2H_2O + O_2$ *(MnO_2 as catalyst)*

✧ Heating Potassium chlorate with manganese dioxide

$2KClO_3 \rightarrow 2KCl + 3O_2$ *(MnO_2 as catalyst)*

IDENTIFICATION

When a **glowing splint** is held at the mouth of a test tube containing oxygen gas, the splint **relights**.

USES

- **Oxygen cylinders** in **hospitals** help people with **breathing problems**
- To burn acetylene gas in an **oxyacetylene torch** when **welding steel**
- **Rockets** burn liquid oxygen as a fuel with liquid hydrogen to form water. This is a very **lightweight fuel**.
- Oxygen masks are used in an aircraft if there is an air leak / low pressure
- To kill bacteria in the treatment of sewerage
- Used in the production of steel to oxidise any impurities in iron before producing the type of steel required

THREE CHEMICAL PROCESSES INVOLVING OXYGEN:

1. **COMBUSTION** takes place when any substance **reacts with oxygen** to **produce heat**. If **flames** are produced it is called **burning**.

e.g. $C + O_2 \rightarrow CO_2$ $\quad 2H_2 + O_2 \rightarrow 2H_2O$ $\quad 4Na + O_2 \rightarrow 2Na_2O$

$S + O_2 \rightarrow SO_2$ $\quad 2Mg + O_2 \rightarrow 2MgO$ $\quad 2Fe + 3O_2 \rightarrow Fe_2O_3$

2. **RESPIRATION** is the **oxidation of sugars** in our body to **produce energy**.

$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + energy$

(This is the reverse of photosynthesis).

3. **RUSTING** occurs when **iron** comes into contact with **water and oxygen** to **form rust**. Rust is hydrated iron (iii) oxide, $Fe_2O_3.nH_2O$.

Combustion, respiration and rusting are all processes using up oxygen.

→ Combustion of fuels and respiration produce carbon dioxide. However the approximate composition of gases in air remain unchanged overall because **PHOTOSYNTHESIS** by green plants **converts carbon dioxide** back into **oxygen** and sugar using sunlight.

$6CO_2 + 6H_2O + energy \rightarrow C_6H_{12}O_6 + 6O_2$

NITROGEN

Nitrogen is the first element in Group V of the periodic table. It is a colourless, odourless gas which makes up 78% of the air. It is an unreactive gas but does have some uses.

PREPARATION

⇨ In industry: Fractional distillation of liquid air is applied

IDENTIFICATION

It does **not support burning** of other substances.

AMMONIA

Ammonia is a colourless, pungent gas, NH_3, that is less dense than air.

It is the most soluble of all gases and dissolves in water to form an alkali called aqueous ammonia NH_3 (aq).

It is only common alkaline gas and makes most red litmus paper blue.

Commercially ammonia is very important and prepared by **HABER PROCESS**

[redacted]

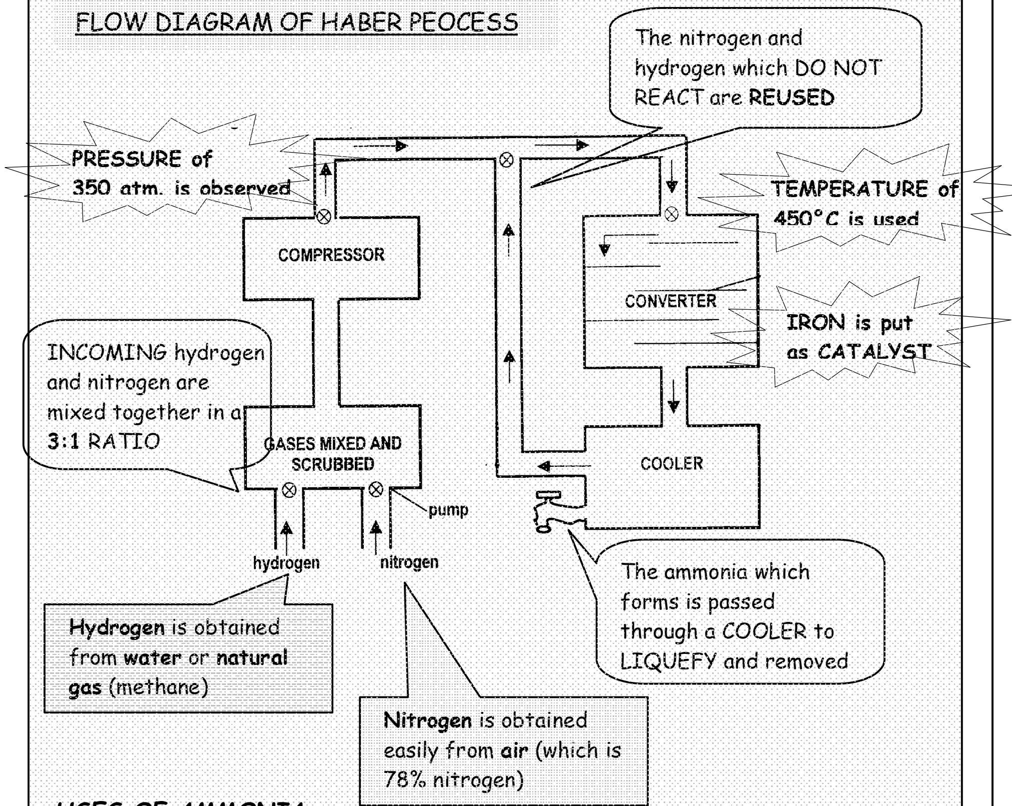
HABER PROCESS

This is the process for the manufacture of ammonia gas from direct combination of **NITROGEN** and **HYDROGEN** gases.

The **Haber Process** is a **REVERSIBLE reaction**

$$N_2(g) + 3H_2(g) \Longleftrightarrow 2NH_3(g) + heat$$

FLOW DIAGRAM OF HABER PEOCESS



USES OF AMMONIA:

- to make fertilisers *e.g. ammonium nitrate, ammonium sulphate*
- to make nitric acid

FERTILISERS

Plants need three **essential** elements: **nitrogen**, **phosphorous** and **potassium**. Because of a growing demand for food to feed an increasing population, farmers need to rely on fertilisers to provide essential elements needed for crops.

Ammonium nitrate, NH_4NO_3, is an especially **good fertiliser** as is contain nitrogen from two sources (NH_4 and NO_3).However **EXCESS nitrate fertiliser** washed into streams and rivers can cause **EUTROPHICATION**.

Eutrophication is when the excess fertilisers cause the plant life to grow too much, they then die and bacteria then takes over processing the decaying matter, this uses up the oxygen and causes the animal life to also die. The water then becomes stagnant.

CARBON

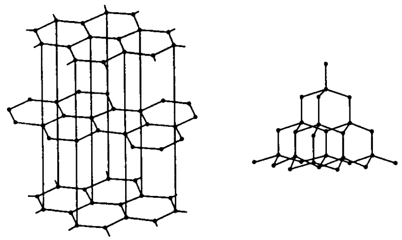
Carbon is the lightest non-metallic element in GroupIV of the periodic table.
It forms the basis of life chemistry. It forms allotropes.

Text box

ALLOTROPES

Allotropes are solid forms of an element with **different molecular structures**.
DIAMOND and GRAPHITE occur naturally as allotropes of carbon.

- DIAMOND is suitable for
  - ⇨ ***Cutting and Grinding tools*** because it is the **hardest** naturally occurring substance
- GRAPHITE is suitable for
  - ⇨ **Lubricant** because it is **soft and flaky** due to a layered structure which is held by a week interaction.



GRAPHITE **DIAMOND**

LIME

[redacted]

10. ORGANIC CHEMISTRY

COTENT	LEARNING OBJECTIVE (Pupils should be able to)
Introduction of Organic Chemistry	• Describe a homologous series as a group of compounds with a general formula, similar chemical properties and showing a gradation in physical properties. • Describe the general characteristics of any homologous series • Define the functional group. • Name, and draw the structure of the organic compounds
Hydrocarbons	• Describe the properties of alkans • Describe the properties of alkenes • Distinguish between saturated and unsaturated hydrocarbons: • Describe the manufacture of alkenes and of hydrogen by cracking hydrocarbons. • Name natural gas and petroleum as sources of fuels • Describe the separation of petroleum by fractional distillation. • Name the uses of petroleum fractions
Alcohols and Acids	• Describe the properties and use of alcohols • Describe formation of ethanol by fermentation • Describe the formation of ethanoic acid by the oxidation of ethanol • Describe the reaction of ethanoic acid with ethanol to form the ester, ethyl ethanoate
Polymer	• Describe the structure of the polymer product from a given monomer and vice versa. • Describe the pollution problems caused by non-biodegradable plastics. • Identify carbohydrates, proteins and fats as natural polymers • Describe the formation of addition polymers • Describe the formation of condensation polymers • Describe some natural polymers as possessing the same linkages as some synthetic polymers • Describe the hydrolysis of carbohydrates gives simple sugars • Describe the hydrolysis of proteins to amino acids • Describe soap as a product of hydrolysis of fats

10.1 Introduction of Organic Chemistry

Organic Chemistry is the branch of chemistry concerned with the compounds of carbon (except carbonates and oxides of carbon). *"Organic" relates to living "organisms," and all organic compounds are or have been associated with living material.*

Food, fibres, fuels, tyres, plastics and most medicines are all carbon containing compounds known as organic compounds. Although they are made of a few elements such as carbon, hydrogen and oxygen, there are such a large variety of compounds. It is because of the ability of carbon atoms to form strong covalent bonds.

Here we learn general properties of organic compounds and try to classify them.

GENERAL PROPERTIES of ORGANIC COMPOUNDS:

- They do **not conduct electricity**
- They have **low melting point**
- They are **flammable** and **volatile** (evaporate so easily)
- They are **insoluble in water** but soluble in organic solvents like ethanol, acetone, etc.
- Most of them burn forming **carbon dioxide and water**
  [E.g.] Methane : $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

HOMOLOGOUS SERIES

Homologous Series is a group of compounds with increasing number of carbon atoms where each member differs from the next consecutive member by another $-CH_2$ unit.

The chemical and physical properties of an organic compound are determined by its FUNCTIONAL GROUP. Organic compounds with the same functional group are grouped into a family called a HOMOLOGOUS SERIES.

There are many homologous series and each series is given a name.

All homologous series have the following characteristics:

1 They have the same general formula.
2 They have similar chemical properties because they have the same **functional group**, i.e. they undergo the same type of reactions.
3 They show a trend in physical properties as the molecular mass increases.
- As the number of carbon atoms INCREASES:
  - Melting points and boiling points INCREASE
  - Flammability DECREASES (don't catch fire so easily)
  - Viscosity INCREASES (don't flow so easily)
  - Volatility DECREASES (don't evaporate so easily)

A functional group is an atom or group of atoms that give an organic molecule its **typical chemical properties**.

