1.1 | Measurements of Physical Quantities (Experimental Chem)

Measurement of Time

For measurement of time, a stopwatch is used. Common apparatus include digital stopwatch. Depending on the type of digital stopwatch, it can measure time up to 2 decimal places in seconds (e.g: 20.01 s). The SI unit for time is the second (s). Other units such as minute (min) and hour (h) could also be used.

Table 1.1.1. The precisions and uncertainties of digital stopwatches

Digital Stopwatch			
Precision/s	**Uncertainty of measurement/s**	**D.P.**	**Example timing/s**
0.1	± 0.1	1	28.1
0.01	± 0.01	2	28.00 28.11

> ⓘ **Note**
> *Precision* is determined by the smallest division on the instrument whereas *uncertainty* of measurement corresponds to the range where the actual value could fall into. For example, if a stopwatch has an uncertainty of ± 0.01s, if its reading shows 28.00 s, it means the actual value could be anywhere in the range of 27.99 s to 28.01 s.

Measuring of Temperature

For measurement of temperature, a *thermometer* is used — *alcohol thermometers* and *mercury thermometers*. Temperature measured by a thermometer is recorded to the nearest 0.5 °C. The *SI unit* for temperature is the kelvin (K). Other units used include degree Celsius (°C).

Temperature can also be recorded using a *datalogger connected to a temperature sensor*. The advantages of using a datalogger connected to a temperature sensor are:

- It is more accurate than the mercury or alcohol thermometer.
- It can record data continuously over a period of time.
- It saves data (in a computer) which can be used to produce graphs and charts

Table 1.1.2. The precision and uncertainty of laboratory thermometers

Thermometer			
Precision/°C	**Uncertainty of measurement/**°C	**D.P.**	**Example temperature/**°C
1	±0.5	1	23.0 23.5

Measurement of Length
For measurement of length in laboratory, *metre rule* and *measuring tape* are used. The *SI unit* for length is metre (m). Other units include mm, cm, dm.

Table 1.1.3. The precision and uncertainty of a metre rule

Metre rule			
Precision/mm	**Uncertainty of measurement/mm**	**D.P.**	**Example length/mm**
1	±0.5	1	20.5

Measurement of Mass
For measurement of mass, an *electronic balance* is used. Mass measured by an electronic mass balance is recorded to the nearest 0.001 g, depending on the type of electronic mass balance used. The *SI unit* for mass is the kilogram (kg). Mass can also be measured in grams (g) or milligrams (mg).

Table 1.1.4. The precision and uncertainty of an electronic balance

Electronic Balance			
Precision/g	**Uncertainty of measurement/g**	**D.P.**	**Example mass/g**
0.001	±0.001	3	2.253

Measuring of Volume (of liquids and solutions)
For measurement of volume of liquids, pipette, volumetric flask, measuring cylinder and burette can be used. (Refer: *table 1.1.5*)

Table 1.1.5. The precision and uncertainty of various instruments used to measure volume

Apparatus	Precision/ cm^3	Uncertainty of measurement/ cm^3	D.P.	Example volume/ cm^3	Primary use
Measuring Cylinder	1	0.5	1	15.0	Measures approximate volumes of liquids and solutions
Pipette	-	0.1	1	25.0 10.0	Measures accurate, fixed volumes of liquids and solutions
Burette	0.1	0.05	2	25.45	Measures very precise volumes of liquids and solutions
Volumetric Flask	-	-	-	100	Measures accurate fixed volumes that are larger

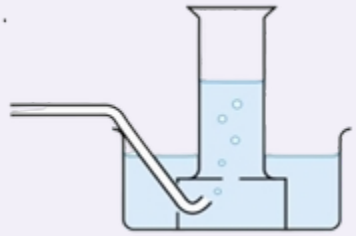
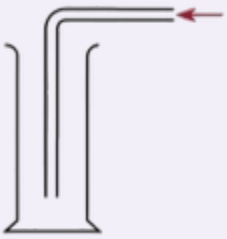
> ⓘ **Note**
> To avoid parallax error, when reading from a burette or measuring cylinder, we should position our eyes at the meniscus, to avoid parallax error.

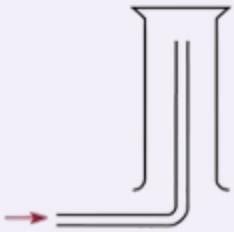
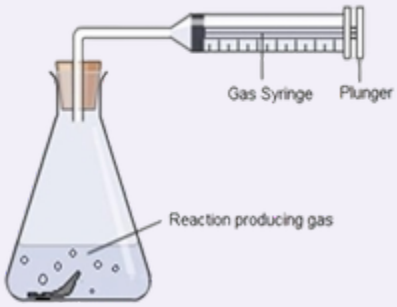
Measuring Volume of Gases

To measure the volume of a gas, a gas syringe is used. Typical gas syringes measures volume up to 100 cm^3.

1.2 | Collection and Drying of Gases (Experimental Chem)

Table 1.2.1. Methods of collecting gases

Water displacement	Used to collect **insoluble or slightly soluble** gases. Density **does not** affect gas collection. **Description** Gas released is bubbled into an inverted test tube filled with water. Gas will take up the volume and push down the water level in the test tube. (Since it does not dissolve in water) **Example Gases** Oxygen, hydrogen, carbon dioxide  **Fig 1.2.1.** Collection of gases using water displacement
Downward delivery	Used to collect gases that are **denser** than air. **Description** The gas collected will sink in the gas jar and displace the air. Gas Jar must not be sealed in order for the air to be pushed out by the gas so that the gas can enter. **Example Gases** Chlorine, hydrogen chloride, sulfur dioxide, nitrogen dioxide  **Fig 1.2.2.** Collection of gas using downward delivery

Upward delivery	Used to collect gases that are **less dense** than air. **Description** The gas collected will rise up the gas jar and displace the air inside. Gas Jar must not be sealed in order for the air to be pushed out by the gas so that the gas can enter. **Example Gases** Ammonia, Hydrogen  **Fig 1.2.3.** Collection of gases using upward delivery
Gas syringe	Used to **measure** and collect any type of gas **Description** Gas produced would be collected in the gas syringe and push the plunger. The gas can be measured directly.  **Fig 1.2.4.** Collection of gases using gas syringe

ⓘ **Note**
Approximate relative mass of air is 29.0. Gases with a higher M_r than air can be collected using downward delivery while gases with a lower M_r than air can be collected using upward delivery.

Table 1.2.2. Densities, solubilities and methods of collection for various gases

Gases (chemical formula)	Density compared to air	Solubility in water	Methods of collection
Hydrogen (H_2)	Less dense	Not soluble	Upward delivery & Water displacement
Ammonia (NH_3)	Less dense	Highly soluble	Upward delivery
Oxygen (O_2)	Slightly denser	Slightly soluble	Downward delivery & Water displacement
Hydrogen chloride (HCl)	Denser	Highly soluble	Downward delivery
Carbon dioxide (CO_2)	Denser	Slightly soluble	Downward delivery & Water displacement
Nitrogen dioxide (NO_2)	Denser	Highly soluble	Downward delivery
Sulfur dioxide (SO_2)	Denser	Highly soluble	Downward delivery
Chlorine (Cl_2)	Denser	Highly soluble	Downward delivery

Methods for Drying Gases
Neutral gases: Oxygen and Nitrogen gas
Acidic gases: Hydrogen Chloride, Carbon Dioxide, Nitrogen Dioxide and Sulfur Dioxide
Alkaline gases: Ammonia

Table 1.2.3. Drying agents used to dry various gases

Concentrated sulfuric acid	Used to dry **<u>neutral/acidic</u>** gases.

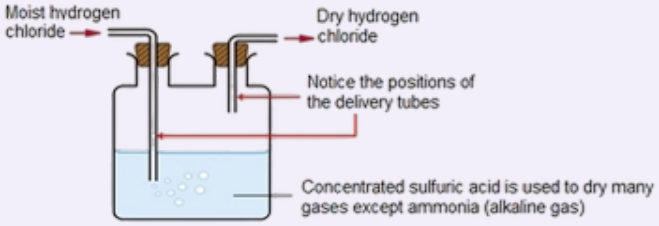
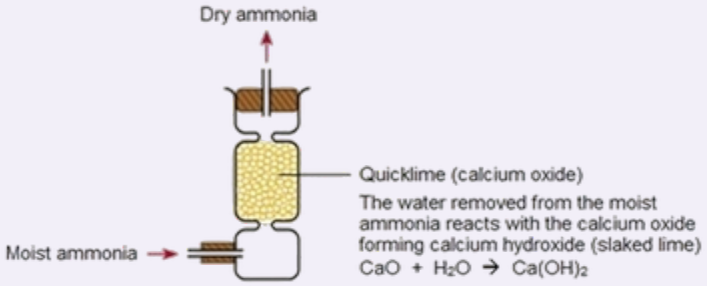
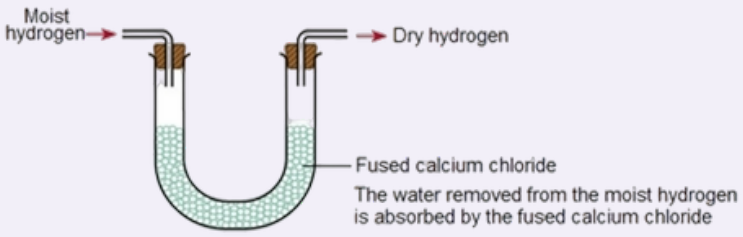


Fig 1.2.5. Setup to dry gases with concentrated sulfuric acid

	ⓘ Note This method is not suitable for alkaline gases which react with sulfuric acid. (Neutralisation) ***Key Qns 1.2.1*** *(Drying H_2)* ***Q. Concentrated sulfuric acid shouldn't be used to dry hydrogen gas, which is neutral. Suggest a reason why.*** When concentrated sulfuric acid absorbs water vapour from gas, a lot of heat is given out. This causes hydrogen gas, which is highly flammable, to catch fire.
Quicklime (calcium oxide)	Used to dry **neutral/alkaline** gases Calcium oxide absorbs moisture and carbon dioxide from the air, so it must be **freshly heated** before use.  **Fig 1.2.6.** Setup to dry gases using calcium oxide ⓘ Note This method is not suitable for acidic gases which react with calcium oxide. (Neutralisation)
Fused calcium chloride (anhydrous calcium chloride)	Used to dry **neutral/acidic** gases Calcium chloride readily absorbs moisture from the air, so it must be **freshly heated** before use.  **Fig 1.2.7.** Setup to dry gases using fused calcium chloride

> ⓘ **Note**
> This method is not suitable to dry ammonia as anhydrous calcium chloride reacts with ammonia.

1.3 | How Are Substances In Mixtures Separated? (Experimental Chem)

Definition 1.3.1 *(Mixture)*

A mixture is a substance that is made up of two or more substances that are not chemically combined.

Definition 1.3.2 *(Pure Substance)*

A pure substance is made up of one single element or compound. It is not mixed with any other substances.

Table 1.3.1. Table showing the methods of separation for several types of mixture. Solid-liquid mixtures can be either miscible or immiscible, and they require different separation techniques

Solid-Solid	Solid-Liquid	Liquid-Liquid	Gaseous-Gaseous
Magnetic attraction	Filtration	Separating funnel	Fractional distillation (after liquefying)
Sieving	Evaporation to dryness	Chromatography	
Using suitable solvents	Crystallisation	Fractional distillation	
Sublimation	Simple distillation		

Separating Solid-Solid Mixtures

1. Magnetic Attraction
A magnet can be used to separate magnetic solids (eg. Iron, Cobalt, Nickel) from non-magnetic solids.

2. Sieving
When a mixture consists of bigger and smaller particles, they can be separated by using a sieve with a suitable pore size. A sieve can be used to separate solids with different particle sizes.

3. Using Suitable Solvents

To separate a mixture of two solids, we use a solvent in which only **one solid is soluble** while the other solid is insoluble. Eg. A mixture of sand and solid sodium chloride can be separated using water as the solvent.

4. Sublimation

Definition 1.3.3 *(Sublimation)*

Sublimation is a process in which a substance changes from a solid to a gas without passing through the liquid state.

The resulting solid deposit formed during sublimation is known as the **sublimate**. Examples of solids that sublime at room temperature and pressure are iodine, ammonium chloride, naphthalene (mothballs) and dry ice (solid carbon dioxide).

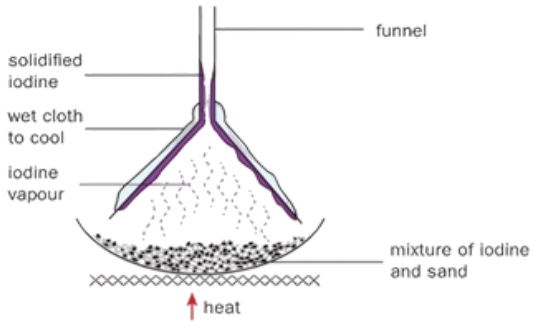


Fig 1.3.1. Setup to separate iodine from sand using sublimation

Separating Solid-Liquid Mixtures

1. Filtration

Filtration is the technique used to separate an insoluble solid from the liquid in a solid-liquid mixture. The liquid that passes through the filter paper is called the **filtrate**. The solid that remains on the filter paper is known as the **residue**.

2. Evaporation to dryness

Evaporation to dryness is used to separate a **dissolved solid** (solute) from its solvent by heating the mixture until all the solvent has vaporised.

> ⓘ Note
> Evaporation to dryness is only suitable for substances where solute **does not decompose** on strong heating. The solid (solute) obtained by evaporation to dryness is **not always pure**.
> When all the water has been removed, any soluble impurities will be left together with the solid.

3. Crystallisation

Crystallisation is used to obtain a pure solid (solute) from its saturated solution. This method is suitable for substances (solute) that **decompose** on strong heating.

In crystallisation, excess water from an aqueous solution is removed by heating the solution. Heating is stopped at the stage when a hot saturated solution is formed. When the hot, resulting solution is allowed to cool to room temperature, the dissolved solid (solute) will be formed as pure Crystals.