Table lists some homologous series of organic compounds.

HOMOLOGOUS SERIES	GENERAL FORMULA	FUNCTIONAL GROUP
ALKANES	C_nH_{2n+2}	**Nil**
ALKENES	C_nH_{2n}	***-C=C-*** Double bonds
ALCOHOL	$C_nH_{2n+1}OH$	***-OH*** Hydroxyl group
CARBOXYLIC ACIDS	$C_nH_{2n+1}COOH$	***-C=O*** / | / **O-H** Carboxyl group
ESTER	$C_nH_{2n+1}COOC_mH_{2m+1}$	***-COO-*** Ester functional group

You could also say -C-C- Single Bonds

Table : Common homologous series

'n' stands for NO. of Carbon atom in a molecule of the compound

A FEW OF RULES

NAMING ORGANIC COMPOUNS

Organic compounds are named according to **how many carbon atoms** they contain and which **functional group** they possess. Table below gives the prefixes and the suffixes assigned.

PREFIX (start with) +	*SUFFIX (end with)*
1 C atom → "***METH***" 2 C atoms → " ***ETH***" 3 C atoms → " ***PROP***" 4 C atoms → **"*BUT*"** 5 C atoms → **"PENT"**	Alkane → "***ane***" Alkene → "***ene***" Alcohol → "***ol***" Carboxylic acid → "***oic acid***"

[EXAMPLE 1]
An organic molecule belongs to the alcohol series and contains <u>4 carbon atoms.</u> Since the names of <u>alcohols</u> end with '-ol', the molecule will be called **butanol**.

[EXAMPLE 2]
The name of the molecule with formula C_2H_5 COOH is **propanoic acid**, since it contains <u>3 carbon atoms</u> and belongs to the <u>carboxylic acid</u> series.

DRAWING STRUCTURAL FORMULAE OF ORGANIC COMPOUNDS

To describe organic compounds, we often use STRUCTURAL FORMULAE as well as molecular formulae.

You can find the **same number of lines** as bond around the atom

RULES: * ***Each C*** *has* ***FOUR (4)*** *bonds*
* ***Each H, OH*** *or* ***COOH*** *has* ***ONE (1)*** *bond*

STEPS: [E.g.] Ethane C_2H_6, Ethene C_2H_4 and Ethanol C_2H_5OH

1. Write the ***correct No.*** *of* ***C atoms*** *for compound*

C C (Ethane) **C C** (Ethene) **C C** (Ethanol)

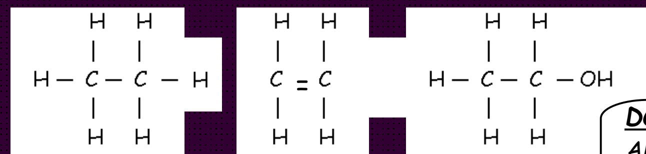
2. ***Draw*** *the* ***bonds connecting*** *these atoms*
→ ***REMEMBER DOUBLE BOND*** *for* ***ALKENES***

C – C   C = C   C – C

3. ***Add*** *in any* ***functional groups*** *(OH or COOH) (usually to last C atom)*

C – C   C = C   C – C ***-OH***

4. Lastly ***fill in*** *the correct no. of* ***H atoms*** *(so that each C atom has 4 bonds)*



Ethane   Ethene   Ethanol

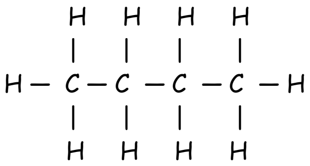
DOUBLE-CHECK;
ALL bonds are drawn (no missing bonds)

ISOMERS

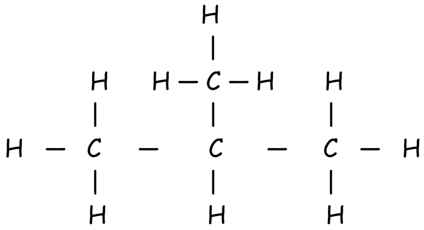
Isomers are **different compounds** which have the **SAME MOLECULAR formula** but **DIFFERENT STRUCTURAL formula**

- The more carbon atoms the more isomers which are possible

[EXAMPLE 1] *C_4H_{10}*



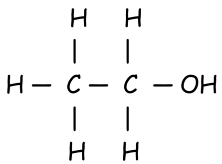
Butane
B.P. = -0.5°C



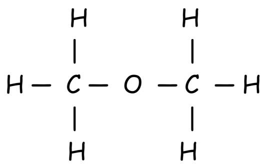
Isobutane
B.P. = - 12°C

Isomers of the SAME HOMOLOGOUS SERIES have SIMILAR CHEMICAL PROPERTIES
but DIFFERENT PHYSICAL PROPERTIES (like B.P, M.P.)

[EXAMPLE 2] *C_2H_6O*



Ethanol
In ALCOHOL series



Diethyl ether
In ETHER series

Isomers of DIFFERENT HOMOLOGOUS SERIES have DIFFERENT CHEMICAL PROPERTIES

10.2 Hydrocarbons

Oil is an essential item to us. Crude oil is mainly composed of some hydrocarbons.
It can be separated into Petrol, Kerosene, Diesel oil and so on.
We are going to see the hydrocarbons that are basic organic compounds.

Hydrocarbons are organic compounds that contain only carbon and hydrogen atoms

All hydrocarbons have covalent molecules. They are found naturally in PETROLEUM and NATURAL GAS.

ALKANES and ALKENES are HYDROCARBONS (contain only C and H atoms)
Alcohols and Carboxylic Acids are **not** hydrocarbons (contain O atoms as well)

We can classify hydrocarbons in terms of bonding.
We will see this issue later.

SATURATED Hydrocarbons (e.g. *ALKANES*)

- Contain ONLY SINGLE C-C bonds

UNSATURATED Hydrocarbons (e.g. *ALKENES*)

- Contain DOUBLE **C=C** bonds

ALKANES *General formula:* C_nH_{2n+2}

Where 'n' is the number of carbon atoms in one molecule.

The alkanes are a family of hydrocarbons, i.e. they contain hydrogen and carbon atoms only.
They are the main hydrocarbons found in petroleum and natural gas.

PROPERTIES OF ALKANES

- Alkanes have ALL **C-C SINGLE** BONDS
- Alkanes are **insoluble** in water
- Alkanes become more viscous, i.e. more difficult to pour out as the number of carbon atoms increase.

Table below shows the first 4 members of the alkane series.

No. of C atoms	ALKANE	CHEMICAL FORMULA	STRUCTURAL FORMULA	Physical State at Room Temperature
1	Methane	CH_4	H │ H – C – H │ H	GAS
2	Ethane	C_2H_6	H H │ │ H – C – C – H │ │ H H	GAS
3	Propane	C_3H_8	H H H │ │ │ H – C – C – C – H │ │ │ H H H	GAS
4	Butane	C_4H_{10}	H H H H │ │ │ │ H – C – C – C – C – H │ │ │ │ H H H H	GAS

Table; Properties of the first 4 members of the alkane series.

Every name ends with '**-ane**'.

Alkanes are covalent compounds **with weak intermolecular forces** between the molecules. As the number of carbon increase, <u>the melting point and boiling point increase;</u> the first four members are gases, the next thirteen members are liquids and the rest are solids.

The carbon atoms in alkanes are held together only by **–C-C- SINGLE COVALENT BONDS**. Thus alkanes are said to be **SATURATED**.

An organic molecule is said to be saturated if it contains only single carbon-carbon covalent bonds. In all organic compounds, each carbon atom will form 4 covalent bonds, while H will form 1 covalent bond. If oxygen atoms are present, each oxygen atom will form 2 covalent bonds.

CHEMICAL REACTION OF ALKANES

Alkanes are fairly unreactive molecules as their single bonds are strong. They are used mainly as fuels to provide heat energy.

Thus they don't form polymers

<u>COMBUSTION OF ALKANES</u>

Alkanes burn in air (oxygen) to form carbon dioxide and water.

[EXAMPLE 1] Methane + oxygen → carbon dioxide + water vapour

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Alkanes can be used as **fuels**

[EXAMPLE 2]

When there is not enough air, burning is incomplete. In this case, soot and carbon monoxide are also produced.

Ethane + insufficient oxygen

→ carbon + carbon monoxide + carbon dioxide + water vapour

$3C_2H_6 + 6O_22 \rightarrow 4C + CO + CO_2 + 9H_2O$

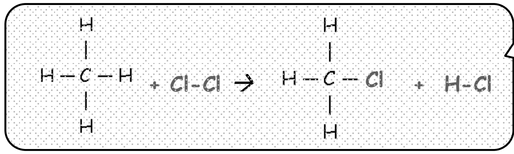
<u>SUBSTITUTION REACTION</u>

In presence of **SUNLIGHT**, it undergoes a Substitution Reaction with **chlorine** to form chloroalkanes. (*i.e. H atoms replaced by Cl atoms*)

[EXAMPLE]

Methane + Chlorine → Chloromethane + hydrogen chloride

$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$



This reaction does not take place in the dark. Sunlight is needed to provide **energy to break the Cl-Cl bond** to produce chlorine atoms which then react with the alkane molecule.

ALKENES *General formula:* C_nH_{2n}

The alkenes also form a family of hydrocarbons-they contain only carbon atoms and hydrogen atoms.
They are formed when petroleum fractions undergo cracking.

PROPERTIES OF ALKANES

Table below shows the first 3 members of the alkene series.

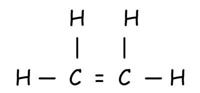
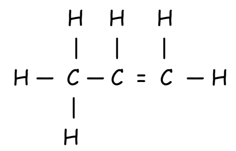
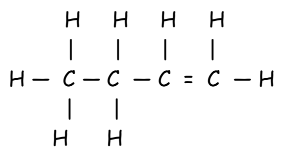
No. of C atoms	ALKENE	CHEMICAL FORMULA	STRUCTURAL FORMULA	Physical State at Room Temperature
2	Ethene	C_2H_4		GAS
3	Propene	C_3H_6		GAS
4	Butene	C_4H_8		GAS

Table: Properties of the first 3 members of alkene family

Every name ends with '**-ene**'

<u>Note that</u> alkene family starts with ethene where **n=2**.
Methene, where **n=1** to give the formula CH2, does <u>not</u> exist

The formulae of each member differs from the previous one by an extra **-CH₂-** group

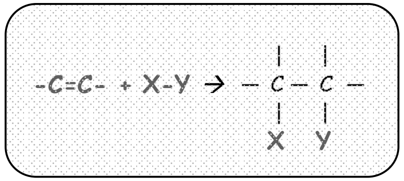
The alkenes contain <u>carbon – carbon double bonds (- C = C -).</u> This carbon double bond is known as the **functional group of the alkene** family. All alkenes must have this functional group.