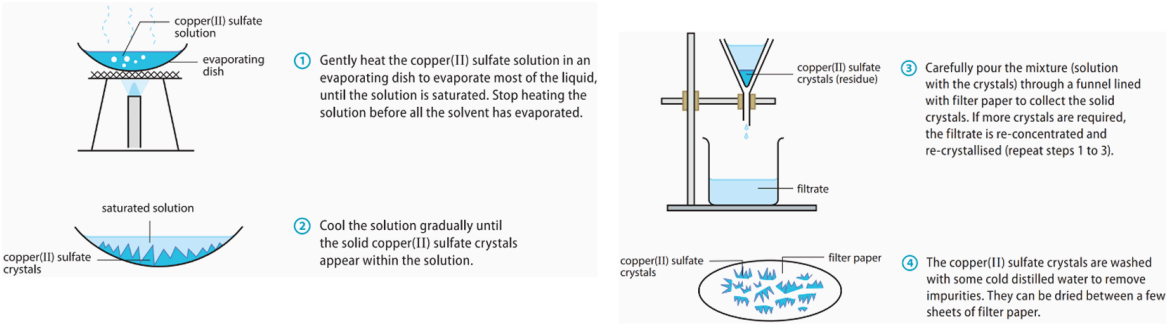


Fig 1.3.2. Diagram shows how crystallisation is used to obtain pure copper(II) sulfate crystals from an aqueous solution of copper(II) sulfate

Definition 1.3.4 *(Saturated Solution)*

A saturated solution is one in which a maximum amount of solute is dissolved in a given volume of solvent.

Proving that a solution is saturated

Dip a clean glass rod into the solution and remove it. If small crystals form on the rod as the solution cools, the solution is at its saturation point.

Advantages of crystallisation vs evaporation to dryness

The substance obtained from evaporation may not be pure whereas crystallisation results in the formation of a pure solid. This is because evaporation to dryness vaporises all the solvent whereas crystallisation removes just enough solvent to obtain a saturated solution.

Key Qns 1.3.1 *(Key Concept of Crystallisation)*

Q. Why doesn't the compound remain dissolved in water (start forming crystals) when the saturated solution is cooled?

The **solubility of the compound decreases** when temperature decreases during cooling, causing the compound to precipitate out as crystals.

Simple Distillation

This method is used to separate a pure **solvent (liquid)** from its solution. It is a process of boiling a liquid and condensing the vapour. The solvent has a much lower boiling point compared to the solute in the solution.

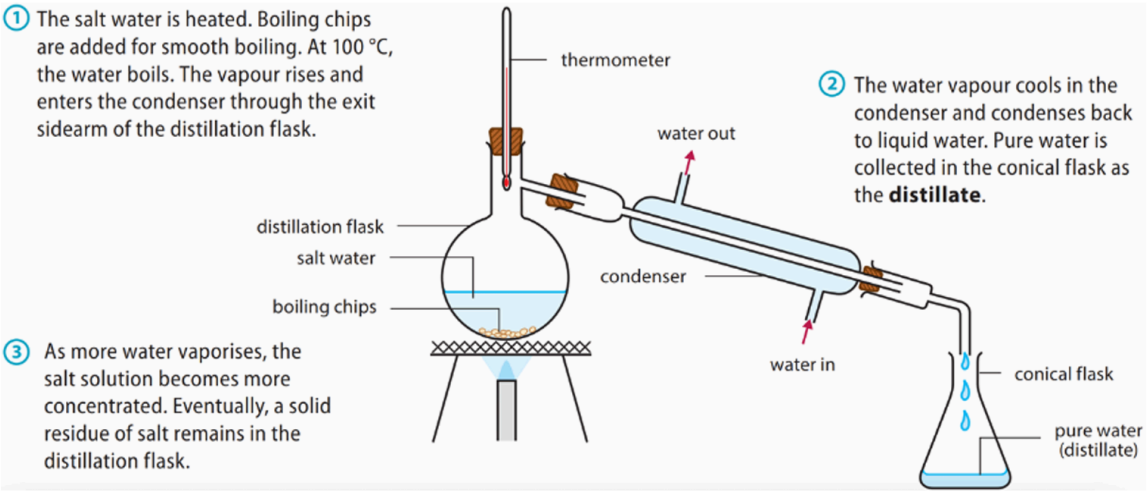
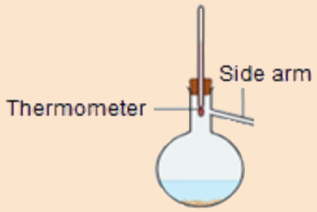


Fig 1.3.3. Diagram showing the setup to distill pure water from saltwater using simple distillation

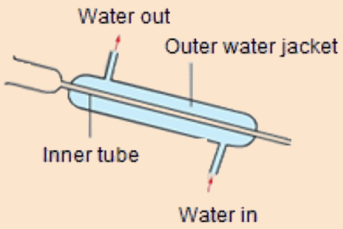
Key Qns 1.3.2 *(Simple Distillation)*

Q. Why is the bulb of the thermometer placed beside the side arm of the distillation flask?



To accurately measure the temperature of the vapour that enters the condenser, which is the boiling point of the substance.

Q. Why is cold running water allowed to enter the water jacket from the bottom of the condenser and leave from the top instead of the other way round?



If water enters from the top of the condenser, it will exit the condenser before the water jacket can be completely filled. Thus, the water has to enter the condenser from the bottom to ensure that the entire jacket is always completely filled so that vapour can be cooled efficiently and condensed into liquid quickly.

Separating Liquid-Liquid Mixtures

Definition 1.3.5 *(Miscible Liquids)*

Miscible liquids are those which form a uniform (homogeneous) solution when mixed together. These liquids are soluble in each other and will dissolve when mixed together.

Definition 1.3.6 *(Heterogeneous Mixtures)*

A *heterogeneous* mixture contains immiscible liquids that do not form a uniform solution. These liquids are insoluble in each other. Thus, when mixed together, these liquids separate into layers known as *phases*.

1. Separating Funnel

A separating funnel is used to separate **immiscible** liquids. When a heterogeneous mixture is left to settle into separate layers, each component can be removed by opening the tap at the bottom and collecting it in separate flasks or beakers.

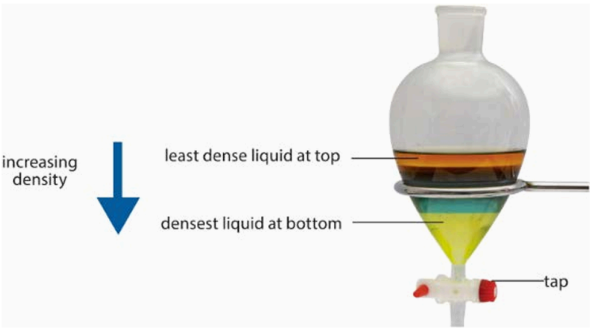


Fig 1.3.4. The figure shows a heterogeneous mixture of various liquids separating into phases of highest density (bottom) and lowest density (top).

2. Fractional Distillation

Fractional distillation is used to separate **miscible liquids** with different boiling points.

A fractionating column is attached to the round-bottomed flask and the condenser. The column contains glass beads or other small solid objects which provide a **larger surface area** for the vapours to condense on.

The vapours of the liquids will condense along the fractionating column and fall back into the round-bottomed flask.

The liquid with the **lowest boiling point** will distill over first. This means that the initial distillate collected will contain the substance with a lower boiling point.

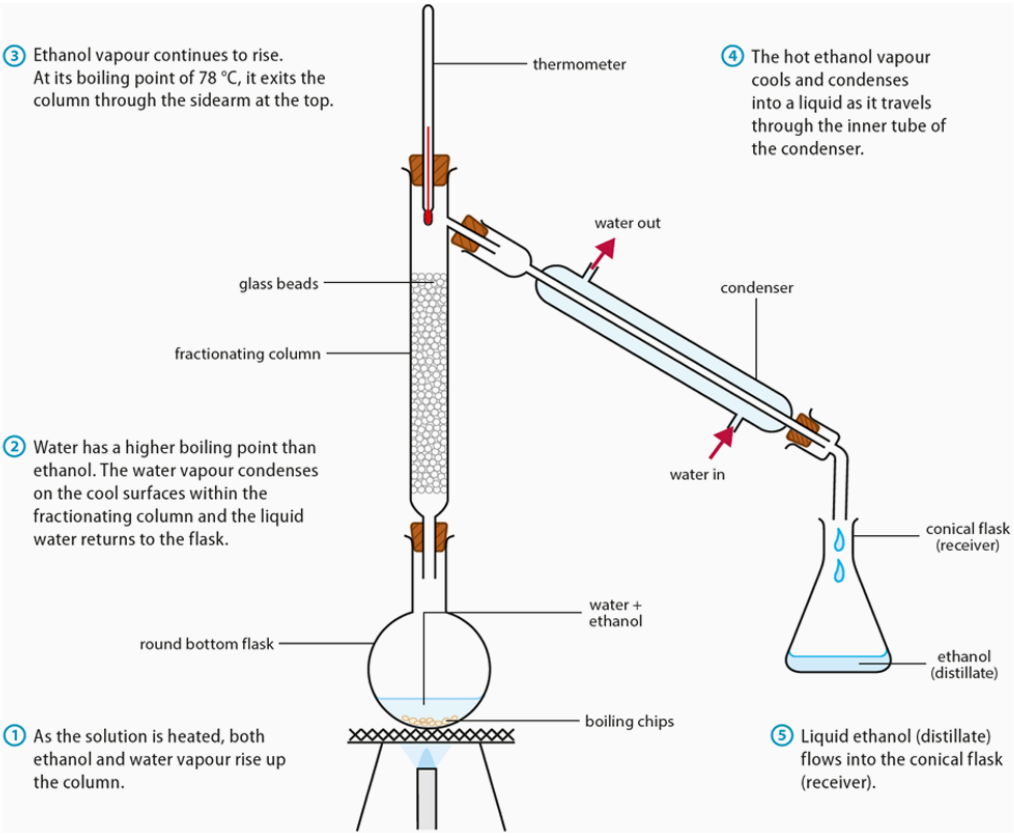


Fig 1.3.5. Fractional distillation setup to separate ethanol from a solution of ethanol and water

3. Chromatography

Paper chromatography is used to separate a mixture of miscible liquids which have different solubilities in a given solvent. (eg. food colourings & amino acids)

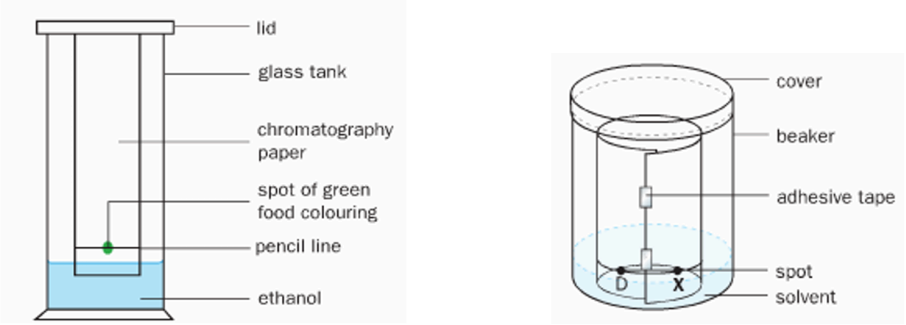


Fig 1.3.6. The diagram shows two setups for chromatography

The components in the food colouring mixture can be separated and identified using a suitable solvent due to the differences in solubilities of the different components in the chosen solvent.

> ⓘ **Note**
> The chromatography paper which contains the separated components is called a *chromatogram*

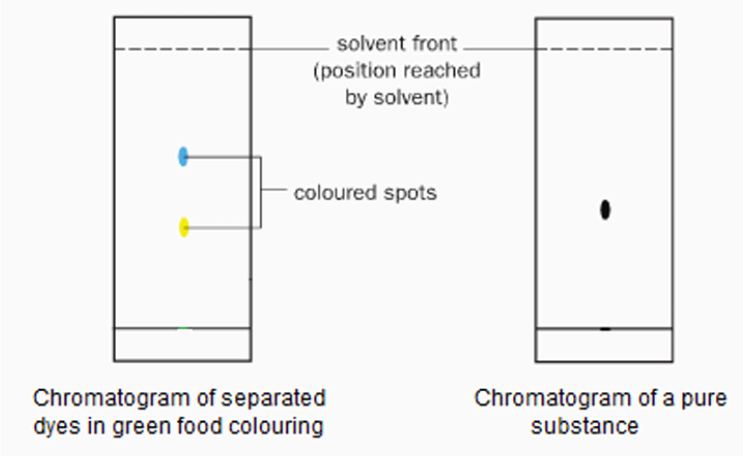


Fig 1.3.7. Figure shows the comparison between the chromatogram of a pure substance and a green mixture of dyes.

The more soluble the substance is in the solvent used, the further it will travel up the chromatogram and they will move more rapidly towards the solvent front. The less soluble substance will move at a slower rate. Identical components will travel the same distance.

Key Qns 1.3.3 *(Chromatography)*

Q. Why do you think the green food colouring disappears from the start line in the resulting chromatogram?

The green food colouring has dissolved in the solvent and separated into its components which moved up the paper.

Q. Why must the start line be drawn in pencil?

Pencil lead is made of graphite, which is insoluble in most substances. Thus, it would not introduce additional spots in the chromatogram.

Q. At the start of the experiment, why must the solvent be below the green food colouring?

To ensure that the green food colouring does not dissolve into the solvent below.

Definition 1.3.7 *(Retention Factor)*

R_f value of a substance is the ratio of the distance travelled by the solute to the distance travelled by the solvent.

This ratio is a constant as long as the chromatography is carried out using the same solvent under the same temperature. The R_f value for every substance is unique.

> ⓘ **Note**
> R_f value is always written in 2d.p.

Definition 1.3.8 *(Locating Agents)*

Locating agents are chemicals that react with the colourless substances to form coloured spots.

Locating agents can be used to identify colourless substances such as amino acids. A locating agent is sprayed onto the resulting chromatogram to make the colourless components visible to the eyes.

Examples of locating agents: Ninhydrin, silver nitrate and potassium permanganate

1.4 | Determining Purity of Substances (Experimental Chem)

Melting Point
Impurities **lower the melting point**. Pure substance (elements and compounds) will melt at a fixed temperature. A mixture which is an impure substance, will melt over a range of temperatures which is lower than the melting point of the pure substance.

Boiling Point
Impurities **increase the boiling point**. Pure substance (elements and compounds) will boil at a fixed temperature. A mixture which is an impure substance, will boil over a range of temperature which is higher than the boiling point of the pure substance.

Chromatography
If there are two or more spots on the chromatogram, the substance is impure. Paper chromatography only works for **substances that are soluble in the solvent**. In cases where any substances are insoluble in the solvent, it will not show up as spots on the chromatogram.

2.1 | How Are Solids, Liquids and Gases Different? (KPT)

Definition 2.1.1 *(Kinetic Particle Theory)*

The kinetic particle theory states that all matter is made up of tiny particles that are in constant random motion. These tiny particles can be atoms, molecules or ions.

These particles have **mass and occupy space.** This explains why matter is a substance that **has mass and occupies space**.