Any organic compounds with a **CARBON = CARBON DOUBLE BOND** is said to be **UNSATURATED**. If a molecules has more than one set of carbon – carbon double bonds, it is said to be polyunsaturated.

CHEMICAL REACTION OF ALKENES

Alkenes are more reactive than alkanes because of the **carbon = carbon double bond.** The reaction of alkenes takes place **at** the carbon = carbon double bond. During a reaction, the carbon = carbon double bond **opens up**, allowing the addition of other molecule onto the alkenes:

Thus they can form polymers



→ The unsaturation in the alkene molecule is destroyed. A saturated product is formed in which <u>the double bond is replaced by single bonds.</u> An addition reaction is said to have taken place.

An addition reaction is a reaction in which one molecule adds to another to form a single molecule product.

In addition reactions, molecules are always added across a carbon = carbon double bond, i.e. the addition is across adjacent carbon atoms. Hence the final structure of the product will always take the appearance above.

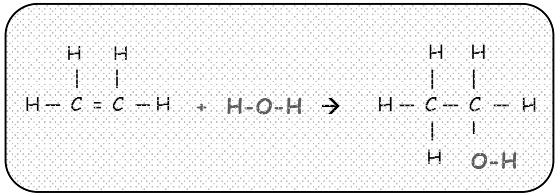
ADDITION OF STEAM

Alkenes react **with water (steam)** in the presence of phosphoric (V) acid (H_3PO_4) catalyst at high temperature and pressure **to form alcohols.**

[EXAMPLE]

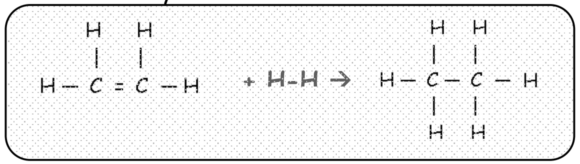
$C_2H_4 + H_2O \rightarrow C_2H_5OH$

ethene + steam → ethanol



ADDITION OF HYDROGEN

Alkenes undergo addition reaction **with hydrogen gas** in the presence of a nickel catalyst **to form alkanes.**



[EXAMPLE]

$C_2H_4 + H_2 \rightarrow C_2H_6$

ethene + H_2 → ethane

→ This process is known as **HYDROGENATION.**

Hydrogenation is used in MARGARINE manufacture to change UNSATURATED VEGETABLE OILS into a solid product.

COMBUSTION OF ALKENES

Alkenes **burn** in plenty of air (oxygen) to form **carbon dioxide** and **water**.

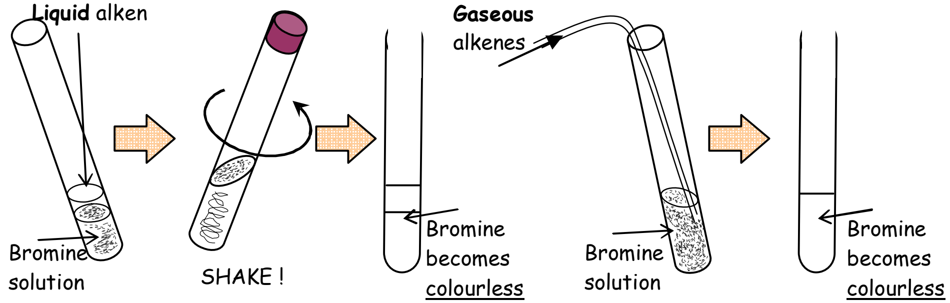
[EXAMPLE] ethane + oxygen -> carbon dioxide + water vapour

$$C_2H_4\,(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$$

Alkenes will produce soot and carbon monoxide when there is insufficient oxygen for complete combustion.

TEST FOR UNSATURATION

We can use the addition reaction as a test to find out if a hydrocarbon is an alkane or alkene. Fig below shows the testing process.



The **reddish-brown colour** of the bromine solution is **decolourised** as the bromine is used in the reaction

- This is an addition reaction of bromine

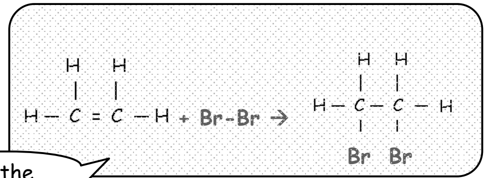
ADDITION OF AQUEOUS BROMINE

Alkenes undergo addition reaction with the aqueous bromine.

[EXAMPLE]

$C_2H_4 + Br_2 \rightarrow C_2H_4Br_2$

Ethene +bromine → 1,2-dibromoethane



The bromine molecule adds onto the double bond of the ethene molecule.

→*FROM THE OBSERVATIONS*

→ [redacted]

ALKENE → • The reddish brown colour of bromine is quickly decolourised, i.e. the colour of the mixture in the test tube **changes from reddish brown to colourless.**

ALKANE → • There is **no reaction**. Alkanes do not undergo addition reactions because they are saturated.

PREPARATION OF ALKENES

Alkenes are formed when petroleum fractions undergo CRACKING, while Alkanes are the main hydrocarbons found in petroleum and natural gas.

CRACKING

Big hydrocarbon molecules can be **broken up** into **smaller molecules** by a process called cracking. The big molecules are passed over a solid CATALYST (aluminium oxide or silicon (V) oxide) at a high temperature (about 600℃), where they break up to give smaller molecules.

The products of cracking CAN NOT BE PREDICTED accurately.
What we know is that at the end of the process, smaller hydrocarbon molecules (either alkanes or alkenes) and/or hydrogen may be formed.

[EXAMPLE 1] Big alkane → Smaller alkane + Alkene

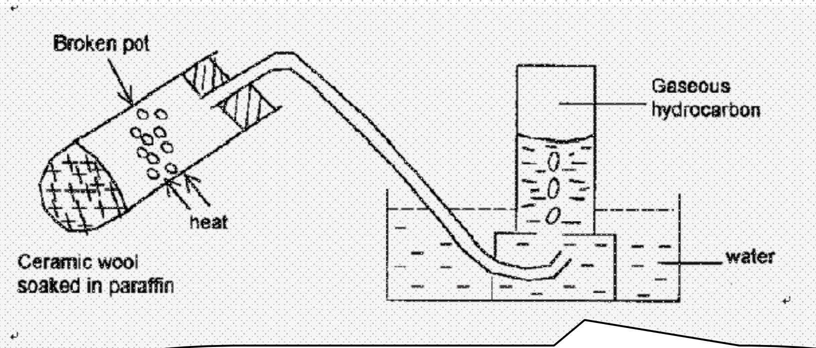
$C_{20}H_{42}$ → $C_{12}H_{26}$ + C_8H_{16}

⇨ Cracking is also used to make hydrogen gas

[EXAMPLE 2] Big alkane → Alkene + Alkene + Hydrogen

$C_{18}H_{38}$ → C_8H_{16} + $C_{10}H_{20}$ + H_2

➡ In a laboratory, in order to form alkenes from paraffin oil (big alkane $C_{10}H_{22}$), Cracking takes place in the way (CATALYTIC CRACKING) shown below.



On heating over a suitable catalyst, it will break down into smaller molecules. These molecules may include alkenes.

• Cracking is essential to match the demand for fractions containing smaller molecules from the refinery process. Some of these smaller molecules are used as chemical feedstock, while others are used to produce high grade petrol for motor vehicles.

SOURCES OF ENERGY

Most of our energy comes from the burning of CRUDE OIL and NATURAL GAS. In some countries, solid COAL is used as fuel.

Crude oil is also known as PETROLEUM

FOSSIL FUELS

Fossil fuels are found in the form of **crude oil**, **natural gas** and **coal**

→They are formed as dead plant and animal material are subjected to intense pressure and heat over millions of years.

Consequently, these fuels are NON—RENEWABLE energy sources.

NATURAL GAS

Natural gas is mostly **methane gas (CH_4)**. It burns cleanly in air to form carbon dioxide gas and water: $CH_4(g)$ + $2O_2(g)$ → $CO_2(g)$ + + $2H_2O(g)$

This reaction is highly EXOTHERMIC.

COAL

Coal is mainly carbon, with small amounts of hydrogen, oxygen, nitrogen and sulphur. When it burns in air, the main products are carbon dioxide and water:

Coal + Oxygen in air → Carbon dioxide + Water

At the same time, small amounts of soot, oxides of sulphur and nitrogen and ash (a solid residue) are formed.

- Coal is not a clean fuel. The sulphur dioxide and nitrogen dioxide gases present in the waste gases of a coal burning power station are removed by passing them through wet limestone before the waste gases are emitted into the atmosphere.

CRUDE OIL (petroleum)

Crude oil (petroleum) is **a mixture of hydrocarbons** with different carbon chain length. Petroleum is quite useless as a mixture;

→ It is usually refined **by fractional distillation** to **separate** out its different compounds to make useful fuels and petrochemicals.

- Crude oil is separated into 7 fractions

Boiling Point	FRACTION	USE	No. of C atoms
Below 40°C	Petroleum Gas	Gas fuel	1 ~ 3
40 -75 °C	Petrol / Gasoline	Car fuel	4 ~ 8
75-150°C	Naphtha	Chemical feed stock	7 ~ 14
160-250°C	Paraffin / Kerosene	Stove fuel / Jet fuel	11~ 15
250-300°C	Diesel	Diesel fuel	16 ~ 20
300-350°C	Lubricant oil	Lubricating oil, waxes and polishes	20 ~ 35
Over350°C	Bitumen	Making roads	More than50

The process is unable to give pure fractions because the **boiling points** of the hydrocarbons found in crude oil are **too close for efficient separation**.
Hence, in table, the data quoted shows a range of boiling points instead of a single boiling point.

Molecules with a **short** carbon chains have **low boiling points** while those with **long** carbon chains have **high boiling points**.

10.3 Alcohols and Carboxylic acids

Most of organic compounds in living things contain oxygen. What we have seen are hydrocarbons which have carbon and hydrogen only. Now we are going to see Alcohols and Carboxylic acids which contain oxygen.

ALCOHOLS *General formula:* $C_nH_{2n+1}OH$

Alcohols are
- **colourless**, **flammable liquids**
- **good solvent** and **fuel**.
- **soluble in water**.

Functional group of the alcohols

-OH (hydroxyl group)

PROPERTIES OF ALCOHOLS

Table below shows the first 4 members of the alcohol series.

No. of C atoms	ALCOHOL	**CHEMICAL FORMULA**	**STRUCTURAL FORMULA**	**Physical State at Room Temperature**
1	Methanol	CH_3OH	H | H – C – OH | H	LIQUID (B.P.=64℃)
2	Ethanol	C_2H_5OH	H H | | H – C – C – OH | | H H	LIQUID (B.P.=78℃)
3	Propanol	C_3H_7OH	H H H | | | H – C – C – C – OH | | | H H H	LIQUID (B.P.=97℃)
4	Butanol	C_4H_9OH	H H H H | | | | H – C – C – C – C – OH | | | | H H H H	LIQUID (B.P.=117℃)

Every name ends with '**-ol**'.

Table: Properties of the first 4 members of alcohol family

- Alcohols are NOT hydrocarbons since they contain oxygen atoms in addition to carbon and hydrogen atoms.
- Alcohols are NOT alkalis even though they contain the – OH group.

They are NEUTRAL liquids

Alcohols do not have the same empirical formula. For example, METHANOL is **CH_3OH**, or **CH_4O**. ETHANOL is **C_2H_5OH** or **C_2H_6O**. Both formulae cannot be reduced to any simpler form.

As the number of carbon atoms in the alcohol increases,
1 the boiling point increases
2 the solubility in water decreases

CHEMICAL REACTION OF ALCOHOLS

COMBUSTION OF ALCOHOLS

Alcohols burn in plenty of air (oxygen) to give **carbon dioxide** and **water vapour**.

[EXAMPLE] Ethanol + oxygen → carbon dioxide + water vapour

$2C_2H_5OH + 6O_2 \rightarrow 4CO_2 + 6H_2O$

The reaction gives out lots of heat energy and is exothermic. In some countries such as Brazil, ethanol is sometimes used as a fuel in cars in place of petrol.

OXIDATION OF ALCOHOLS

Alcohols can be oxidized **to carboxylic acids**.

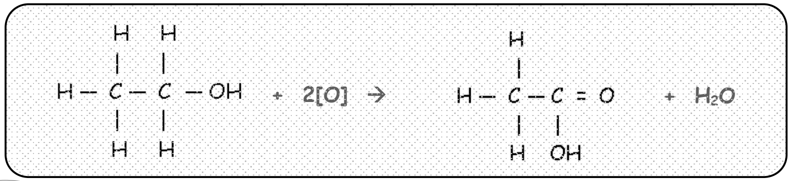
This reaction takes place in the presence of an oxidizing agent such as acidified potassium manganate (VII) or acidified potassium dichromate (VI)

[EXAMPLE]

Ethanol can be oxidized into an organic acid called ethanoic acid

$$C_2H_5OH + 2[O] \rightarrow CH_3COOH + H_2O$$

ethanol + from oxidising agent → ethanoic acid + water



Oxidation of alcohols can also take place they are left exposed to oxygen in the air for a few days.

For example, if ethanol is left exposed in the air it turns "SOUR." This is the common fate of wines and beers which are opened but no drunk. The reason is that the ethanol has been oxidised to ethanoic acid

$$C_2H_5OH + O_2 \rightarrow CH_3COOH + H_2O$$

The product is a dilute solution of ethanoic acid called vinegar. This reaction takes place in the presence of bacteria in the air.

ETHANOL C_2H_5OH

This is the commonest alcohol, and is a colourless, water-soluble liquid.

USE of ethanol is:

1 as fuel for vehicles
2 as solvent for paints and varnishes
3 in alcoholic drinks such as beer and wine

→ Ethanol is **Produced** by **fermentation** or the **addition reaction** of an **ethene** with **steam**

It is sometimes more economical to produce ethanol from ethene gas obtained by cracking petroleum fractions.

As for Process of the addition reaction, you can refer to ' ALKENE'

FERMENTATION

This is one of methods to prepare ethanol. This process is used all over the world for baking, wine-making, and brewing beer.

Fermentation is the conversion of sugars into ethanol and carbon dioxide gas by the action of micro-organism such as yeast, in the absence of air.

A solution containing glucose (a sugar) is mixed with water and yeast and allowed to react for a few days in the absence of air.

$$\text{Glucose} \xrightarrow[37°C]{\text{YEAST}} \text{Ethanol} + \text{Carbon dioxide}$$

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

If air is present in the mixture, oxidation of ethanol by the bacteria in the air will take place and the end products will be water and ethanoic acid.

CONDITIONS FOR FERMENTATION

- This process must take place **in the absence of air (oxygen).**
- This process takes place **at an optimum temperature of 37 ℃.**

If the temperature goes above 40 ℃, the enzymes in yeast which catalyse the reaction become denatured so that they can no longer act as catalysts.

The fermentation of sugars produces only a dilute solution of ethanol (up to 15%). When the ethanol content exceeds this value, the yeast dies and fermentation stops. Higher concentrations of ethanol can be obtained by fractional distillation of the solution.

CARBOXYLIC ACIDS

General formula: $C_nH_{2n+1}COOH$

The carboxylic acids form a homologous series.
They are generally WEAK ACIDS.
So they exhibit **normal acidic properties**.

They exist mainly as molecules and do not form hydrogen ions as easily as mineral acids. That is why the acidity of compounds in this series is weak

Functional group of the Carboxylic acids

-COOH or – C = O
|
OH

DO YOU REMEMBER?

Since solution of carboxylic acids are acidic, they will undergo **typical reactions of acids** – they will react with metals above hydrogen in the reactivity series to form hydrogen, with metal carbonates to form salt, carbon dioxide and water, and with bases to form salt and water.

PHYSICAL PROPERTIES OF CARBOXYLIC ACIDS

Table below shows the physical properties of the first 4 members of the series

No. of C atoms	CARBOXYLIC ACID	CHEMICAL FORMULA	STRUCTURAL FORMULA	Physical State at Room Temperature
1	Methanoic Acid	HCOOH	H – C = O │ OH	LIQUID (B.P.=101℃)
2	Ethanoic Acid	CH_3COOH	H │ H – C – C = O │ │ H OH	LIQUID (B.P.=118℃)
3	Propanoic Acid	C_2H_5COOH	H H │ │ H – C – C – C = O │ │ │ H H OH	LIQUID (B.P.=141℃)
4	Butanoic Acid	C_3H_7COOH	H H H │ │ │ H – C – C – C – C = O │ │ │ │ H H H OH	LIQUID (B.P.=164℃)

Table: Properties of the first 4 members of carboxylic acid family

The general formula for the carboxylic acids is $C_nH_{2n+1}COOH$, where **'n' STARTS WITH 0** for the first member of the series.

The first 4 members are all liquids at room temperature.
As the number of carbon atoms in the molecule increase, the boiling point increase

The most important carboxylic acid is ETHANOIC ACID. It is used for **flavourings** and as a **preservative**

PREPARATION OF CARBOXYLIC ACIDS

Carboxylic acids are prepared by oxidation of alcohols.

You can check the previous page!

OXIDATION IN AIR

When a solution of a carboxylic acids, for example, ethanol is exposed to air, the oxygen present slowly oxidises ethanol into ethanoic acid in the presence of bacteria. Vinegar, which is a solution of ethanoic acid in water, is made this way

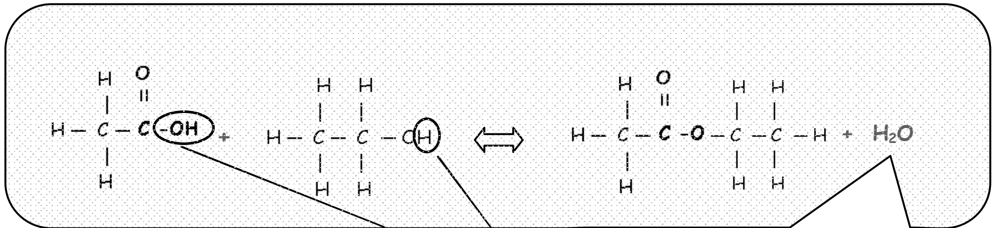
OXIDATION OF ETHANOL USING OXIDISING AGENT

The orange acidified potassium dichromate solution turns green in this reaction.

CHEMICAL REACTIO OF CARBOXYLIC ACIDS

Carboxylic acids react with alcohols to form water in a reaction called **esterification**. Concentrated sulphuric acid (H_2SO_4) is used as a catalyst.

[EXAMPLE] ethanoic acid + ethanol ethyl ⇔ ethanoate + water



ETHANOIC ACID loses the –OH group while ETHANOL loses the –H group to form WATER.

The organic compound formed in esterification is called **ESTER**
Esters also form a homologous series.

Esterification is not the same as neutralization even though water is produced in both reactions. In neutralization, the hydrogen ion reacts with the hydroxide ion to form water. In esterification, an alcohol reacts with a carboxylic acid to form water.

ESTERS *General formula:* $C_nH_{2n+1}COOC_mH_{2m+1}$

Esters are organic compounds formed by the reaction of a carboxylic acid and alcohol.
Esters are **volatile fragrant** substance

Functional group of the esters

-COO-

or

$$-\overset{\overset{O}{\|}}{C}-O-$$

USE OF ESTERS

- **Flavourings in food**
- Ingredients **in Perfume** (*sweet smelling*)

DETERMINATION of NEMES and CHENICAL FORMULAE

The names and formulae of the esters formed follows;

Reaction:	*ORGANIC ACID*	+	*ALCOHOL*	→	*ESTER*	+	*WATER*
Name:	*[ ] anoic acid*	+	*{ } anol*	→	*{ } yl [ ] anoate*	+	*water*
Formula:	*[ ] COOH*	+	*{ } OH*	→	*[ ] COO { }*	+	H_2O

[EXAMPLES]

I) **eth**anoic acid + **eth**anol → **eth**yl **eth**anoate + water
$CH_3COOH + \underline{C_2H_5}OH \rightarrow CH_3COO\underline{C_2H_5} + H_2O$

II) **prop**anoic acid + **but**anol → **but**yl **prop**anoate + water
$C_2H_5COOH + \underline{C_4H_9}OH \rightarrow C_2H_5COO\underline{C_4H_9} + H_2O$

III) **prop**anoic acid + **meth**anol → **meth**yl **prop**anoate + water
$C_2H_5COOH + \underline{CH_3}OH \rightarrow C_2H_5COO\underline{CH_3} + H_2O$

SOME ORGANIC REACTIONS

alkane + O_2 → CO_2 + H_2O

alkene + O_2 → CO_2 + H_2O
alkene + H_2 → *alkane*
alkene + H_2O → *alcohol*

alcohol + *carboxylic acid* → *ester* + H_2O
COMBUSTION: alcohol + O_2 → CO_2 + H_2O
OXIDATION: alcohol + $2[O]$ → *carboxylic acid* + H_2O

10.4 Polymers

Most organic compounds have at most a few tens of atoms. Some have as many as one million atoms. Large molecules can natural materials like proteins, DNA, cellulose or starch, or Man-made materials like plastics. Such large molecules have properties which depend on the functional groups they contain and the overall shape of the molecule itself.