Table 2.1.1. Distance, arrangement and movement of particles at various states of matter

State	Attractive forces between particles	Arrangement (Distance + Arrangement)	Movement
Solid	Very strong	Particles are **very closely packed** together. Particles are **orderly arranged** with fixed positions.	Particles can only vibrate and rotate about their fixed positions.
Liquid	Strong	Particles are **closely packed** together. Particles are **disorderly arranged**.	Particles can slide past one another freely throughout the liquid.
Gaseous	Weak	Particles are **spaced far apart**. Particles are **disorderly arranged**.	Particles can move about randomly and freely in any direction rapidly, occupying any available space.

Key Qns 2.1.1 *(Density of Bromine)*

Q. Why does liquid bromine have a much higher density than gaseous bromine?

In liquid state, the bromine molecules are closely packed together. In gaseous state, these molecules are far apart.

In the same unit volume of bromine, there are a lot of bromine particles in liquid state and much fewer particles in gaseous state.

Thus, the total mass of molecules per unit volume of liquid bromine is much higher than that of gaseous bromine.

***Definition 2.1.2** (Temperature)*

Temperature is the measure of the average kinetic energy of the particles in a substance.

Heating
When a substance is heated, thermal energy is transferred to the substance. Some of the energy is converted to kinetic energy of the particles. The increased kinetic energy of the particles result in an increase in temperature of the substance.

Cooling
When a substance is cooled, thermal energy is transferred from the substance to its surroundings. The kinetic energy of the particles in the substance is reduced, which results in a drop in temperature of the substance.

***Key Qns 2.1.2** (Changes of state)*

Q. Why does temperature remain constant during changes of state?

The temperature remains constant during a change of state because the average kinetic energy of the particles does not change. Instead, the thermal energy supplied (or removed) is used to overcome (or strengthen) the attractive forces between particles.

Melting of A Substance

***Definition 2.1.3** (Melting)*

Melting is the process by which a solid changes to a liquid at the melting point of the substance.

Particles in the solid vibrate and move faster as they are heated as they gain kinetic energy, until they have sufficient energy to overcome the forces of attraction holding them in their fixed positions. The particles eventually break away from their fixed positions and move around one another randomly.

Freezing of A Substance

***Definition 2.1.3** (Freezing)*

Freezing is the process by which a liquid changes to a solid at the freezing point of the substance.

Particles in the liquid slow down and kinetic energy decreases, as energy is lost to the surroundings as heat when the liquid is cooled. The particles eventually settle into their fixed positions in an orderly arrangement, forming a solid.

Expansion and Contraction of Solids

When a solid is heated at temperatures below its melting point, the particles cannot spread out freely as they are **strongly attracted to each other**. Instead, the particles **vibrate more quickly about their positions**, with a slightly wider spacing than before. The solid has undergone expansion.

When a solid is cooled, thermal energy is transferred from the particles to the surroundings. The particles possess less kinetic energy, so they **vibrate slower and come closer than before**. The solid has undergone contraction.

Evaporation

Definition 2.1.4 *(Evaporation)*

Evaporation is the process by which a liquid changes to a gas at temperatures lower than its boiling point. Heat is taken in from the surroundings during the process.

At the surface of the liquid in contact with air, some particles have enough energy to overcome the attractive forces to escape into the air as a vapor.

Boiling

Definition 2.1.5 *(Boiling)*

Boiling is the process by which a liquid changes to a gas throughout the liquid at the boiling point of the substance. Heat is taken in from the surroundings during the process.

Particles in liquid state move faster when heated as they gain kinetic energy, until they have sufficient energy to completely overcome the strong forces of attraction holding them together. The particles become far apart from each other, moving freely and randomly in all directions, forming a gas.

Table 2.1.2. Differences between evaporation and boiling

Evaporation	Boiling
Occurs only on the surface of the liquid	Occurs throughout the liquid
Occurs at any temperature	Occurs only at boiling point
A slow process	A fast process

Condensation

Definition 2.1.6 *(Condensation)*

Condensation is the process by which a gas changes to a liquid. Heat is given out to the surroundings during the process.

When a gas is cooled, kinetic energy of particles is converted to thermal energy, transferred to the surroundings. The volume of gas decreases as the spacing between particles reduces as it cools. The gas continues to contract until the particles come close enough to form a liquid.

Sublimation

Definition 2.1.7 *(Sublimation)*

Sublimation is a process in which a substance changes from a solid to a gas without passing through the liquid state. (Same as *Definition 1.3.3*)

During sublimation, thermal energy from the surroundings is transferred to the solid, the solid particles gain kinetic energy and completely overcome the forces of attraction, resulting in the substance converting from a solid to a gas. Temperature remains constant until all the substance is in gaseous state.

The particles at the surface of the solid have enough kinetic energy to break away from the solid and escape as gas. Heat is taken in from the surroundings during this process.

Vapour deposition

Definition 2.1.8 *(Vapour Deposition)*

Vapour deposition is the process by which a gas changes directly to a solid without going through the liquid state.

During vapour deposition, the particles in the gas are cooled until they slow down and arrange themselves directly into solid state. Heat is given out to the surroundings during this process.

2.2 | How Do Particles Move? (KPT)

Diffusion

Definition 2.2.1 *(Diffusion)*

Diffusion is the net movement of particles from a region of higher concentration to a region of lower concentration.

Simple diffusion occurs mainly in liquids and gases. Vacancy diffusion or interstitial diffusion occurs mainly in solids such as metals and alloys.

Conditions that affect rate of diffusion

Table 2.2.1. What affects diffusion rate?

Conditions	Rate of diffusion	Reason
Increase temperature	Increase	More thermal energy is converted to kinetic energy of the particles.
Increase M_r	Decrease	Particles with higher mass require more kinetic energy to move at the same speed.

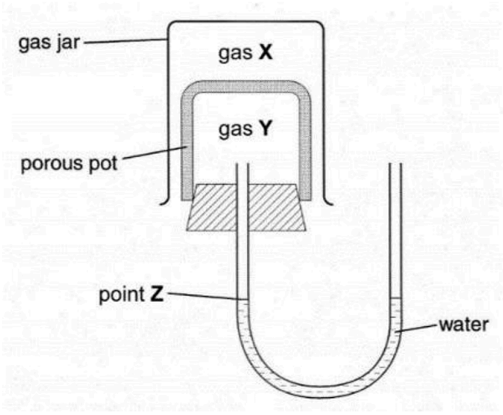


Fig 2.2.1. Experiment using a porous pot to compare rate of diffusion of two gases X and Y

Gas X	Gas Y	Observation + Reason
Hydrogen	Air	M_r of gas X (hydrogen) is 2 while the approximate relative mass of air is 29. Since hydrogen gas has a lower M_r compared to air, hydrogen will diffuse into the porous pot faster than air can diffuse out. This results in the pressure in the porous pot being greater than the pressure outside the porous pot, hence resulting in the water level at point Z to drop.

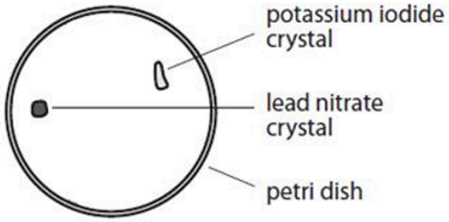


Fig 2.2.2. Figure shows an experiment carried out, dissolving KI and $Pb(NO_3)_2$ crystals in water to investigate diffusion.

Explanation for the formation of yellow (lead(II) iodide crystals)
Yellow precipitate is observed to form in the region closer to where lead(II) nitrate crystal was added. The crystals dissolved into the water as ions and diffused into the spaces between the water molecules, moving from a higher concentration to a lower concentration.

The rate of diffusion is **inversely related** to the relative formula mass of the particles at constant temperature. The relative formula mass of lead(II) ions is 207, higher than the relative molecular mass of iodide ions' 127. Thus, the rate of diffusion of lead(II) ions is slower and the yellow precipitate of lead(II) iodide forms closer to lead nitrate.

3.1 | What Is An Atom Made Up Of? (Atomic Structure)

Substances			Particles					
Elements	Compound	Mixtures	Atoms	Molecules	Ions	Electrons	Protons	Neutrons

Definition 3.1.1 *(Element)*

Element is a pure substance that cannot be broken down into two or more simpler substances by chemical methods.

Definition 3.1.2 *(Atom)*

An atom is the smallest particle that can still have the chemical characteristics of an element.

Definition 3.1.3 *(Compounds)*

Compound is a pure substance containing two or more types of elements that are chemically combined in a fixed ratio.

3.2 | Subatomic Particles (Atomic Structure)

Subatomic Particles

Subatomic particle	Relative atomic mass	Charge	Location
Proton (p)	1	1+	In the nucleus
Neutron (n)	1	0	
Electron (e^-)	$\frac{1}{1840}$	1-	Orbiting the nucleus

An atom has an equal number of protons and electrons, thus it is electrically neutral.

Nuclide Notation

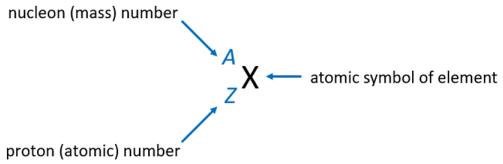


Fig 3.1.1. Diagram showing nuclide notation

Electron Shells

Electron shells have different energy levels, with the one closest to the nucleus as the lowest energy level. The further the electron shell from the nucleus, the higher its energy level and the less stable the electrons will be. Electrons will occupy electron shell with the lowest energy level (inner shell) first.

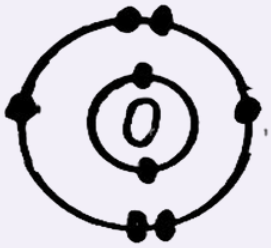
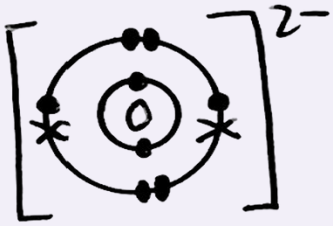
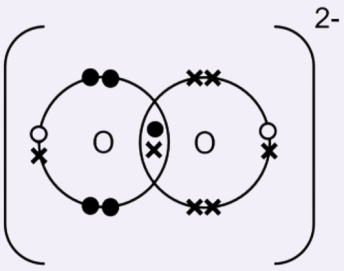
The outermost shell is also known as the *valence shell*, and the electrons it contains are known as *valence electrons.*

Ions

Definition 3.1.4 *(Ions)*

An ion is a charged particle formed when an atom or group of atoms gains or loses electron(s) but the number of protons and neutrons remains the same.

Dot-And-Cross Structure (Prerequisite)

Oxygen atom	Oxide ion	Peroxide ion
		

Isotopes

Definition 3.1.5 *(Isotopes)*

Isotopes are atoms with the same number of protons and electrons but have different numbers of neutrons.

Chemical Properties of Isotopes

Isotopes of an element have similar chemical properties as they have the same number of valence electrons. This means the isotopes will lose, gain or share the same number of electrons during a chemical reaction to achieve a noble gas configuration.

Physical Properties
Different physical properties as their masses are different due to different numbers of neutrons. Will differ in density, melting point and boiling point.

Table 3.1.1. Applications of isotopes

Technetium-99 (Tc)	Detection of tumors
Iodine-131 (I)	Treatment of thyroid disorder
Californium-252 (Cf)	Detection of explosives
Americium-241 (Am)	Used in smoke detectors
Carbon-14 (C)	Estimating the age of things that have carbon
Uranium-238 (U)	Estimating the age of rocks

Calculating Relative Atomic Mass (A_r) of an element

$$A_r = \sum(isotopic\ mass\ *\ isotopic\ abundance\ (\%))$$

Table 3.1.2. Chemical formula of ions

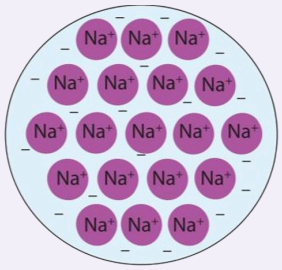
Cations		Anions	
Silver	Ag^+	Hydroxide	OH^-
Ammonium	NH_4^+	Nitrate	NO_3^-
Hydrogen	H^+	Ethanoate	CH_3COO^-
Zinc	Zn^{2+}	Sul**fate**	SO_4^{2-}
ⓘ Note Copper, Iron, Lead and Chromium can form multiple ions as they are transition metals. (II) and (III) refer to 2+ and 3+ charges in their respective ions.		Sul**fite**	SO_3^{2-}
		Carbonate	CO_3^{2-}
		Thiosulfate	$S_2O_3^{2-}$
		Phosphate	PO_4^{3-}
		Manganate (VII)	MnO_4^-

4/5.1 | Ionic Bonding (Chem Bonding & Structure)

Elements

Elements can appear in the following forms. (Note *Definition 3.1.1*)

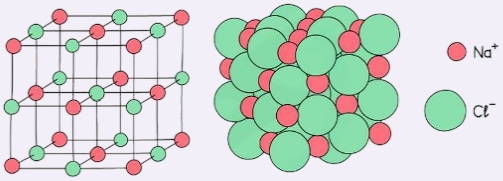
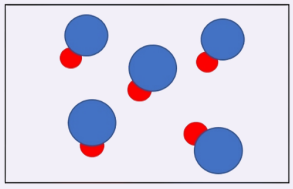
Table 4.1.1. Types of elements

Atoms of the same element	Molecules of the same element	A lattice of positive metal ions with a sea of negatively charged delocalised electrons
		

Compounds

All compounds have a chemical formula. Elements are bonded to form two types of compounds. (Note *Definition 3.1.3*)

Table 4.1.2. Types of compounds

Ionic Compounds	Covalent Compounds
 The above ionic compound, $NaCl$, exists as cations & anions which are formed when atoms of a metal lose electrons to atoms of a non-metal to achieve stable electronic configuration of noble gas.	 The above covalent compound, HCl, exists as diatomic molecules. These molecules are formed when the 2 types of atoms share electrons with each other to achieve stable electronic configuration of noble gas.

Mixtures

***Definition 4.1.1** (Mixtures)*

A mixture is an impure substance containing two or more substances (elements and/or compounds) that are physically mixed in any ratio.

Types of Mixtures

Mixture	Example
Elements	Alloys such as steel
Compounds	Aqueous solution of salts
Elements and Compounds	Air

Compounds vs Mixtures

	Compounds	Mixtures
What is it made up of?	Made up of two or more types of elements that are chemically combined	Made up of two or more elements and/or compounds that are physically combined
Ratio of its constituents	Fixed ratio	No fixed ratio
Properties	Different properties from its constituents elements	Similar properties to its constituents elements
Melting/Boiling point	Fixed	Range of temperatures

Separation of Compounds

- Thermal decomposition (exposing the compound to strong heating)
- Electrolysis (passing electric current through the compound)

Ion Formation

Definition 4.1.2 *(Cation)*

A cation is a positively charged particle formed when an atom loses electrons to other atoms.