POLYMERS are **very large molecules** that are formed when **thousands** of **smaller units** of identical molecules called **MONOMERS** are joined together.

[redacted]

Polymers are also called MACROMOLECULES

- The process of joining monomers to form a polymer is called **POLYMERISATION**.

Bonding that takes place between monomers are **covalent bonds**

There are 2 types of polymerisations;

1 **ADDITION polymerisation** – the addition reaction between monomers
2 **CONDENSATION polymerisation** – the successive linking together of monomers

→ Polymers occur NATURALLY or may be SYNTHETIC (man-made) .
Hence there are 2 groups of polymers.

1 **SYNTHETIC polymers**
2 **NATURAL polymers**

It is often called **PLASTIC**

[redacted]

SYNTHETIC POLYMERS

Synthetic polymers are substances such as plastics, and man-made fibers such as nylon and terylene. They are formed by either addition or condensation polymerisation.

USES OF SYNTHETIC POLYMERS

- **Poly(ethene):** plastic bags, mineral water bottles, cling film
- **Nylon**: can be made into fibres to make strong ropes(e.g. fishing lines) or woven into cloth to make sleeping bags, parachutes, etc.
- **Terylene**: can be made into fibres and woven into cloth

POLLUTION PROBLEMS OF SYNTHETIC POLYMERS

- Plastics burn easily and may produce **poisonous gases** on combustion.

They need to be coated with fire retardants to reduce the risk of fire

If you look at it as a bright side, plastics are resistant to corrosion

- They are also **non-biodegradable**, i.e. they are **not decomposed** by bacteria in the ground.

Disposal of plastics is difficult and gives rise to environmental pollution when they are incinerated or buried in landfills.

NATURAL POLYMERS

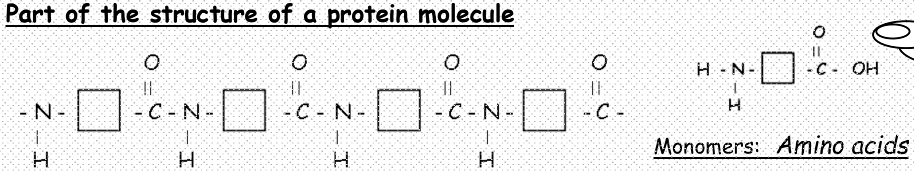
Proteins, carbohydrates and fats are natural polymers found in plants and animals as well as in our food.

All natural polymers are **biodegradable**, unlike most synthetic polymers

PROTEINS

Proteins are needed by both plants and animals mainly for growth, and also to provide enzymes and some energy. Proteins are made by polymerizing amino acids.

Part of the structure of a protein molecule



Monomers: *Amino acids*

Amino acids are the monomer units of all proteins and contain the **–COOH** and **–NH$_2$** groups at either end of the molecule.
[e.g.] H_2N-CH(COOH)H glycine
H_2N-CH(COOH)CH_3 alanine

FATS

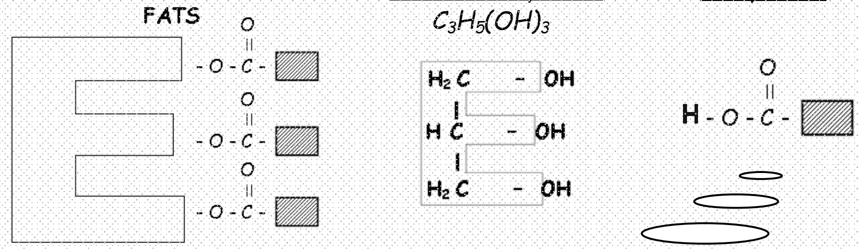
Fats are naturally occurring esters of a fatty acid and glycerol.

They are solid at body temperature and are widely **used by plants and animals** as a means of storing food which can be **used as a fuel(energy).**

the structure of fats

Monomers: *Glycerol* *Fatty acids*

$C_3H_5(OH)_3$

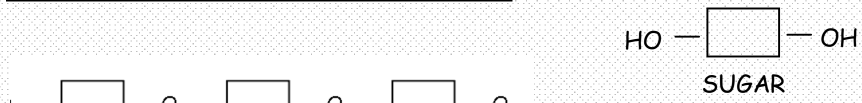


Fatty acids are general names for hydrocarbons which contain the **–COOH. Carboxylic acids** belong to this.

CARBOHYDRATES

Carbohydrates are important nutrients for energy to plants and animals. Examples of carbohydrates are starch, sugar, glycogen and cellulose. Carbohydrates are made from small sugar molecules joined together.

Part of the structure of a Starch molecule



Monomers: *Sugar*

Chemical formula of sugar is
$C_6H_{12}O_6$
General formula of carbohydrates is
$C_x(H_2O)_y$

HYDROLYSIS OF NATURAL POLYMERS

The condensation polymerisation reaction in natural polymers is easily reversible by the process of hydrolysis

For example, if starch hydrolysis takes place, starch (polymer) will break down into smaller molecules, eventually into sugars (monomers).

Hydrolysis is the chemical reaction of a compound with water which causes it to break down.

HYDROLYSIS
Polymer to Monomer

- <u>STARCH</u> - - - - - -▸ SUGARS
- <u>PROTEINS</u> - - - - - -▸ AMINO ACIDS
- <u>FATS</u> - - - - - -▸ FATTY ACIDS

SOAP is formed **by boiling fats** with aqueous sodium hydroxide.
This is because the fatty acids react with sodium hydroxide to form their sodium salts.

- These salts have a WATER-LOVING END and a WATER-HATING END. The water-hating end attaches itself to the grease and the water-loving end causes the grease to detach i.e. dissolve therefore 'cleaning'

<u>STARCH HYDROLYSIS</u>

STARCH is broken down to form SUGAR by:

- ACID HYDROLYSIS (heated with dilute acid)
  Acid hydrolysis is slow but eventually the starch is broken down into glucose, which is the monomer and will not undergo further hydrolysis.

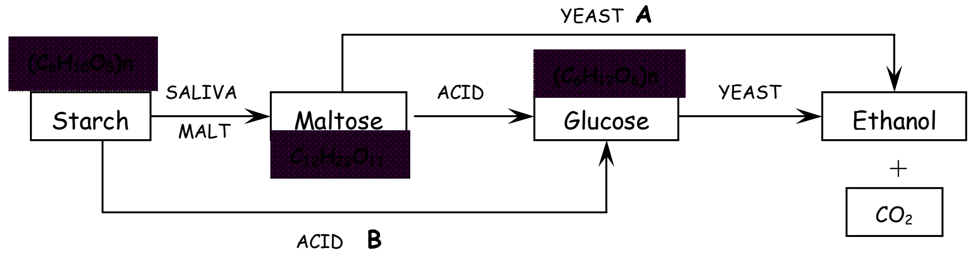
  It takes place in the stomach of mammals.

- ENZYME HYDROLYSIS (by enzyme amylase)
  Enzyme breaks down the starch down into the disaccharide maltose which contains two glucose units minus a water molecule.

  Amylase is found in saliva in the mouth.

[EXERCISE]

The diagram below gives a summary of the breakdown of starch to maltose and glucose and then to ethanol. (From 2003 national exam.)
(a) Name the processes represented by the letters A and B.
(b) What is the purpose of the yeast in process A?



<u>ANSWER</u>

(a) A: Fermentation B: Hydrolysis
(b) For catalyst / For the reaction to speed up

<u>FORMATION OF POLYMERS</u>

As we have seen, we can classify polymers into two groups in terms of how to form.

ADDITION POLYMERS

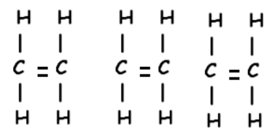
Addition polymers are **synthetic polymers** made from **unsaturated monomers** (e.g. alkenes) through an **addition reaction.** In addition polymerisation, monomers add onto one another to form a single polymer
.

[EXAMPLE] <u>Formation of poly (ethene)</u>

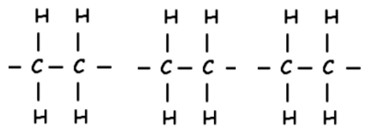
Polyethene is made from ethene molecules. The molecules contain a **carbon-carbon double bond (-*C*=*C*-)** that can add onto one another.

The steps below show how to draw the structure of poly(ethene).

<u>Step 1</u> Draw some ethene molecules side by side:

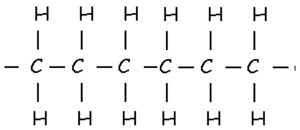


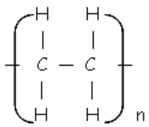
<u>Step 2</u> open the double bonds in the molecules:



This is an **ADDITION REACTION**. To join monomers togeter they must have either ***C* = *C* double bonds** or reactive functional groups that will link them together on the left as well as on the right to form a chain strucure.
Hence monomers should be **UNSATURATED**

Step 3 join the molecules together. The structure can be represented simply as shown on the right.

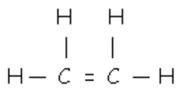
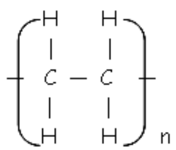
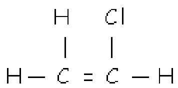
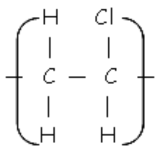
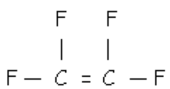
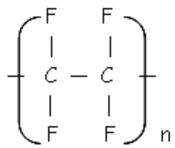




Where 'n' stands for the number of monomers in the structure

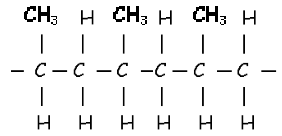
The group $-CH_2CH_2-$ in the simplified structure is called the **REPEATING UNIT** of the polymer.

- Table shows monomers, polymers and their uses.

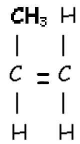
MONOMER	POLYMER	USES
 Ethene	 Poly(ethene)	• Plastic bags • Cling film • Waterproof sheets • Plastic plates
 Chloroethene	 Poly(chloroethene) or polyvinylchloride (PVC)	• Water pipes • Waterproof sheets • Electrical insulators
 Tetrafluoroethene	 Poly(tetrafluoroethene) (PTFE or teflon)	• Coating for non-stick cooking utensils • Sealing and Bearings

[EXERCISE]

The structure of a polymer is shown.
From what hydrocarbon is the polymer made?
Draw its structure



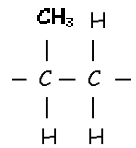
ANSWER



The polymer is an addition polymer.
The monomer from which it is made must contain a carbon – carbon double bond.

To determine the structure of the monomer first identify the repeating unit in the polymer. The monomer is obtained by 'closing the ends' of the repeating unit to obtain the double bond.

The repeating unit is →



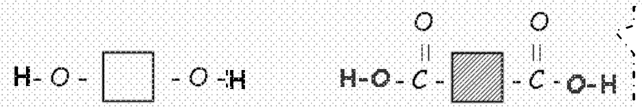
CONDENSATION POLYMERS

Condensation polymers are made from **monomers** containing **alcohol, carboxylic acid** or **amino** functional groups which **link** together. Condensation polymers can be **natural or synthetic**

In condensation polymerization, monomers join together to form a polymer with the **elimination** of small molecules such as water or ammonia.

[EXAMPLE 1] Formation of terylene

Two different monomers join together to form Terylene.
The diol and the dicarboxylic acid can react as monomers to form an ester linkage.

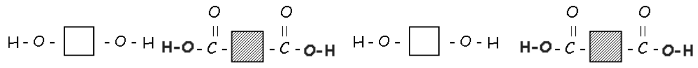


A condensation polymer may contain two kinds of monomer.

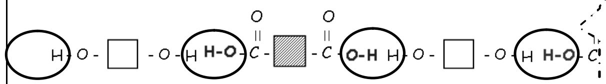
Diol is an alcohol containing 2 -OH groups in its molecules

Dicarboxylic acid is an carboxylic acid containing 2 -COOH groups in its molecules

Step1 draw the monomers alternately



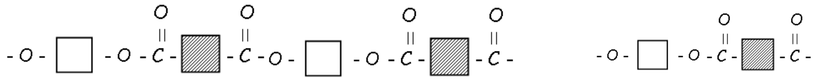
Step2 take away one water molecule from each pair of monomer molecules.



The molecules are said to **condense** together to give water.

Step3 join the remaining parts of the monomers together.

The repeating unit of terylene

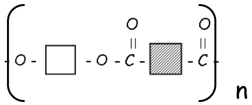


➡ The units in Terylene are joined by the  group of atoms.

We say that terylene has an **ESTER LINKAGE**.

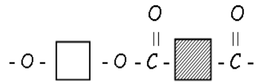
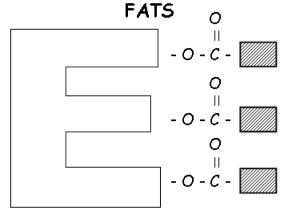
- Polymers containing ester linkages are also known as **POLYESTERS**

Hence, the structure of terylene can be represented as



Natural polymers can be formed by ester linkages as well as synthetic polymers.

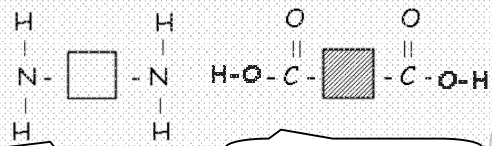
-Table below shows example polymers which have an ester linkage

POLYMER LINK	SYNTHETIC POLYMERS	NATURAL POLYMERS
	POLYESTERS	
ESTER 	**TERYLENE**  ALCOHOL ACID *Monomers; Alcohol +Acid*	**FATS**  GLYCEROL FATTY ACID *Monomers; Glycerol + Fatty acid*

[EXAMPLE 2] Formation of Nylon

Two different monomers join together to form nylon.

The acid and the amine ends on the monomers can react to form an amide linkage.



Diamine is an amine containing 2 $-NH_2$ groups in its molecule

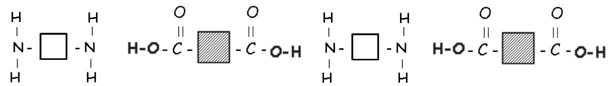
Dicarboxylic acid

AMINE form a homologous group which has $-NH_2$ as its functional group.

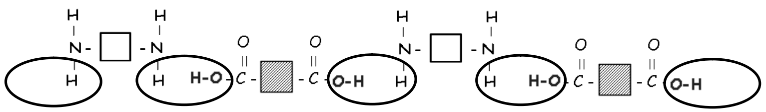
- It is derived from ammonia by replacement of one or more H atoms by hydrocarbons.

[E.g.] Aniline $C_6H_5NH_2$

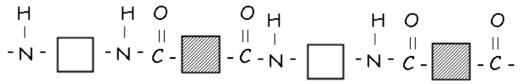
Step1 Draw the monomers alternately



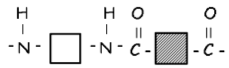
Step2 Take away one water molecule from each pair of monomer molecules in order for molecules to condense together to give water.

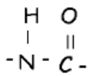


Step3 join the remaining parts of the monomers together.



The repeating unit of nylon

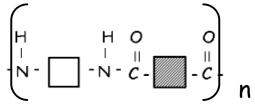


➡ The units in nylon are joined by the  group of atoms.

We say that nylon has an **AMIDE LINKAGE**.

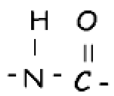
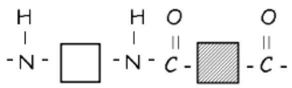
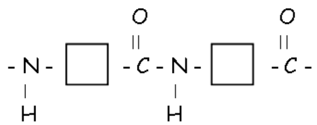
- Polymers containing amide linkages are also known as **POLYAMIDES**

Hence, the structure of nylon can be represented as



Natural polymers can be formed by amide linkages as well as synthetic polymers.

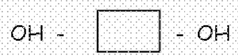
-Table below shows example polymers which have an amide linkage

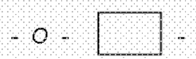
POLYMER LINK	SYNTHETIC POLYMERS	NATURAL POLYMERS
	POLYAMIDES	
AMIDE 	NYLON  <u>*Monomers; Amine +Acid*</u>	PROTEIN  <u>*Monomers; Amino acid*</u>

[EXAMPLE 3] <u>CARBOHYDRATES</u>

As you have seen, carbohydrates (starch and cellulose) are **natural condensation polymers** made up of **smaller sugar** molecules **joined together**

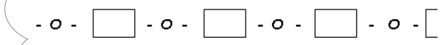
Monomer → *Repeating unit*



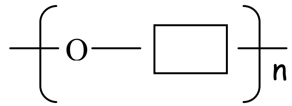


Hence, the structure of carbohydrate can be represented as

The polymer of Carbohydrates is like



It has single monomers.

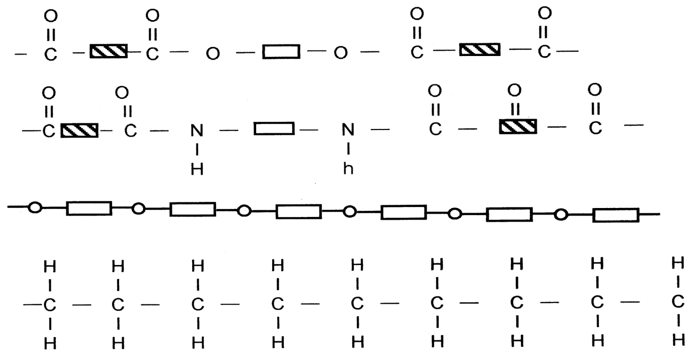


➡ The units in carbohydrates are joined by the - O -

- We say that carbohydrates have **CELLULOSE LINKAGES**.

[EXERCISE 1]

Which of the following structures represents Terylene?
(From 2004 National Exam.)



ANSWER: ***A***

→Terylene is a condensation polymer.

Therefore, the key to work out is to identify the linkages of the polymers.

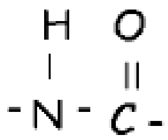
Terylene has ESRTER LINKAGES -COO.

As you can find between monomers, Figure A has ester linkages.

[EXERCISE 2]

Which pair of substances both contain the linkage shown?

A. nylon and terylene
B. sugars and protein
C. nylon and protein
D. terylene and poly(ethene)



<u>*ANSWER:*</u> ***C***

→This is AMIDE LINKAGE. Hence polymers to be answered should be polyamides
Nylon is a synthetic polyamide, while protein is a natural polyamide.

<u>KILONA LONA</u> has finished. But your learning still continues.
As long as you're studying, you're making a progress.

REVISION QUESTIONS

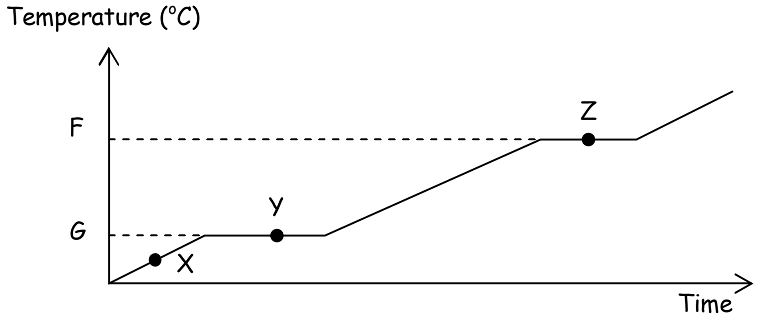
1. Matter

1. The figure shows the changes of state of matter

$$\text{solid} \underset{A}{\overset{\text{melting}}{\rightleftarrows}} \text{Liquid} \underset{\text{condensation}}{\overset{B}{\rightleftarrows}} \text{gas}$$

Name the process represented by A and B.

2. Study the heating curve and answer the questions below.



(a) Name the temperature G and F.
(b) At what state or states can a substance exists at the point X, Y and Z in graph.

3. Study the table below. This table shows the melting points and boiling points of oxygen, iron, diamond and sulphur.

Substance	Melting point (°C)	Boiling Point (°C)
Oxygen	-219	-183
Iron	1540	2900
Diamond	3550	4832
Sulphur	119	445

Which substance in the table is…
(a) a liquid at 200 °C
(b) a gas at 0 °C
(c) a solid at 1600 °C

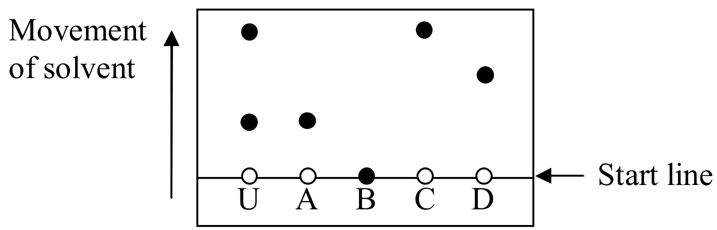
4.The table below shows the boiling points of the elements found in a sample of liquid air.

element	argon	helium	neon	nitrogen	oxygen
boiling point/°C	-186	-269	-246	-196	-183

Which elements would be gaseous at -190°C?

2. Experimental Techniques

1. U is soluble mixture. A, B, C and D are known pure substances. A pupil carried out a chromatography experiment to separate the mixtures. Below is the chromatogram showing the result of the separation.



(a) Which of the dyes are present in U?
(b) Suggest which of the dye is insoluble in this solvent.

2. What is the name of substance
(a) which remains on the filter paper after filtration.
(b) which is able to pass through the filter.