Definition 4.1.3 *(Anion)*

An anion is a negatively charged particle formed when an atom gains electrons from other atoms.

Definition 4.1.4 *(Ionic Compounds)*

Ionic compounds are pure substances which are made up of cations and anions which are held together by strong ionic bonds.

Definition 4.1.5 *(Ionic Bonding)*

Ionic bonding is the strong electrostatic forces of attraction between cations and anions.

Formation of Ionic Bonds

When a metal reacts with a non-metal, electrons are transferred from the atoms of metal to the atoms of non-metal. In the reaction, the metal atom loses electrons to the non-metal atom to form a cation. The non-metal atom gains electrons from the metal atom to form an anion. The product of the reaction is an ionic compound which is made up of cations and anions. The strong electrostatic forces of attraction between cations and anions in an ionic compound is known as *ionic bond*.

Giant Ionic Lattice Structure

Type of matter	Ionic compounds
Type of structure	Giant ionic lattice structure
Types of particles in the structure	Cations and anions
Forces of attraction between particles in structure	Strong electrostatic forces of attraction (strong ionic bonds)
Movement and arrangement of particles in the structure	Ions are orderly arranged. Ions vibrate about fixed positions and are very closely packed together, in a fixed ratio.

Property #1: High melting and boiling points

Explain, in terms of structure, why calcium oxide is a solid at r.t.p.

- Calcium oxide is an ionic compound with a giant ionic lattice structure. [1]
- In this structure, the calcium ions and oxide ions are held together by strong electrostatic forces of attraction (ionic bonds). [1]
- A lot of energy is required to overcome these strong forces to melt CaO. Hence CaO has a high melting point and is a solid at r.t.p. [1]

Explain, in terms of structure, why calcium oxide has a higher melting point than sodium chloride.

- The charge of calcium ion is 2+ and the charge of oxide ion which is 2-. [1] The charge of sodium ion which is 1+ and the charge of chloride ion which is 1-.
- Thus, the electrostatic forces of attraction between Ca^{2+} and O^{2-} ions is stronger than the forces between Na^{+} ions and Cl^{-} ions. [1]
- Thus, more energy is required to overcome the stronger forces to melt CaO than NaCl. [1] Hence, CaO has a higher melting point than NaCl.

Property #2: Very hard and have regular shapes

Explain why the lattice structure of magnesium oxide causes the solid crystals to be 'very hard and have regular shapes.'

- To separate the Mg^{2+} ions and O^{2-} ions from each other in the giant ionic lattice structure, a lot of force must be applied to overcome the strong electrostatic forces of attraction which hold these ions in fixed positions. Hence MgO is hard. [1]
- The strong electrostatic force of attraction held the ions closely packed together and orderly arranged in fixed positions. Hence the crystal of MgO has a regular shape. [1]

Property #3: Good conductors of electricity when molten or in aqueous state. Cannot conduct electricity in solid state.

Explain, in terms of structure, why molten sodium bromide is a good conductor of electricity.

- The Na^{+} and Br^{-} ions in molten $NaBr$ are mobile hence these ions can move freely throughout the substance and act as charge carriers.

Explain why sodium chloride has to be dissolved in water, or melted, before it will conduct electricity.

- To overcome the strong electrostatic forces of attraction which held the Na^{+} ions and Cl^{-} ions in fixed positions [1]
- in the giant ionic crystal lattice so that the ions can be mobile[1] and act as charge carriers and conduct electricity.

Explain, in terms of structure, why sodium bromide crystals does not conduct electricity but aqueous sodium bromide is a good conductor of electricity

- Sodium bromide crystal has a giant ionic lattice structure. In this structure, the sodium and bromide ions are not mobile as they are held in fixed positions by strong electrostatic forces of attraction.
- In aqueous sodium bromide, the sodium and bromide ions are mobile. Hence these ions can act as charge carriers and can conduct electricity.

Property #4: Usually soluble in water (with some exceptions)

Explain why sodium nitrate is soluble in water.

- When sodium nitrate is mixed with water, water molecules are attracted to the sodium and nitrate ions.
- The forces of attraction between the water molecules and the Na^+ and NO_3^- ions are stronger than the electrostatic forces of attraction between the Na^+ and NO_3^- ions.
- As a result, the ions are pulled away from their fixed positions in the ionic lattice structure and sodium nitrate dissolves in water to form an aqueous solution.

Explain why magnesium oxide is insoluble in water.

- In the giant ionic lattice structure of MgO, the magnesium ions and oxide ions are held together by strong electrostatic forces of attraction. [1]
- These forces are stronger than the forces of attraction between the water molecules and the Mg^{2+} and O^{2-} ions. Thus, the ions remained in fixed positions in the lattice structure and could not dissolve in water. [1]

4/5.2 | Electronic Struct of Covalent Substances (Chem Bonding & Structure)

Molecules

When non-metal atoms react with one another, these atoms become chemically combined together by sharing electrons with each other to achieve the stable electronic configuration of noble gas.

Definition 4.2.1 *(Molecules)*

A molecule is an electrically neutral particle made up of two or more atoms which are chemically combined together through the sharing of electrons.

Usually, all atoms in a molecule of any covalent substance have the **stable electronic configuration of a noble gas**. There are some exceptions to this rule though (eg. group 5 elements *B* and *Al*).

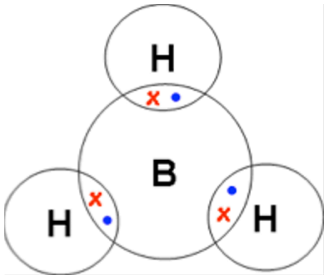


Fig 4.2.1. Boron trihydride — an exception to the octet rule

Definition 4.2.2 *(Molecule of Element)*

Molecule of element is an electrically neutral particle made up of two or more atoms of the same kind which are chemically combined together through the sharing of electrons.

Definition 4.2.3 *(Molecule of Compound)*

Molecule of compound is an electrically neutral particle made up of two or more atoms of different elements which are chemically combined together through the sharing of electrons.

Formation of Molecules

When non-metal atoms react with one another, these atoms become chemically combined together by sharing electrons with each other to achieve the stable electronic configuration of noble gas. The atoms in a molecule are held together by *covalent bonds.*

Definition 4.2.4 *(Covalent Bonds)*

The strong electrostatic forces of attraction between positively charged nuclei of the two atoms and negatively charged electrons which are shared between two atoms in the molecule.

Types of Covalent Bonds

Single bonds	Fluorine: $F - F$ Hydrogen chloride: $H - Cl$
Double bonds	Oxygen: $O = O$ Carbon Dioxide: $O = C = O$
Triple bonds	Azide radical: $N \equiv N$ Acetylene: $H - C \equiv C - H$

Definition 4.2.5 *(Valency)*

Valency refers to the number of electron(s) that must be lost, gained or shared in order for the atom to attain the stable electronic configuration of noble gas.

Polyatomic Ions and Dative Bonding

Polyatomic ion is an ion that contains more than one atom. The atoms are held by covalent bonds and the charges on the polyatomic ions are due to losing or gaining of electrons by the ion.

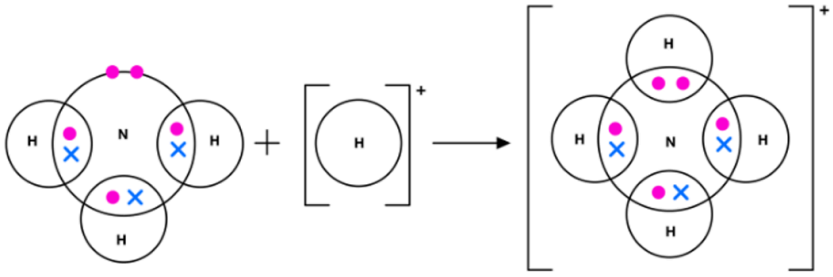


Fig 4.2.2. Dative (coordinate) bonding of the lewis base NH_3 with H^+ to form NH_4^+ polyatomic cation.

Table 4.2.1. Chemical names and chemical formula of common covalent substances

Molecules of elements	Molecules of compounds
Hydrogen, H_2	Methane, CH_4
Nitrogen, N_2	Ammonia, NH_3
Fluorine, F_2	Hydrochloric acid, HCl
Oxygen, O_2	Nitric acid, HNO_3
Iodine, I_2	Sulfuric acid, H_2SO_4
Chlorine, Cl_2	Ethanoic acid, CH_3COOH
Bromine, Br_2	Ethanol, C_2H_5OH
Astatine, At_2	Hydrogen peroxide, H_2O_2
Ozone, O_3	

Naming Binary Covalent Compounds

1) Identify the element furthest to the left and bottom of the periodic table
2) Add an "ide" ending to the other element
3) Depending on the number of atoms in this other element, select the prefix

Number of Atoms in Compounds	Prefix for the element
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

4/5.3 | Structure and Properties of Covalent Substances (Chem Bonding & Structure)

Overview of Structure of Covalent Substances

- Simple covalent molecules
- Giant covalent structure
- Macromolecules

Simple Covalent Molecules

Type of matter	Covalent substance
Type of structure	Simple covalent molecules
Types of particles in the structure	Molecules
Forces of attraction between particles in structure	Weak intermolecular forces of attraction **between the molecules**

> ⓘ Note
> *Empirical formula* is the simplest ratio of atoms of the covalent substance from its molecular formula.

Property #1: Low melting and boiling points

Explain, in terms of structure, why methane is a gas at r.t.p. [3]

- Methane is a covalent substance that exists as simple covalent molecules. [1]
- In this structure, there are only weak intermolecular forces of attraction between methane molecules. [1]
- Only a small amount of thermal energy is required to overcome these weak forces to melt and boil CH_4. Hence, CH_4 has low melting and boiling points [1] and is a gas at r.t.p.

Explain, in terms of structure, why ethanol is a volatile liquid at r.t.p. [3]

- Ethanol is a covalent compound that exists as simple covalent molecules. [1]
- In this structure, there are only weak intermolecular forces of attraction between ethanol molecules. [1]
- Only a small amount of thermal energy is required to overcome these weak forces to melt and boil ethanol. As ethanol has low melting and boiling points, [1] it is a volatile liquid as it evaporates quickly at r.t.p.

Property #2: Poor conductors of electricity (with some exceptions)

Explain why pure water is a poor conductor of electricity.

- Pure water exists as simple covalent water molecules that are electrically neutral. Pure water does not have freely moving ions or delocalised electrons that can act as charge carriers to conduct electricity.

Explain why hydrogen chloride gas cannot conduct electricity but once the gas is dissolved in water, the aqueous solution can conduct electricity.

- Hydrogen chloride gas exists as simple covalent HCl molecules that are electrically neutral. Hydrogen chloride gas does not have freely moving ions or delocalised electrons to act as charge carriers to conduct electricity.
- When hydrogen chloride gas dissolves in water, an aqueous solution of hydrochloric acid is formed. In aqueous state, HCl molecules ionise in water to form H^+ and Cl^- ions which are mobile. These ions can act as charge carriers and conduct electricity.

Property #3: Generally insoluble in water

These covalent substances are usually soluble in organic solvents such as oil, benzene, esters, tetrachloromethane, dichloromethane, hexane

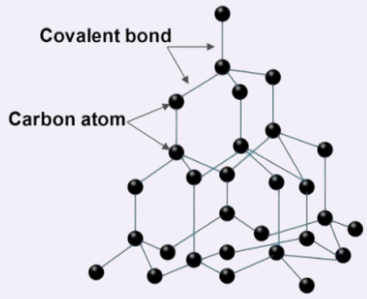
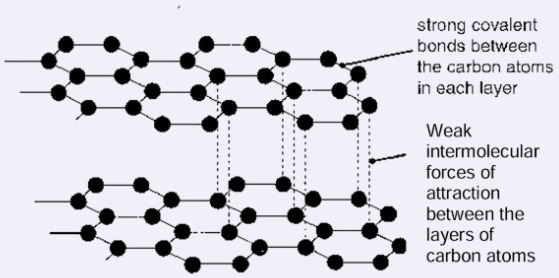
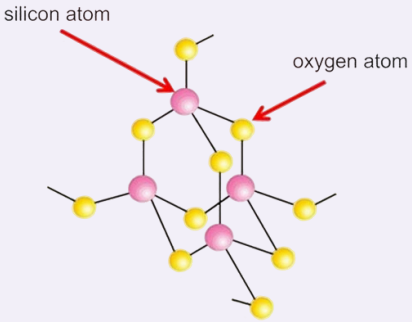
(i) Chlorine (Cl_2)
(ii) Hydrogen chloride (HCl)
(iii) Ammonia (NH_3)
(iv) Sulfur dioxide (SO_2)
(v) Sulfur trioxide (SO_3)
(vi) Ethanol (C_2H_5OH)
(vii) Glucose ($C_6H_{12}O_6$)

Fig 4.3.1. Just some exceptions that are soluble in water

Giant Covalent Structure

Type of matter	Covalent substance
Type of structure	Giant covalent structure
Types of particles in the structure	Atoms
Forces of attraction between particles in structure	Strong covalent bonds between atoms
Arrangement of particles	Atoms are held in fixed positions in the giant rigid tetrahedral structure

Table 4.3.1. Substances with giant covalent structures

Diamond Eg. Drill tips		C
Graphite Eg. Dry lubricants/ Pencil lead		C
Silicon dioxide Eg. Glassware		**Deducing chemical formula for silicon dioxide** Each oxygen atom is bonded to a maximum of 2 silicon atoms Each silicon atom is bonded to a maximum of 4 oxygen atoms Simplest ratio of Si to O is 1:2, hence SiO_2.

Definition 4.3.1 *(Allotropes)*

Allotropes are different forms of the same element with different structural arrangement of atoms.

Property #1: High melting and boiling points

Explain, in terms of structure, why graphite has high melting and boiling points. [3]

- Graphite is a covalent substance with a giant covalent structure. [1]
- Graphite is made up of a giant network of **planar hexagonal ring layers** of carbon atoms. Within each layer, each carbon atom is bonded to three other carbon atoms by strong covalent bonds. [1]
- A large amount of thermal energy is required to break the strong covalent bonds to melt and boil graphite. [1] Thus, graphite has high melting and boiling points.