3. Name one method to separate the following mixtures: crude oil, liquid air.

4. Name a method to separate a mjixture of

(a) water and cooking oil
(b) a mixture of iodine and sodium chloride

5. A liquid is thought to be pure ethanol. What is the best way to test its purity?
   A. Test with universal indicator paper
   B. Burn its completely in oxygen
   C. Mesure its boiling point.
   D. React with aqueous sodium hydroxide

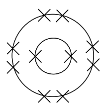
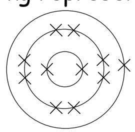
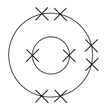
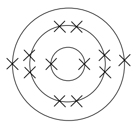
3. Atoms, elements and compounds

1, Which of the following atomic particles have almost the same mass as a neutron.
A Proton B Electron C Sodium ion D Alpha particle

2. An element x forms an ion ${}^{45}_{21}X^{3+}$ How many protons, electrons and neutrons are there in the ion of X?

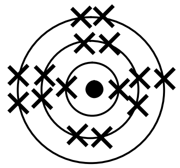
	Protons	Electrons	Neutrons
A	21	21	24
B	21	18	45
C	21	18	24
D	18	21	24

3. Which of the following represents a magunisium ion?

A  B  C  D 

4. The proton number of Chlorine is 17, and Sodium is 11.
(a) Draw the electronic structure of the elements Chlorine and Sodium.
(b) Write down the formula of a compound formed when Sodium combines with chlorine.
(c) Draw the structure of compound (b).

5. The diagram below represents electron (x) arrangement of a particular atom. Study this diagram and answer the questions that follow



The relative atomic mass of the atom represented is 31
(a) What is its proton number?
(b) What is its neutron number?

6. Complete the table below for these two isotopes of chlorine.

Chlorine	Mass number	Number of protons	Number of neutrons
${}^{35}Cl$	35	17	
${}^{37}Cl$			20

4. Stoichiometory

1. Calcium and nitrate ions combine to form calcium nitrate. The formula of the compound formed is…
A $CaNO_3$ B Ca_2NO_3 C $Ca(NO_3)_2$ D $CaNO_6$

2. Sodium phosphate has a formula Na_3PO_4. Then the total number of atoms in the formula of iron (II) phosphate is …
A 6 B 8 C 13 D 17

3. Consider the reaction
$CaCO_3 \rightarrow CaO + CO_2$
(a) What mass of Lime (CaO) would be produced from 20 tonnes of lime stone ($CaCO_3$)
(b) What mass of carbon dioxide will be given off by heating 20 g of calciumcarbonate?
4. Magnesium reacts with dilute hydrochloric acid to form magnesium chloride and hydrogen gas.
(a) Write a balanced chemical equation, including state symbols, for the reaction of magnesium with dilute hydrochloric acid.
(b) Calculate the mass of magnesium chloride formed when 6.0g of magnesium react with an excess of dilute hydrochloric acid.

5. Balanced equations below.

(a) Na_2CO_3 + HCl → $NaCl$ + H_2O + CO_2

(b) H_2S + O_2 → H_2O + SO_2

5. Periodic Table

1 Complete the sentences.

(a) In the periodic table, elements are arranged in the order of the _____________.

(b) Elements in Group I are called ____________________.

(c) Elements Group I have ____________ boiling points and melting points, and low ___________. compare with other metals.

(d) An atom of halogen has _______________ valency electrons. So it forms the ion which has one ___________ charge.

(e) Elements in Group 0 are called ______________.

2. Complete the table below.

Element	Chemical formula	State at r.t.p
Chlorine		
Bromine	Br_2	
Iodine		Solid

3. Which one displaces bromine from aqueous sodium bromide?

A Chlorine B Iodine C Sodium D Sodium chloride

6. Chemical Reaction

1. State four factors to speed up the reaction

2. Name the substance which reduced in each reaction below.

(a) $CuO\ (s) + H_2\ (g) \longrightarrow Cu\ (s) + H_2O\ (g)$

(b) $SO2\ (g) + 2Mg\ (s) \longrightarrow S\ (s) + 2MgO\ (s)$

3. Explain an (a) exothermic reaction and (b) endothermic reaction

4. The rate of the reaction between a magnesium ribbon and an excess of dilute hydrochloric acid could be measured. .

The volume of hydrogen produced was recorded every minute as shown in the table below.

Time (min)	0	1	2	3	4	5	6	7
Volume of Hydrogen (cm^3)	0	14	23	31	38	40	40	40

(a)Plot the results on a graph paper and draw the graph.

(b) What was the total volume of hydrogen produced when the reaction was over?

(c) Why did the reaction stop?

(d) How could you make the reaction go faster?

7. Acids, Bases and Salts

1. (a) Name the ion which all acids solution contain.

(b) Name the ion which all alkali solution contain.

2. What is the difference between strong acid and weak acid?

3. Write a chemical equation for the reaction between hydrochloric acid and aqueous sodium hydroxide.

4. Which one of the following equation represents neutralization?

A $Zn\ (s) + CuSO_4\ (aq) \longrightarrow ZnSO_4(aq) + Cu\ (s)$

B $Zn\ (s) + H_2SO_4\ (aq) \longrightarrow ZnSO_4(aq) + H_2\ (g)$

C $ZnO\ (s) + H_2SO_4\ (aq) \longrightarrow ZnSO_4(aq) + H_2O\ (l)$

5. There are some oxides, CO_2, ZnO, CuO, SO_2, Al_2O_3, CO, MgO

(a) Define amphoteric oxide.

(b) Which oxides are basic oxides?

(c) Which oxides are amphoteric oxides?

(d) Which oxides are neutral acids?

(e) Which oxide cause acid rain?

6. There are salts: $NaCl$, $BaSO_4$, $CuSO_4$, $AgCl$, K_2CO_3, $CaCO_3$

(a) Which salts are soluble in water?

(b) Which salts can be obtained as precipitates?

8. Metals

1. Give two advantage of alloy.
2. What is use of stainless?

3. In extraction on iron, state three materials which are added to the top of the furnace with the heated ore.

4. Name the ore and chemical formula thet is used in the extraction of iron.

5. What is the function of carbon monoxide in extraction of iron?

9. Non-Metals

1. Name four common gases in the clean air and state the composition each.

2. What method is used for manufacture of oxygen from air?

3. Name the reactants and products, and give the condition in Haber process.

4. Write the balanced chemical equation of Haber process.

10. Organic Chemistry

1 Which of the following compounds **cannot** react with hydrogen?
A C_4H_8 B C_5H_{10} C C_6H_{12} D C_7H_{16}

2. Explain isomers using butane as an example.

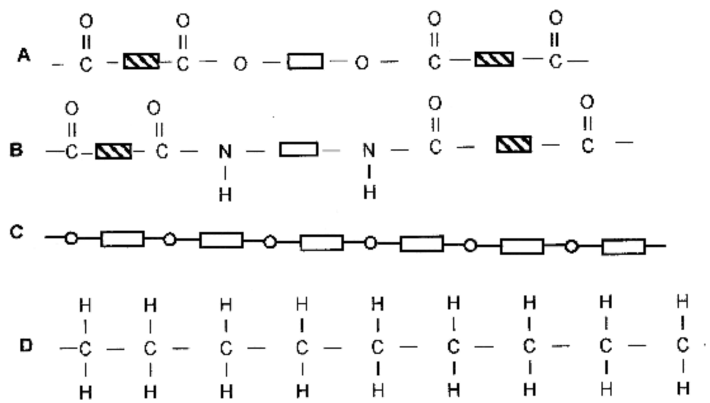
3. Which substance is the main component of natural gas?

4. How crude oil could be refined?

5. What does the 'unsaturated' mean?

6. Write the equation for the reaction of ethene and steam with their structural formulae.

7. Which of the following structures represents nylon?



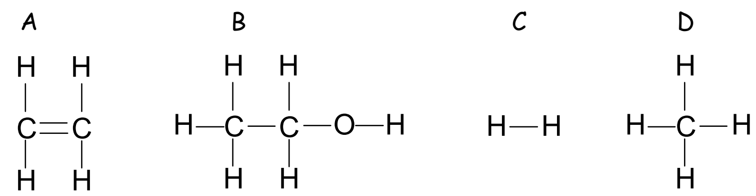
8. Which substance below belongs to homologous series of the general formula C_nH2_n?
A. Ethane B. Ethanoic acid C. Ethene D. Ethanol

9. Which substance below belongs to a homologous series of general molecular formula C_nH_{2n+2}
A. CH_3COOH B. C_2H_5OH C. C_3H_8 D. C_2H_4

10. Write a balanced chemical equation for the complete combustion of ethanol.
$C_2H_5OH + O_2 \rightarrow$

11. State the final product for the hydrolysis of carbohydrates.
A. Simple sugers B. Ethanol C. Polymers D. alcohol

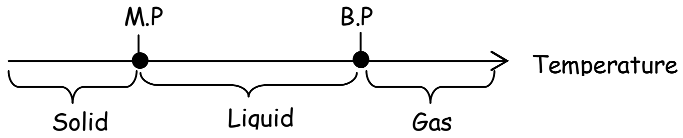
12. Which substance react with steam?



ANSWER

1. Matter

1. A Freezing B Evaporation
2. (a) G: Melting point F: Boiling point
   (b) X: Soild Y: Solid and Liquid Z: Liquid and Gas
3. (a) Sulphur (b) Oxygen (c) Diamond



4. helium, neon and nitrogen
A substance is a gas at a temperature above its boiling point.

2. Experimental Techniques

1. (a) A and C (b) B : Spot B doesn't move because of insoluble in this solvent,
2. (a) residue (b) filterate
3. Fractional distillation
4. (a) a separating funnel (b) sublimation
5. C
When a substance is pure, it has a sharp melting and boiling points.

3. Atoms, elements and compounds

1. A

The mass of a neutron and proton are one (it doesn't have any unit). The mass of an electron is approximately 1/1840. The alpha particle is the nucleus of helium atom.

2. C

The number of protons in an atom of the element is given by the atomic number, which is 21. It is a cation and it has a charge of +3, so the number of electrons is the number of protons – 3, that is 21 – 3 = 18. The number of neutrons is equal to the mass number – the atomic number, which is 45 – 21 = 24.

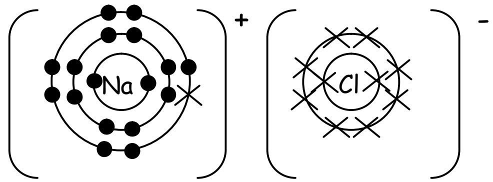
3, A
The atomic number of magnesium is 12 and it loses two electrons to become full outer-shell, so there must be 10 electrons in a sodium ion

4. (a)

(b) NaCl
The valencies of sodium and chlorine are 1 and 1. The formula is given by exchanging their valencies as shown below,

$Na^1 \; Cl^1 \rightarrow NaCl$

(c)



5. (a) 15
   Number of Electrons = Number of Prtons
   (b) 16
   Number of Neutrons = Mass Number – Number of protons (31-16=16)

6.

Chlorine	Mass number	Number of protons	Number of neutrons
^{35}Cl	35	17	18
^{37}Cl	37	17	20

- Number of **Neutrons** = **Mass Number – Atomic Number**
- Number of **Electrons** = Number of Protons = **Atomic Number**

(So that **Atoms** should be **electrically neutral)**

4. Stoichiometory

1. C

The valencies of calcium and nitrate are 2 and 1. The formula is given by exchanging their valencies as shown below,

$Ca^{2} NO_3^{1} \rightarrow Ca(NO_3)_2$

2. C

Sodium ion has valency of 1, the valency of phosphate ion must be equal to three times the valency of sodium ion, which is 3 x 1 or 3. The valency of iron (II) ion is 2. Therefore the formula of iron (II) phosphate is $Fe_3(PO_4)_2$. One PO_4 molecule is composed of one P atom and four O atoms, which is in total five atoms. In $Fe_3(PO_4)_2$, there are three Fe atoms and two PO_4 molecules, which are 3 + 2 x 5 = 13 atoms.

3. Relative atomic mass of Ca: 40, C: 12 and O: 16, so relative molecular mass of $CaCO_3$: 100 (40+12+3x16) and CaO: 56 (40+16).
So,
(a)

	$CaCO_3$	$\rightarrow$	$CaO + CO_2$
Relative mass	100		56
Actual mass	20 tonnes		X tonnes

100X = 56 x 20
<u>X = 11.2 tonne</u>

(b)

	$CaCO_3$	$\rightarrow$	$CaO + CO_2$
Relative mass	100		44
Actual mass	20 g		X g

100X = 44 x 20
<u>X = 8.8 g</u>

4 (a) $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
(b) Relative atomic mass of Mg:24, H:1 and Cl:35.5, so relative molecular mass of $MgCl_2$: 95 (24+2x35.5).

	$Mg(s) + 2HCl(aq)$	$\rightarrow$	$MgCl_2(aq) + H_2(g)$
Relative mass	24		95
Actual mass	6.0 g		Xg

24X = 95 x 6.0
<u>X = 23.75 g</u>

5.

(a) $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$

No. of Na	2	2x1=2
C	1	1
O	3	3
H	2x1=2	2
Cl	2x1=2	2x1=2

(b) $2H_2S + 3O_2 \rightarrow 2H2O + 2SO_2$

No. of H	2x2=4	2x2=4
S	2x1=2	2x1=2
O	3x2=6	2x1+2x2=6

5. Periodic Table

1. (a) atomic number (b) alkali metals (c) low, density (d) seven, negative (e) noble gases

2.

Element	Chemical formula	State at r.t.p
Chlorine	Cl_2	Gas
Bromine	Br_2	Liquid
Iodine	I_2	Solid

3. A
This is the question about displacement reaction.

More reactive halogens will DISPLACE less reactive halogens from their aqueous salt solutions

F Cl Br I

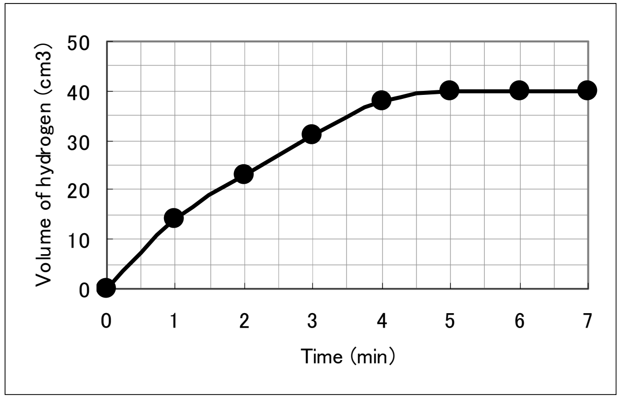
More reactive $\longleftrightarrow$ less reactive

A $Cl_2 + 2NaBr \longrightarrow 2NaCl + Br_2$

B $I_2 + 2NaBr \longrightarrow$ X

6. Chemical Reaction

1. Consentration, Pressure, Surface area of solid, Temperature, Catalyst (You can choose four factors)
2. (a) CuO (CuO is said to be reduced to Cu because it has lost oxygen)
   (b) SO_2 (SO_2 is said to be reduced to S because it has lost oxygen)
3. (a) A reaction that given out heat to the surrounding.
   (b) A reaction that taken in heat from the surrounding.
4.
   (a)



(b) $40cm^3$
(c) Because all magnesium has been used up.
(d) Use more concentrated hydrochloric acid or increase the temperature of substances.

7. Acids, Bases and Salts

1. (a) Hydrogen ion H^+
   (b) Hydroxide ion OH^-
2. Strong acid ionize completely, weak acid ionize partially. In weak acid, ionization is reversible
3. $HCl\,(aq) + NaOH\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l)$
4. C
   Acids react with bases to give salts and water only. This reaction is called neutlisation

ZnO (s)	+ H_2SO_4 (aq)	→ $ZnSO_4$ (aq)	+ H_2O (l)
Base	Acid	Salt	Water

5. (a) The oxides which react with both acids and bases.
   (b) CuO, MgO (c) ZnO, Al_2O_3 (d) CO (e) SO_2

6. (a) NaCl, $CuSO_4$, K_2CO_3
   (b) $BaSO_4$, AgCl, $CaCO_3$

8. Metals

1. Stronger and harder than the pure metals, resistance of corrosion
2. Car bodies, Cutlery
3. Lime stone, Cokes and haematite (iron ore)
4. Haematite, Fe_2O_3
5. Reducing agent (reduce oxygen from iron ore)

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

Reduced

9. Non-Metals

1. Nitrogen 78%, Oxygen 21%, Carbon dioxide 0.04%, Noble gases 0.94%
2. Fractional distilation
3. Reactants: Hydrogen H_2, Nitrogen N_2
   Products: Ammonia NH_3
   Pressure: 350 atm, Temperature: 450°C, Iron is put as catalyst
4. $3H_2 + N_2 \rightleftarrows 2NH_3$

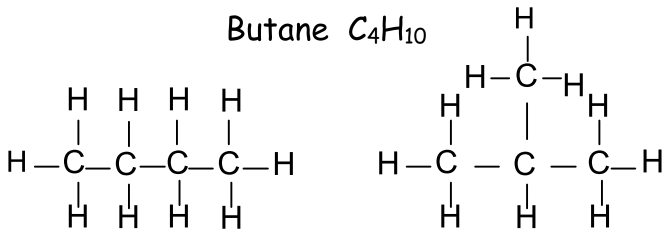
10. Organic Chemistry

1 D

C_7H_{16} has formula of C_nH_{2n+2}, and it is the general formula of alkanes. It does not undergo hydrogenation because alkanes are all saturated hydrocarbons. Other compounds have formula of C_nH_{2n}, and it is the general formula of alkenes. They can undergo hydrogenation because alkenes are unsaturated hydrocarbons.

2.

Butane C_4H_{10}



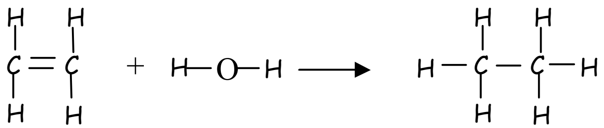
Isomers are different compounds which have the same molecular formula but different structural formula.

3. Methane CH_4

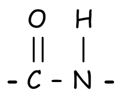
4. Fractional distillation

5. Compounds which have molecules containing double or triple covalent bonds are said to be unsaturated.

6.



7. B

The unit in nylon are joined by the  group of atoms

8. C

Substance	Formula
Ethane	C_2H_6
Ethanoic acid	CH_3COOH
Ethene	C_2H_4
Ethanol	C_2H_5OH

9. C

C_nH_{2n+2} is C_3H_8 when n = 3

10 $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

11. A

12.A

C_2H_4 (g) + H_2O $\rightarrow$ C_2H_5OH (g)

This is an addition reaction. In this reaction, molecules are always added across a carbon = carbon double bond (C=C).

EXTRA. HOW TO ANSWER EXAM QUESTIONS

- **READ** the **QUESTION** carefully
  - ☑ look out for **KEY WORDS** for **CHEMISTRY**
  - ☑ collect any **DATA** relevant for the question

- **THINK** how you can use what you **KNOW** to ANSWER the question
  - ☑ Use the **PERIODIC TABLE** if necessary

- **ANSWER** what the question **ASKS FOR**
  - ☑ Look out for **KEY WORDS** to answer the **EXAM QUESTION** e.g. *state, define, describe, calculate, draw*
  - ☑ Use the number of MARKS as a guide to HOW DETAILED your answer should be:
    EACH MARK should be **ANOTHER POINT** for your answer!!
  - ☑ ALWAYS SHOW any **WORKING OUT**

Common Exam Terminology
"*State*": one or two words
"*Define*": one sentence
"*Describe*" / "*Discuss*" / "*Explain*": a few sentences
"*Compare*" / "*Contrast*": differences / similarities between
"*Draw*": diagram
"*Calculate*": calculations with working out

[EXAMPLE 1]
(a) State the name of your Chemistry teacher [*1 mark*]
ANSWER: MR. T. IGUCHI
(b) Define the role of a Chemistry teacher [*1 mark*]
ANSWER: To teach chemistry to students
(c) Describe your Chemistry teacher in terms of physical appearance
THINK: 5 marks = 5 points!! [*5 marks*]

- *Average height*
- *Short hair*
- *Wears glasses*
- *Mkuwa (Mzungu)*
- *Male*

NOT (as need physical appearance):

- *Punctual*
- *Friendly*
- *Strict*
- *Too lenient*

Now write these 5 points into short sentences →

ANSWER:
My chemistry teacher is a male of average height. He has black short hair and sometimes wears glasses. He comes from Japan.

[EXAMPLE 2]
(a) State the S.I. unit of matter [*1 mark*]
ANSWER: kilogram

(b) Define diffusion [*1 mark*]
ANSWER: Diffusion is the movement of particles from an area of high concentration to an area of low concentration

(c) Describe in terms of Kinetic Theory the process of melting [*3 marks*]
THINK: melting = solid → liquid {heating}
Kinetic Theory = MOVING particles
→ Answer should include all these details and cover at least 3 points (3 marks!!)
ANSWER:

- *As a solid is heated, the particles begin to vibrate more and more*
- *Eventually the particles vibrate so much that they overcome the attractive forces and begin to move about like particles in a liquid*
  i.e. the solid has melted to become liquid