Explain, in terms of structure, why sand is non-volatile. [3]

- Sand is mainly made up of silicon dioxide which is a covalent substance with a giant covalent structure. [1]
- Each silicon atom is bonded to four oxygen atoms by strong covalent bonds throughout the giant rigid tetrahedral structure. Each oxygen atom is bonded to two silicon atoms by strong covalent bonds throughout the giant rigid tetrahedral structure. [1]
- A large amount of thermal energy is required to overcome the strong covalent bonds to melt and boil silicon dioxide. Thus, sand has a high melting and boiling point range. [1] Hence, sand is non-volatile.

Property #2: Some covalent substances with giant covalent structures are hard while others are soft.

Explain, in terms of structure, why silicon dioxide is hard. [3]

- Silicon dioxide is a covalent substance with a giant covalent structure. [1]
- Each silicon atom is bonded to four oxygen atoms by strong covalent bonds throughout the giant rigid tetrahedral structure. Each oxygen atom is bonded to two silicon atoms by strong covalent bonds throughout the giant rigid tetrahedral structure. [1]
- A large amount of force must be applied to overcome the strong covalent bonds to displace the atoms from their fixed positions in the structure. [1] Thus, silicon dioxide is hard.

Explain, in terms of structure, why graphite is used to make dry lubricant. [4]

- Graphite is a covalent substance with a giant covalent structure. [1]
- Graphite is made up of a giant network of **planar hexagonal ring layers of carbon atoms.** The layers of carbon atoms are held together by weak intermolecular forces of attraction [1]
- Only a small amount of force needs to be applied to overcome these weak forces for the layers of atoms to slide past each other. [1]
- Thus, graphite is soft and slippery [1] and can be used as dry lubricant for machinery.

Property #3: Most are poor electrical conductors (EXCEPT graphite)

Explain, in terms of structure, why graphite is a good conductor of electricity. [2]

- Graphite is a covalent substance with a giant covalent structure. [1]
- Graphite is made up of a giant network of planar hexagonal ring layers of carbon atoms. Within each layer, each carbon atom is bonded to three other carbon atoms by strong covalent bonds.
- Each carbon atom has one valence electron that is not used to form covalent bonds. These valence electrons are delocalised and can move freely throughout the layers of carbon atoms in the structure and act as charge carriers to conduct electricity. [1]
- Thus, graphite is a good conductor of electricity,

Explain, in terms of structure, why silicon dioxide is a poor conductor of electricity. [2]

- Silicon dioxide is a covalent substance with a giant covalent structure. [1]
- In this structure, each silicon atom is bonded to four oxygen atoms by strong covalent bonds throughout the giant rigid tetrahedral structure.
- Each oxygen atom is bonded to two silicon atoms by strong covalent bonds throughout the giant rigid tetrahedral structure.
- All the valence electrons of each silicon atom and oxygen atom are involved in covalent bonding. There are no delocalised electrons that can act as charge carriers and move freely throughout the structure. [1]
- Thus, silicon dioxide is a poor conductor of electricity.

Property #4: Usually insoluble in water and organic solvents

Explain, in terms of structure, why silicon dioxide is insoluble in water. [3]

- Silicon dioxide is a covalent substance with a giant covalent structure. [1]
- Each silicon atom is bonded to four oxygen atoms by strong covalent bonds throughout the giant rigid tetrahedral structure. Each oxygen atom is bonded to two silicon atoms by strong covalent bonds throughout the giant rigid tetrahedral structure. [1]
- The forces of attraction between the atoms in the giant molecular structure and water molecules are not strong enough to overcome the strong covalent bonds within the structure.
- Thus, silicon dioxide cannot dissolve in water. [1]

Macromolecules

A macromolecule (polymer) comprises long chains of large molecules which are made up of many covalent molecules (monomers) that are joined together by covalent bonds.

Type of matter	Covalent substance
Type of structure	Macromolecules
Types of particles in the structure	Long-chain molecules
Forces of attraction between particles in structure	Weak intermolecular forces of attraction between the long – chain molecules.

Natural polymers: silk, wool, rubber, starch, proteins, lipids
Man-made polymers: polyester, poly(ethene), nylon, polystyrene and other plastics

Property #1: No fixed melting and boiling points (relatively low range)

Explain why polymers do not have fixed melting and boiling points.

- A polymer is a macromolecule which is made up of chains of large molecules. The long-chain polymer molecules do not have a fixed length as the number of monomers joined together to form the polymer is not fixed and the monomers may have a range of sizes.
- Thus, the amount of intermolecular forces of attraction between the chain of molecules will vary. Hence, polymers do not have fixed melting and boiling points.

When heated, polymers typically soften over a range of temperatures.

- Instead, when heated, polymers typically soften over a range of temperatures as the weaker intermolecular forces of attraction are overcome by the molecular vibrations with higher kinetic energy.
- This range of temperatures is known as softening temperature.
- The **larger the polymer, the higher the softening temperature**.

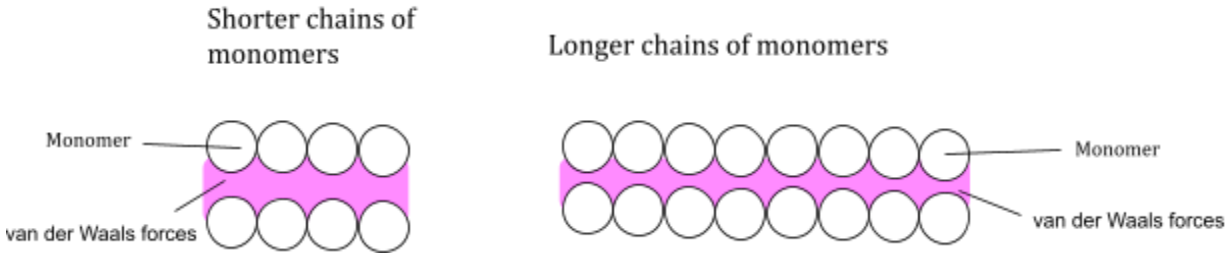


Fig 4.3.1. The polymer made up of longer chains of monomers usually have stronger intermolecular forces of attraction between the long-chain molecules as there is more surface area for contact.

Poly(ethene) which is a polymer has relatively low melting and boiling points range. Explain why.

- Poly(ethene) is a macromolecule which is made up of chains of large molecules. Only a small amount of thermal energy is required to overcome the weak intermolecular forces of attraction between the chains of molecules.
- Thus, poly(ethene) polymers which are macromolecules has low melting and boiling points range.

Although covalent substances that are macromolecules have low melting and boiling points range but most of them are solids at r.t.p ($25\ °C$**). Why?**

- A polymer is a macromolecule which is made up of chains of large molecules.
- The intermolecular forces of attraction between the chains of larger molecules are stronger.
- The melting and boiling points range of polymers are above room temperature of $25\ °C$, Hence, most of them will be solids at $25\ °C$.

Property #2: Poor electrical conductors at solid, liquid and gaseous state

Explain why poly(ethene) is a poor conductor of electricity. [1]

- Poly(ethene) is a macromolecule that consists of long chains of poly(ethene) molecules that are electrically neutral.
- Poly(ethene) does not have freely moving ions or delocalised electrons to act as charge carriers to conduct electricity.

Property #3: Solubility

Most polymers are **insoluble in water** but **soluble in organic solvents** (oil, benzene, esters, tetrachloromethane, dichloromethane, hexane)

4/5.4 | Metallic Bonding (Chem Bonding & Structure)

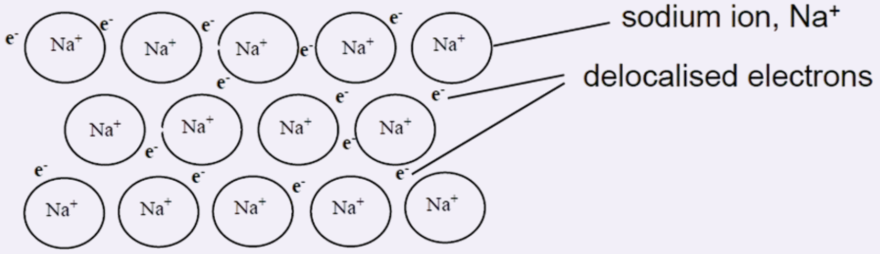
Some physical properties of metals
Shiny (lustrous)
Mostly solids at r.t.p (except for mercury which is a liquid)
Usually have high melting and boiling points
Malleable and ductile
Good conductor of heat
Good conductor of electricity
Usually high density (except lithium, sodium and potassium)

Definition 4.4.1 *(Metallic Bonding)*

Metallic bonding is the strong electrostatic forces of attraction between positively charged metal ions and a **sea of delocalised mobile negatively charged** valence electrons.

Type of matter	Metals
Type of structure	Giant metallic lattice structure
	In the **giant metallic lattice structure**, the metal atoms lose their valence electrons and become positively charged metal ions. The valence electrons no longer belong to any metal atom and are delocalised. These electrons move freely between the positive metal ions like a 'sea of delocalised mobile negatively charged valence electrons'. The metal ions are held together by metallic bonds — the electrostatic force of attraction between the positively charged metal ions and the sea of delocalised mobile negatively charged valence electrons.
Types of particles in the structure	Positively charged metal ions & negatively charged sea of delocalised electrons
Forces of attraction between particles in structure	Strong metallic bonds

Diagram



> ⓘ Note
> Metals are solid at r.t.p, hence the metal ions must be orderly arranged in the lattice structure. There must still be some spaces between the metal to be occupied by the sea of delocalised mobile valence electrons.

Property #1: High melting and boiling points

> ⓘ Note
> Exceptions are alkali metals (group 1 metals) that have relatively low melting and boiling points (although they are still solid at r.t.p). Another exception is mercury, which is liquid at r.t.p.

Explain, in terms of structure, why aluminium is a solid at r.t.p. [3]

- Aluminium is a metal with a giant metallic lattice structure. [1]
- In this structure, there are strong electrostatic forces of attraction between positively charged aluminium ions and negatively charged **sea of delocalised mobile valence electrons**. [1]
- A large amount of thermal energy is required to overcome these strong forces to melt *Al*. Hence Al has a high melting point and is a solid at r.t.p. [1]

Use the structure of metal to explain why magnesium has a higher melting point of (650 °C) than sodium (97.79 °C). [3]

- Sodium atoms have an electronic configuration of 2.8.1 and magnesium atoms have an electronic configuration of 2.8.2. This means that each magnesium atom will lose 2 valence electrons to form a higher positively charged metal ion, Mg^{2+} compared to sodium atom which will lose only 1 valence electron to form a *singly positively charged* metal ion, Na^{+}. [1]
- Hence, the electrostatic forces of attraction between metal ions and the delocalised valence electrons in magnesium is stronger than in sodium. [1]
- A large amount of thermal energy is needed to overcome the stronger electrostatic forces of attraction. Therefore, magnesium has a higher melting point than sodium. [1]

Which metal has a stronger metallic bond – sodium (melting point of 97.79 °C) or potassium (melting point of 63.5 °C)? Why? [3]

- Sodium ion has an electronic configuration of 2.8 and potassium ion has an electronic configuration of 2.8.8. Sodium ions have one less electron shell compared to potassium ions. Hence, sodium has a smaller ionic radius. [1]
- In the giant metallic lattice structure, the negatively charged valence electrons are closer to the positively charged nuclei of the sodium ions than potassium ions. Hence, the electrostatic force of attraction between the nucleus of the positive ions and the valence electrons would be stronger in sodium than in potassium. [1]
- Hence, more thermal energy is needed to overcome the stronger electrostatic forces of attraction in sodium, leading to a higher melting point. [1]

Property #2: High density

> ⓘ **Note**
> Exceptions are alkali metals (group 1 metals) that have relatively low densities. (Li, Na, K all float on water)

Explain why iron has high density while oxygen has low density at r.t.p. [3]

- Iron is a solid while oxygen is a gas at room temperature. In a given volume, iron atoms are very closely packed together while oxygen molecules are spaced far apart from each other. [1]
- The relative atomic mass of iron is 56 while the relative molecular mass of oxygen is 32. [1]
- Thus, in a fixed volume, there is a much larger number of iron atoms which have a higher relative atomic mass while there is a much smaller number of oxygen molecules which have a lower relative molecular mass. [1] Hence, iron has a higher density than oxygen at room temperature.

Property #3: Good electrical conductivity

Explain why sodium is a good conductor of electricity. [1]

- Sodium has delocalised valence electrons which act as charge carriers that can move freely throughout the giant metallic lattice structure. [1]
- When connected to an electrical circuit, these valence electrons are able to move from the negative terminal to the positive terminal.

Property #4: Good thermal conductivity

Explain why iron is a good thermal conductor. [2]

- Iron has delocalised valence electrons which can move freely throughout the giant metallic lattice structure. [1]
- When heated, these electrons absorb thermal energy, gain kinetic energy and move rapidly from the hotter end to the colder end, carrying the energy with them. [1]
- Hence, thermal energy is transferred easily by these delocalised electrons in the structure.

Property #5: Most metals are malleable and ductile

Definition 4.4.2 *(Malleability)*

Malleability refers to the ability of a material to be hammered into different shapes without breaking.

Definition 4.4.3 *(Ductility)*

Ductility refers to the ability of a material drawn into wires.

Explain why aluminium is malleable and ductile. [2]

- In the giant metallic lattice structure of aluminium, aluminium atoms which have the same atomic size are very closely packed and orderly arranged in layers. [1]
- When a force is applied, the layers of aluminium atoms can slide over one another easily. Thus, **aluminium is soft** and hence, malleable and ductile. [1]

Alloys

Definition 4.4.4 *(Alloy)*

A mixture of a metal and one or more other elements.

Table 4.4.1. Alloys and their properties and uses

Alloy	Composition	Properties	Uses
Brass	**Copper - 70%** Zinc - 30%	Harder and stronger Does not corrode easily Attractive yellow colour (like gold)	Decorative ornaments, coins, musical instruments
Bronze	**Copper - 88%** Tin - 12%	Harder and stronger Does not corrode easily Attractive red-brown colour (like copper)	Decorative ornaments, bells, medals
Mild steel	**Iron - 99.5%** Carbon - 0.5%	Harder and stronger Durable	Car bodies
Stainless steel	**Iron - 73%** Chromium - 18% Nickel - 8% Carbon - 1%	Harder and stronger Durable Highly resistant to corrosion	Cooking utensils, surgical equipment

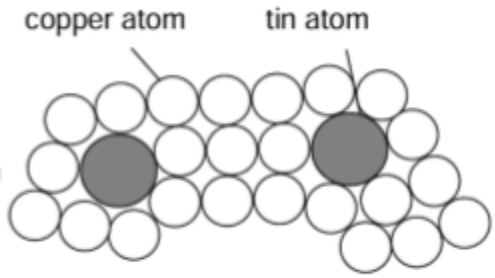


Fig 4.4.1. Structure of bronze (an alloy of copper and tin) showing tin atoms disrupting the orderly arrangement of copper atoms.

Property #1: Alloys are harder and stronger than pure metals

Thus, alloys are less malleable and less ductile compared to pure metals.

Explain why bronze is harder and stronger than pure copper. [3]

- Bronze is an alloy of copper and tin.
- As copper and tin have different atomic sizes, the orderly arrangement of copper atoms is disrupted by the presence of tin atoms. Hence layers of copper atoms cannot slide over each other easily when a force is applied.
- Pure copper is a constituent element in bronze. As copper atoms have the same atomic size, the copper atoms are orderly arranged in layers. Hence layers of copper atoms can slide over each other easily when a force is applied.

Property #2: Alloys do not have fixed melting and boiling points

Alloys are mixtures which are impure substances. Thus, alloys do not have fixed melting and boiling points. These substances melt and boil over a range of temperatures.

Property #3: Alloys are good electrical conductors

As alloys are mixtures containing at least one metal, alloys are good conductors of electricity as they have delocalised mobile valence electrons that can act as charge carriers and move freely throughout the alloy.

Property #4: Alloys are good thermal conductors

Alloys are mixtures containing at least one metal. Hence, like metals, alloys are good thermal conductors as they have delocalised mobile valence electrons.

6.1 | Chemical Formulas and Equations

> ⓘ Note
> Most of the concepts here are lower sec stuff so I will not be covering them. I will cover ionic equations though since that is new.

Monatomic Elements

Definition 6.1.1 *(Monatomic elements)*

Monatomic elements are pure substances that exist as single, uncombined atoms.

Diatomic Molecules

Definition 6.1.2 *(Diatomic molecules)*

Diatomic molecules are particles made up of two atoms that are covalently bonded together.

Polyatomic Molecules

Definition 6.1.3 *(Polyatomic molecules)*

Polyatomic molecules are particles made up of three or more atoms which are covalently bonded together.

> ⓘ Note
> Exception: Even though sulfur exists as S_8 molecule, when writing chemical reactions involving sulfur, we write S instead of S_8.

Chemical Formulae of Elements that have Giant Covalent Structures

Their chemical formula is just the chemical symbol of their element in the periodic table. Note that these giant covalent structures are usually formed by group 14 atoms.

Writing Ionic Equations

1	Write the balanced chemical equation for the reaction (eg. redox reaction for the production of chlorine gas): $$4HCl(aq) + MnO_2(s) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$$
2	Split/ionise ionic compounds or acids in aqueous state $$4H^+(aq) + 4Cl^-(aq) + MnO_2(s) \rightarrow Mn^{2+} + 2Cl^-(aq) + 2H_2O(l) + Cl_2(g)$$
3	Cancel out ions that appear on both sides of the chemical equation "spectator ions" that have the same physical state and charge. $$4H^+(aq) + 2Cl^-(aq) + MnO_2(s) \rightarrow Mn^{2+} + 2H_2O(l) + Cl_2(g)$$

7.1 | Mole Concept (Chem Calc.)

Relative Atomic Mass (A_r)

Definition 7.1.1 *(Relative Atomic Mass)*

The relative atomic mass (A_r) of an element is the average mass of one atom of that element relative to $\frac{1}{12}$ the mass of a carbon-12 atom.

Relative Molecular Mass (M_r)

Definition 7.1.2 *(Relative Molecular Mass)*

The relative molecular mass (M_r) of a molecular substance is the average mass of one molecule of that covalent substance relative to $\frac{1}{12}$ the mass of a carbon-12 atom.

Relative Formula Mass (M_r)

Definition 7.1.3 *(Relative Formula Mass)*

The relative formula mass (M_r) of an ionic compound is the average mass of one formula unit of that ionic compound relative to relative to $\frac{1}{12}$ the mass of a carbon-12 atom.

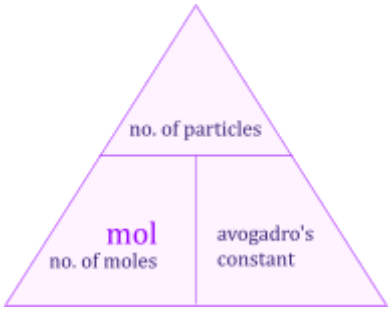
Mass of Element In Compound

$$\text{Mass of element in compound} = \frac{A_r \times No.\ of\ atoms\ of\ that\ element}{M_r\ of\ compound} \times Mass\ of\ compound$$

Mole

Definition 7.1.4 *(Mole)*

One *mole* of any substance always contains 6.02×10^{23} of particles of that substance.



$$\text{Moles of substance} = \frac{Number\ of\ particles}{6.02 \times 10^{23}}$$

Key Qns 7.1.1 *(Moles)*

How many hydrogen atoms are there in 3 moles of water?

comparing ratio of particles,

	H_2O molecule	:	H atoms	:	O atom
=	1	:	2	:	1

hence, comparing mol ratio of particles,

	H_2O molecules	:	H atoms	:	O atoms
	1 mol	:	2 mol	:	1 mol
=	3 mol	:	6 mol	:	3 mol

No. of hydrogen atoms in 3 mol of H_2O molecules
= number of moles of H atoms x Avogadro Constant
= 6 x 6.02 x 10^{23} (also accept 2 x 3 x 6.02 x 10^{23}) [1]
= 3.612 x 10^{24} (also accept 36.12 x 10^{23})
= **3.61 x 10^{24} (3 s.f.) [1]**

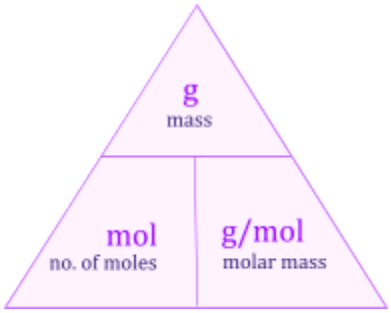
Molar Mass

Definition 7.1.5 *(Molar Mass)*

Molar mass is the mass of one *mole* of any substance.

ⓘ **Note**

Molar mass has the same numerical value of relative atomic mass (A_r) and relative molecular mass (M_r) of the substance, but it has **units** (g/mol)



$$\text{Molar mass} = \frac{Mass\ of\ substance}{Number\ of\ moles}$$

Key Qns 7.1.2 *(Ammonium Chloride)*

Calculate the number of moles in 10 kg of ammonium chloride.

Chemical formula of ammonium chloride is NH_4Cl.

10 kg = 10 x 1000 = 10 000 g

$$\text{number of moles of } NH_4Cl = \frac{mass\ of\ NH_4Cl\ (in\ g)}{molar\ mass\ of\ NH_4Cl\ (in\ g/mol)}$$
$$= \frac{10\,000\ g}{[(1\times 14)+(4\times 1)+(1\times 35.5)]\ g/mol}$$
$$= \frac{10000\ g}{53.5\ g/mol}$$
$$= 187\ \text{mol (nearest whole number)}$$

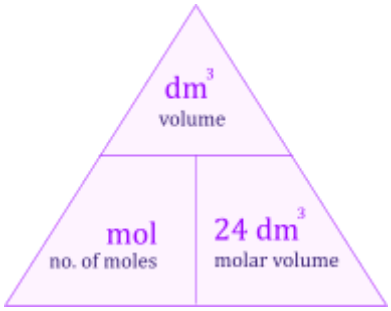
Mass and Volume of Gaseous Substances

Definition 7.1.6 *(Avogadro's Law)*

Avogadro's law states that at the same temperature and pressure, equal volumes of gases contains the same number of particles

Definition 7.1.7 *(Molar Volume)*

1 mol of any gas occupies 24 dm^3 at r.t.p and pressure.



$$\text{Number of moles} = \frac{Volume}{24}$$

Key Qns 7.1.3 *(Carbon Dioxide)*

Calculate the mass of 1500 cm^3 ***of carbon dioxide.***

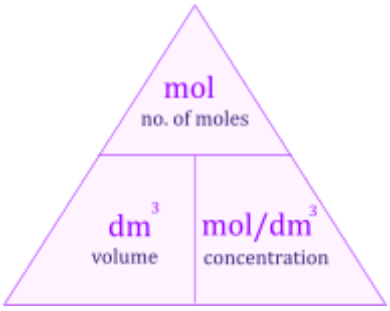
Number of moles of $CO_2 = \frac{1.5}{24} = 0.0625\ mol$

Mass of $CO_2 = 0.0625 \times (12 + 16 \times 2) = 2.75\ g$

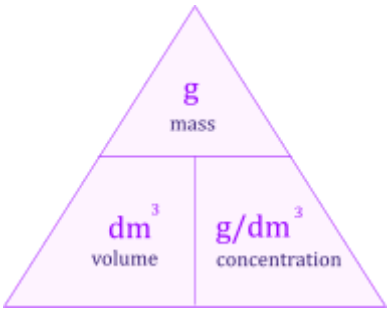
Concentration and Volume

Definition 7.1.8 *(Concentration)*

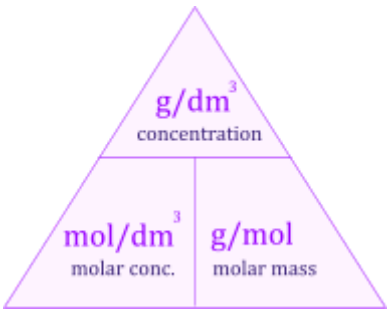
The concentration of a solution is the amount of a solute dissolved in a unit volume of the solvent.



$$\text{Molar Concentration } (mol/dm^3) = \frac{Number\ of\ moles\ of\ solute}{Volume\ of\ solution}$$



$$\text{Concentration } (g/dm^3) = \frac{Mass\ of\ solute}{Volume\ of\ solution}$$



$$\text{Concentration } (mol/dm^3) = \frac{Conc.\ (g/dm^3)}{Molar\ mass\ of\ solute}$$

Key Qns 7.1.4 *(Barium Chloride)*

Calculate the mass of barium chloride in $500 cm^3$ ***of*** $0.250\ mol/dm^3$ ***barium chloride***

No. moles of barium chloride $= 0.500 \times 0.250 = 0.125\ mol$

Mass of barium chloride $= 0.125 \times (137 + 35.5 \times 2) = 26.0\ g$

7.2 | Empirical Formula, Molecular Formula Stoichiometry (Chem Calc.)

Molecular and Empirical Formulas

Definition 7.2.1 *(Molecular Formula)*

The molecular formula shows the types of atoms present and the actual numbers of the atoms in the simplest unit of a covalent substance.

Definition 7.2.2 *(Empirical Formula)*

The empirical formula shows the simplest ratio of atoms of different elements present in a compound.

Key Qns 7.2.1 *(Deducing Empirical Formula From Mass)*

3.53 g of iron reacts with chlorine to form 10.24 g of iron chloride. Find the empirical formula for the iron chloride.

Element	Fe	Cl
Mass of element in iron chloride/g	3.53	6.71
Molar mass in g/mol	56	35.5
Number of moles/mol	$\frac{3.53}{56} = 0.063036$	$\frac{6.71}{35.5} = 0.18901$
Mole ratio	1	$\frac{0.18901}{0.063036} = 3$
Simplest ratio	1:3	
Empirical formula	$FeCl_3$	

Key Qns 7.2.2 *(Deducing Empirical Formula From Percentage)*

An oxide of titanium contains 60% by mass of titanium (Ti). What is the empirical formula for this oxide?

Element	Ti	O
Mass of element in 100g of titanium oxide/g	60	40
Molar mass in g/mol	48	16
Number of moles/mol	$\frac{60}{48} = 1.25$	$\frac{40}{16} = 2.5$
Mole ratio	1	$\frac{0.18901}{0.063036} = 2$
Simplest ratio	1:2	
Empirical formula	TiO_2	

Deducing Molecular Formula

Since,

$$Molecular\ Formula = (Empirical\ Formula)_n$$

We can easily find n, given the molecular mass

$$n = \frac{Molecular\ Mass}{Molar\ Mass\ of\ Empirical\ Formula}$$

Key Qns 7.2.3 *(Deducing Molecular Formula From Percentage)*

A particular compound comprises 87.5% nitrogen and 12.5% hydrogen by mass. Given that the relative molecular mass is 32, calculate the molecular formula of this compound.

Element	N	H
Mass of element in 100g of compound/g	87.5	12.5
Molar mass in g/mol	14	1
Number of moles/mol	$\frac{87.5}{14} = 6.25$	$\frac{12.5}{1} = 12.5$
Mole ratio	1	$\frac{0.18901}{0.063036} = 2$
Simplest ratio	1:2	
Empirical formula	NH_2	

Let the molecular formula be $(NH_2)_n$

$$n = \frac{32}{16} = 2$$

∴ The molecular formula of compound is N_2H_4, hydrazine.

Key Qns 7.2.4 *(Hydrated Copper(II) Sulfate)*

A sample of hydrated copper(II) sulfate weighs 124.8 g. The sample has been determined to contain 31.8 g of copper(II) ions and 48.0 g of sulfate ions.

a) How many molecules of water of crystallisation are present in the sample?

Mass of water of crystallisation in hydrated copper(II) sulfate: $124.8 - 31.8 - 48.0 = 45.0g$

Molar mass of water: $2 + 16 = 18\ g/mol$

Number of moles of water molecules: $\frac{45.0}{18} = 2.5\ mol$

Number of water molecules: $2.5 \times 6.02 \times 10^{23} = 1.51 \times 10^{24}$

b) Deduce the actual formula of hydrated copper(II) sulfate.

Moles of $CuSO_4 = \frac{31.8+48.0}{63.5+32+4\times16} = 0.50031$

	$CuSO_4$	H_2O
Number of moles/mol	0.50031	2.5
Mole ratio	1	5
Simplest ratio	1:5	

∴ The molecular formula of hydrated copper(II) sulfate is $CuSO_4.5H_2O$.

Stoichiometry

Stoichiometry is the relationship between the number of moles of reactants and the number of moles of products involved in a chemical reaction.

Key Qns 7.2.5 *(Hydrogen Sulfide)*

Hydrogen sulfide burns in an excess of oxygen according to the following equation.

$$2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(g) + 2SO_2(g)$$

If 0.68 g of hydrogen sulfide are burned completely in oxygen, **what mass of sulfur dioxide will be formed at room temperature and pressure?**

Moles of H_2S burned $= \frac{0.68}{34} = 0.02$

Comparing mole ratio,

$H_2S : SO_2$

1 : 1

Number of moles of SO_2 produced $= 0.02\ mol$

Mass of SO_2 produced $= 0.02 \times 64 = 1.28\ g$

***Key Qns 7.2.6** (Hydrated Magnesium Chloride Crystals)*

Magnesium chloride crystallises as a hydrate, $MgCl_2.xH_2O$***. When 20.3 g of the hydrate were heated to a constant mass, 9.5 g of the anhydrous salt, MgCl2, is formed according to equation:***

$$MgCl_2.xH_2O(s) \rightarrow MgCl_2(s) + xH_2O(g)$$

Find the value of x.

$$\text{Mass of } H_2O = 20.3 - 9.5 = 10.8\ g$$

$$\text{Number of moles of } H_2O = \frac{10.8}{18} = 0.6\ mol$$

$$\text{Number of moles of } MgCl_2 = \frac{9.5}{24+35.5\times2} = 0.1\ mol$$

Comparing mole ratio,

$$MgCl_2 : H_2O$$

$$1 \quad : \quad 6$$

$$\therefore x = 6$$

Gaseous Reactants (Volume Ratio)

***Key Qns 7.2.7** (Some Hydrocarbon)*

Methane gas reacts with steam to form hydrogen gas and carbon monoxide gas. Hence, calculate the volume of methane reacted when $100cm^3$ ***of hydrogen gas is produced in this reaction .***

$$CH_4(g) + H_2O(g) \rightarrow 3H_2(g) + CO(g)$$

Comparing mole ratio

$$H_2 : CH_4$$

$$3 : 1$$

Hence, by avogadro's law volume ratio $(H_2 : CH_4) = 3:1$

$$\therefore Volume\ of\ methane = \frac{1}{3} \times 100 = 33.3\ cm^3 (3sf)$$

Key Qns 7.2.8 *(Some Hydrocarbon)*

30 cm^3 *of hydrocarbon,* C_nH_{2n}*, burns in excess oxygen to form water vapour and* $120cm^3$ *of carbon dioxide. What is the formula of the hydrocarbon?*

$$wC_nH_{2n}(g) + xO_2 \rightarrow yH_2O(g) + zCO_2(g)$$

$$\text{Moles of } CO_2 \text{ produced} = \frac{0.120}{24} = 0.005\ mol = Moles\ of\ C\ atom\ in\ CO_2$$

$$\text{Moles of hydrocarbon} = \frac{0.03}{24} = 0.00125\ mol$$

$$\text{Moles of } C \text{ atom in hydrocarbon} = 0.00125n\ mol$$

$$0.00125n = 0.005 \Rightarrow n = 4$$

$$\therefore Formula\ of\ hydrocarbon\ is\ C_4H_8$$

Calculations Involving Concentrations (Eg. Titration)

Key Qns 7.2.9 *(Example of Calc. W/ Concentration)*

2.65 g of Na_2CO_3 were dissolved in water and made up to 250 cm^3 solution. 25.0 cm^3 of this solution required 20.0 cm^3 of nitric acid for complete neutralisation.

$$Na_2CO_3(aq) + 2HNO_3(aq) \rightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$$

Hence, find the concentration of nitric acid in g/dm^3.

$$\text{Moles of } Na_2CO_3 \text{ in } 25\ cm^3 = \frac{2.65}{106} \times \frac{25}{250} = 0.0025\ mol$$

Comparing mole ratio

$$Na_2CO_3 : HNO_3$$

$$1 : 2$$

$$\text{Number of moles of } HNO_3 \text{ required} = \frac{2}{1} \times 0.0025 = 0.005\ mol$$

$$\text{Concentration of } HNO_3 = \frac{0.005}{0.02} \times (1 + 14 + 48) = 15.8 g/dm^3$$

7.3 | Limiting Reactants, % Yield and % Purity (Chem Calc.)

Limiting Reactant

The reactant that is completely used up in this reaction is known as the *Limiting Reactant*.

Key Qns 7.3.1 *(Calc. Mass)*

When 6 g of hydrogen is allowed to react with 32 g of oxygen, water is produced. Identify the limiting reactant. Hence, find the mass of water formed.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

Number of moles of H_2 provided $= \frac{6}{1} = 6\ mol$

Number of moles of O_2 provided $= \frac{32}{16} = 2\ mol$

Comparing mole ratio,

$$H_2 : O_2$$

$$2 : 1$$

4 mol of H_2 is required to react with 2 mol of O_2. Since we are given 6 mol of H_2, H_2 is in excess and O_2 is the limiting reactant.

Comparing mole ratio,

$$O_2 : H_2O$$

$$1 : 1$$

Number of moles of H_2O produced $= \frac{1}{1} \times 2 = 2\ mol$

Mass of water produced $= 18 \times 2 = 36\ g$

Key Qns 7.3.2 *(Calc. Mass & Volume)*

When $4.8\ dm^3$ ***of hydrogen at r.t.p is passed over*** $67.0\ g$ ***of heated lead(II) oxide, lead and water vapour are produced.***

$$H_2(g) + PbO(s) \rightarrow Pb(s) + H_2O(g)$$

Calculate the mass of excess reactant remaining at the end of this reaction.

Moles of H_2 provided $= \frac{4.8}{24} = 0.2\ mol$

Moles of PbO provided $= \frac{67}{207+16} = 0.3\ mol$

H_2 is the limiting reactant, there will be $0.1\ mol$ of PbO remaining.

Mass of PbO remaining $= 0.1 \times (207 + 16) = 22.3\ g$

Calculate the mass of lead formed.

Comparing mole ratio,

$$H_2 : Pb$$

$$1 : 1$$

Hence, moles of Pb formed $= \frac{1}{1} \times 0.2 = 0.2\ mol$

Mass of Pb formed $= 0.2 \times 207 = 41.4\ g$

Key Qns 7.3.3 *(Calc. Volume)*

A student reacted $100\ cm^3$ ***of ammonia gas with*** $120\ cm^3$ ***of oxygen gas to produce nitrogen monoxide gas and water vapour.***

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

Calculate the volume of excess reactant remaining at the end of this reaction.

Volume of NH_3 provided $= 0.1\ dm^3$

Volume of O_2 provided $= 0.12\ dm^3$

Comparing mole ratio (and thus volume ratio)

$$NH_3 : O_2$$

$$4 : 5$$

To react with $0.12\ dm^3$ of O_2, $\frac{4}{5} \times 0.12 = 0.096\ dm^3$ of NH_3 is required. Since we are given $0.1\ dm^3$ of NH_3, NH_3 is in excess and O_2 is the limiting reactant.

Volume of unreacted $NH_3 = 0.01 - 0.096 = 0.004\ dm^3$

Calculate the volume of nitrogen monoxide formed.

Comparing volume ratio,

$$O_2 : NO$$

$$5 : 4$$

Thus, volume of NO formed $= \frac{4}{5} \times 0.12 = 0.096\ dm^3$

Key Qns 7.3.4 *(Calc. Conc. & Volume)*

When $25.0\ cm^3$ ***of*** $3.00\ mol/dm^3$ ***aluminium chloride solution is mixed with*** $75.0\ cm^3$ ***of*** $2.00\ mol/dm^3$ ***silver nitrate solution, aqueous aluminium nitrate and silver chloride precipitate are formed.***

$$AlCl_3(aq) + 3AgNO_3(aq) \rightarrow Al(NO_3)_3(aq) + 3AgCl(s)$$

Calculate the mass of silver chloride produced.

Moles of $AlCl_3$ provided $= 0.025 \times 3.00 = 0.075\ mol$

Moles of $AgNO_3$ provided $= 0.075 \times 2.00 = 0.150\ mol$

Comparing mole ratio,

$AlCl_3 : AgNO_3$

$1 : 3$

To completely react with $0.150\ mol$ of $AgNO_3$, $\frac{1}{3} \times 0.150 = 0.0500\ mol$ of $AlCl_3$ is required.

We are provided with $0.075\ mol$ of $AlCl_3$, hence, $AlCl_3$ is in excess and $AgNO_3$ is the limiting reactant.

Comparing mole ratio,

$AgNO_3 : AgCl$

$1 : 1$

Moles of $AgCl$ produces $= \frac{1}{1} \times 0.150 = 0.150\ mol$

Mass of $AgCl$ produced $= 0.150 \times (108 + 35.5) = 21.5\ g\ (3sf)$

Percentage Yield

The percentage yield is a measure of the efficiency of the reaction.

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

Percentage Purity

This is a measure of how pure a substance is. Sometimes, the reactants used in the chemical reaction are impure.

$$\text{Percentage purity} = \frac{Amount\ of\ pure\ sample\ in\ species}{Amount\ of\ species} \times 100\%$$

8.1 | Properties of Acids (Acids and Bases)

What are Acids?

Definition 8.1.1 *(Definition of Acids)*

A substance which produces hydrogen ions in aqueous solution.

Basicity

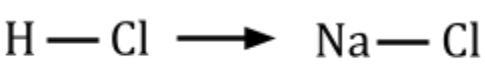
Definition 8.1.2 *(Basicity)*

Number of moles of hydrogen ions produced when 1 mole of acid molecules ionised in aqueous solution.

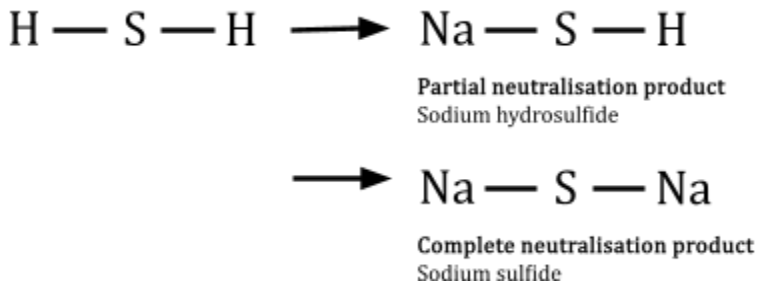
ⓘ **Note**

Basicity of an acid influences the number of salts it forms.

Monobasic acids (eg HCl)

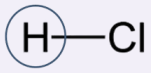
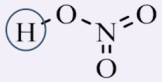
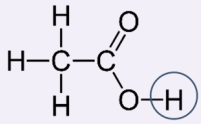
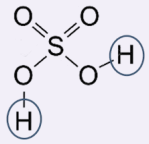
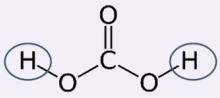
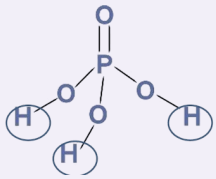


Dibasic acids (eg Hydrosulfic Acid)



Classification of Acids According to Basicity

Table 8.1.1. Acid classification (mono-, di-, tri-basic acids) as well as strength

Basicity	Examples	Chemical formula & Structural formula	Ionisation Equation
Monobasic	Hydrochloric acid	**HCl** 	$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
	Nitric acid	**HNO_3** 	$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$
	Ethanoic acid	**CH_3COOH** 	$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$
Dibasic	Sulfuric acid	**H_2SO_4** 	$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$
	Carbonic acid	**H_2CO_3** 	$H_2CO_3(aq) \rightleftharpoons 2H^+(aq) + CO_3^{2-}(aq)$
Tribasic	Phosphoric acid	**H_3PO_4** 	$H_3PO_4(aq) \rightleftharpoons 3H^+(aq) + PO_4^{3-}(aq)$

Strength of Acids

Definition 8.1.3 *(Strength of Acids)*

The extent of ionisation of acid molecules to produce hydrogen ions in aqueous solutions.

Definition 8.1.4 *(Strong Acids)*

A substance that ionises completely to produce hydrogen ions in aqueous solutions.

Definition 8.1.5 *(Weak Acids)*

A substance that ionises partially to produce hydrogen ions in aqueous solutions.

> ⓘ **Note**
> The double arrow ⇌ in the ionisation equation of acids indicate a weak acid.

Table 8.1.2. Comparison: Strong acids vs Weak acids

Strong Acids	Weak acids
A strong acid is one that **completely ionises** in water.	A strong acid is one that **partially ionises** in water. Usually only **~1%** of acid molecules split up to form ions.
A solution of HCl contains only • H^+ ions • Cl^- ions • H_2O molecules	A solution of CH_3COOH contains only • H^+ ions • CH_3COO^- ions • CH_3COOH molecules • H_2O molecules
Thus, for the same **basicity, volume and concentration**, strong acids have a much higher concentration of hydrogen ions in aqueous solution of strong acids compared to weak acids.	Thus, for the **same basicity, volume and concentration**, weak acids have a much lower concentration of hydrogen ions in solution compared to strong acids.
This explains why the rate of reaction of strong acids with substances is faster than that of weak acids. Strong acids react more vigorously than weak acids. **Observations** Reaction mixture involving strong acids will be **hot to touch** while reaction mixtures involving weak acids will be **warm to touch**.	

Increasing Concentration of Acids

- Increasing mass of acid
- Decreasing mass of water

Property #1: Acids have a sour taste

Property #2: Acids are good electrical conductors

Explain why dilute hydrochloric acid is a good conductor of electricity at r.t.p. [1]

The H^+ and Cl^- ions in dilute hydrochloric acid are mobile hence these ions can move freely throughout the substance and act as charge carriers. [1]

Explain why hydrogen chloride gas cannot conduct electricity but once the gas is dissolved in water, the aqueous solution can conduct electricity. [2]

- Hydrogen chloride gas exists as simple covalent HCl molecules that are electrically neutral. Hydrogen chloride gas does not have freely moving ions or delocalised electrons to act as charge carriers to conduct electricity.
- When hydrogen chloride gas dissolves in water, an aqueous solution of hydrochloric acid is formed. In aqueous state, HCl molecules ionise in water to form H^+ and Cl^- ions which are mobile. These ions can act as charge carriers and conduct electricity. (Refer: *pg 34*)

Reactions of Acids (Chemical Properties)

$reactive\ metal\ +\ dilute\ acid \rightarrow salt\ +\ hydrogen\ gas$

$carbonate\ +\ dilute\ acid \rightarrow salt\ +\ carbon\ dioxide\ gas\ +\ water$

$hydrogen\ carbonate\ +\ dilute\ acid \rightarrow salt\ +\ carbon\ dioxide\ gas\ +\ water$

$metal\ oxide\ +\ dilute\ acid \rightarrow salt\ +\ water$

$metal\ hydroxide\ +\ dilute\ acid \rightarrow salt\ +\ water$

Reactive Metal and Dilute Acids

Observation: A colourless, odourless gas is formed.

Hydrogen gas is insoluble in water. Thus hydrogen gas produced in the reaction will not dissolve in the aqueous solution of the reaction mixture.

> ⓘ **Note**
> Based on the *reactivity series,* copper, silver, gold and platinum will **NOT** react with acids.

Why do some reactions seem to stop after a while?

Lead is a reactive metal which will react with dilute hydrochloric acid to produce lead(II) chloride which is an insoluble salt. This insoluble salt quickly forms a layer around the metal which protects the metal from further reaction with the acid. Thus, reaction stops abruptly.

Solubility Rules of Salts

- All group 1 metal salts and ammonium salts are **soluble in water**
- All nitrates are **soluble in water**
- Most sulfates are **soluble in water** except **lead(II) sulfate** and **calcium sulfate** and **barium sulfate**
- Most carbonates are **insoluble in water** except **group 1 metal carbonates** and **ammonium carbonate**
- All **lead(II)** and **silver** halides are insoluble in water, all other halides are soluble in water.

Carbonates and Dilute Acids

(Neutralisation)

Observation: A colourless, odourless gas is formed.

Carbon dioxide gas is sparingly soluble in water. Thus most of the carbon dioxide gas produced in the reaction will not dissolve in the aqueous solution of the reaction mixture.

General ionic equation

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

Reaction between Alkalis and Dilute Acids

(Neutralisation)

General ionic equation

$$H^{+}(aq) + OH^{-}(aq) \rightarrow H_2O(l)$$

8.2 | Properties of Bases (Acids and Bases)

What Are Bases?

Definition 8.2.1 *(Base)*

Any substance which reacts with acid to produce salt and water.

Types of bases

- Metal oxides
- Metal hydroxides
- Carbonates
- Hydrogen carbonates
- Aqueous ammonia

Property #1: Bases have high melting and boiling points

Explain, in terms of structure, why calcium oxide is a solid at r.t.p. [3]

- Calcium oxide is an ionic compound with a giant ionic lattice structure. [1]
- In this structure, the calcium ions and oxide ions are held together by strong electrostatic forces of attraction (ionic bonds). [1]
- A lot of energy is required to overcome these strong forces to melt CaO. Hence CaO has a high melting point and is a solid at r.t.p.

Property #2: Only some bases are soluble in water

These bases are known as *alkalis.* All **group 1 and last four group 2** metal oxides and hydroxides, as well as ammonia are alkalis.

Metal oxides react in water to form an aqueous solution of the corresponding metal hydroxide.

$$K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$$

Property #3: Bases are good conductors of electricity when molten and in aqueous solutions.

Explain why aqueous barium hydroxide is a good conductor of electricity at r.t.p. [1]

- The Ba^{2+} and OH^- ions in aqueous barium hydroxide are freely moving hence these ions can move about throughout the substance and act as charge carriers.

Explain why white solid of barium oxide cannot conduct electricity but once the white solid dissolved in water, the colourless aqueous solution can conduct electricity. [2]

- In solid state, Ba^{2+} and OH^- ions are held in fixed positions by strong ionic bonds. Hence these ions are not mobile and cannot act as charge carriers to conduct electricity. [1]
- Barium oxide dissolves in water to form barium hydroxide solution. [1]
- In this solution, Ba^{2+} and OH^- ions are mobile hence these ions can move freely throughout the solution and act as charge carriers. [1]

Chemical reactions of bases with acids

$$base + dilute\ acid \rightarrow salt + water$$

Alkalis

Definition 8.2.2 *(Alkalis)*

A substance which produces hydroxide ions in aqueous solution.

Strength of Alkalis

Definition 8.2.3 *(Strength of Alkalis)*

The extent of ionisation of alkali to produce hydroxide ions in aqueous solutions.

Definition 8.2.4 *(Strong Alkalis)*

A substance which ionises completely to produce hydroxide ions in aqueous solutions.

Definition 8.2.5 *(Weak Alkalis)*

A substance which ionises partially to produce hydroxide ions in aqueous solutions.

Examples of Weak Alkalis

- Ammonia
- Methylamine (organic derivative of ammonia)

Table 8.2.1. Comparison: Strong alkali vs Weak alkali

Strong Alkalis	Weak Alkalis
A strong alkali is one that **completely ionises** in water.	A weak alkali is one that **partially ionises** in water. Usually only **~1%** of alkali molecules split up to form ions.
A solution of KOH contains only • K^+ ions • OH^- ions • H_2O molecules	A solution of NH_3 contains only • NH_4^+ ions • OH^- ions • NH_3 molecules • H_2O molecules
Thus, for the same **volume and concentration**, strong alkalis have a much higher concentration of hydroxide ions in aqueous solution of alkalis.	Thus, for the **same volume and concentration**, weak alkalis have a much lower concentration of hydroxide ions in solution compared to strong alkalis.

Property #1: Alkalis have a bitter taste

Property #2: Alkalis feel soapy

Property #3: Aqueous solution of alkalis are good electrical conductors

Property #4: Alkalis are soluble in water

Chemical Properties of Alkalis (Chemical Reactions)

$alkali + dilute\ acid \rightarrow salt + water$

$alkali + ammonium\ salt \rightarrow salt + ammonia\ gas + water$

$alkali + aqueous\ solution\ of\ metal\ B\ salt \rightarrow metal\ B\ hydroxide + salt$

Reactions of alkalis and dilute acids

(Refer to acids section)

Alkalis and Ammonium Salts

Observation: A gas with a pungent smell is evolved.

Ammonia gas is soluble in water. Thus ammonia gas produced in the reaction dissolves in the aqueous solution of the reaction mixture. (No effervescence)

Alkalis and Aqueous Solution of Metal Salts

Alkalis react with some aqueous metal salts to give an insoluble metal hydroxide. This is known as a *precipitation* reaction.

Table 8.2.3. Applications of bases/alkalis

Aqueous ammonia	Window cleaning solutions Making nitrogenous fertilisers
Aqueous sodium hydroxide	Detergents to dissolve grease
Calcium hydroxide	Neutralise acidic soil Neutralise acid spills
Magnesium hydroxide	In toothpaste to neutralise acid on teeth In antacids to neutralise excess stomach acid
Zinc oxide	In calamine solution to neutralise bee stings

8.3 | pH Scale (Acids and Bases)

Water As a Solvent

Water undergoes autoionisation as such

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

This is a reversible reaction. A small proportion of water molecules break up to form hydrogen ions and hydroxide ions. Some of these hydrogen and hydroxide ions then react together again to form water molecules.

Thus, in water and neutral solutions, the concentration of hydrogen ions is equal to the concentration of hydroxide ions. These substances also contain water molecules.

Relationship Between Conc. of Hydrogen & Hydroxide Ions and pH

The lower the pH, the more acidic the solution is. This is because the concentration of H^+ ions are higher than the concentration of OH^- ions.

The higher the pH, the more alkaline the solution is. This is because the concentration of OH^- ions are higher than the concentration of H^+ ions.

Table 8.3.1. Some common indicators

Indicator	Approx. pH range for colour change	Colour change (acid → alkali)
Methyl orange	3.2-4.4	red to yellow
Screened methyl orange	3.8-4.2	violet to green
Thymolphthalein	9.3-10.5	colourless to blue
Litmus	4.5-8.3	red to blue

pH Meter/datalogger connected to a pH sensor

This is an electrical device which is more accurate than indicators. It requires calibration and is dipped into the solution directly to measure pH. The chemical indicators give only approximate pH values. To obtain **more accurate pH values**, we can use the pH meter.

substances	strong acids	weak acids	neutral aqueous solution & pure water	weak alkalis	strong alkalis
an example of solution	dilute hydrochloric acid	dilute ethanoic acid	salt solutions (e.g. sodium chloride solution)	aqueous ammonia	sodium hydroxide solution
extent of ionisation of acid/alkali	complete	partial	-	partial	complete
concentration of H^+ ions	high	lower than concentration of H^+ ions in strong acids but higher than concentration of H^+ ions in alkalis	concentration of H^+ ions is the same as concentration of OH^- ions	higher than concentration of H^+ ions in strong alkalis but lower than concentration of H^+ ions in acids	low
concentration of OH^- ions	low	higher than concentration of OH^- ions in strong acids but lower than concentration of OH^- ions in alkalis		lower than concentration of OH^- ions in strong alkalis but higher than concentration of OH^- ions in acids	high
presence of water molecules	yes	yes	yes	yes	yes
pH range of solution	$0 \leq pH \leq 2$	$2 < pH < 7$	pH 7 (no range)	$7 < pH < 12$	$12 \leq pH \leq 14$
suggested pH value of solution	pH 1	pH 4	pH 7	pH 10	pH 14
Observation made when a few drops of Universal Indicator solution was added to the solution.	Green Universal Indicator turned red.	Green Universal Indicator turned orange.	No visible change. Universal Indicator remained green.	Green Universal Indicator turned blue.	Green Universal Indicator turned violet.

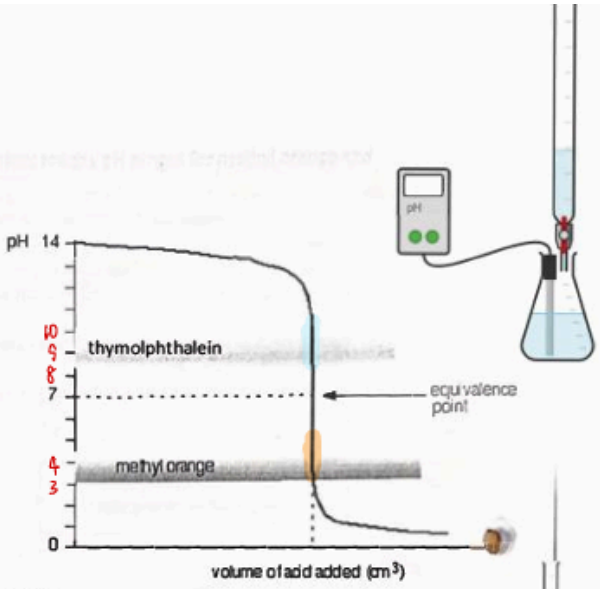
Fig 8.3.1. Colour change of universal indicator on various solutions

(1)STRONG ACID REACTING WITH STRONG ALKALI

The graph shows the pH changes when a **strong acid** (HCl) is added to a **strong alkali** (NaOH)
Evidence :
At start of titration [0 cm^3] pH of solution is **14**
At the end of titration the pH of reaction mixture is approximately **1**

Both thymolphthalein and methyl orange are suitable indicators

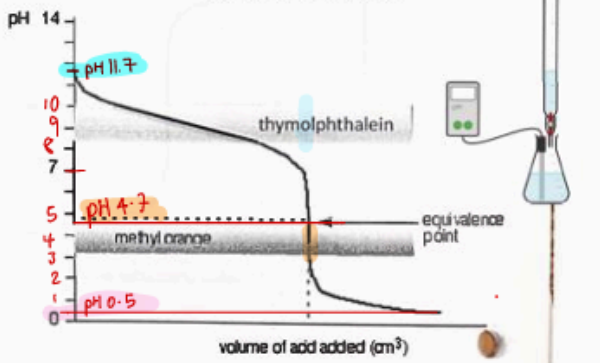
Notice the equivalence point is at pH **7**
Hence the salt formed has pH **7**



(2)STRONG ACID REACTING WITH WEAK ALKALI

The graph shows the pH changes when a strong acid is added to a weak alkali

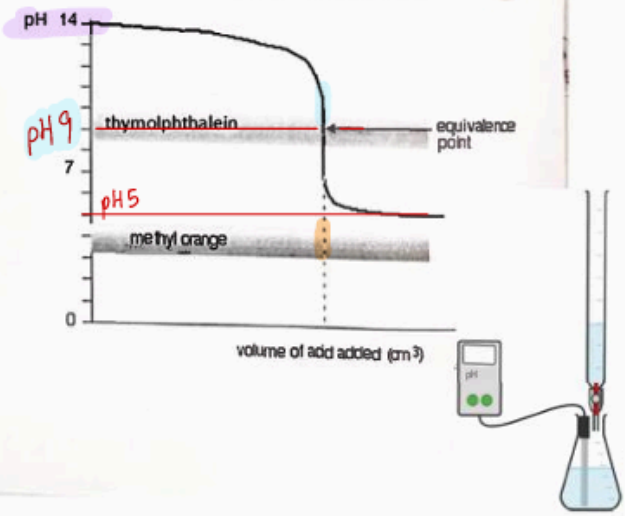
Thymolphthalein cannot be used to indicate the end-point of the reaction.
A suitable indicator is methyl orange
Notice the equivalence point is at pH 4.7
Hence the salt formed has pH 4.7



(3)WEAK ACID REACTING WITH STRONG ALKALI

The graph shows the pH changes when a weak acid is added to a strong alkali

Methyl orange cannot be used to indicate the end-point of the reaction.
A suitable indicator is thymolphthalein
Notice the equivalence point is at pH pH 9
Hence the salt formed has pH pH 9



Key Qns 8.3.1 *(pH change in titration of NaOH and HCl)*

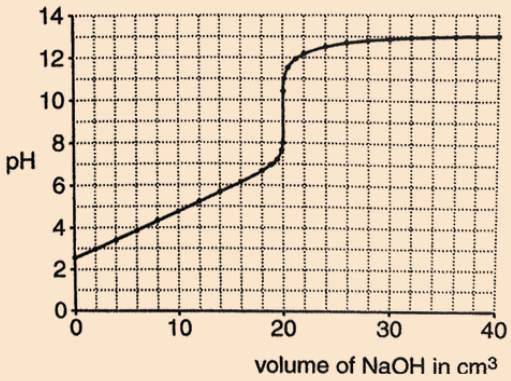
0. 1 mol/dm^3 ***sodium hydroxide, NaOH, is added from a burette into a solution of HCl. Explain why the pH does not change smoothly in the titration.***

At initial point, pH is low as only acid is present in the conical flask. When one drop of NaOH is added from the burette, the concentration of HCl in the conical flask did not change much as only one drop equivalent of HCl reacted with the NaOH. As more NaOH is added, the pH will remain below 7 as H^+ is still present in the conical flask [1].

Only when an equivalent amount of NaOH and HCl is added, complete neutralisation occurs, and the pH increases to pH 7 [1].

As more NaOH is added, there is no more acid present to react with the NaOH and there is excess NaOH. As a result, pH will increase from 7 to 13.

Key Qns 8.3.2 *(pH change in titration of NaOH and HCl)*



0. 1 mol/dm^3 ***sodium hydroxide, NaOH, is added from a burette into a solution of ethanoic acid. Explain the shape of the graph.***

At the initial point (0 cm^3 of $NaOH$), there is only ethanoic acid present. Ethanoic acid is a weak acid, it does not dissociate completely, and thus pH is at 2.5. When some NaOH is added, the OH^- ions reacted with H^+ ions, causing the pH to increase. When an equivalent amount of $NaOH$ and ethanoic acid is added, the equivalence point is at pH 9 and neutralisation occurred [1].

$NaOH$ reacted with CH_3COOH to give CH_3COONa and water. When more $NaOH$ is added, there is no more ethanoic acid present to react with the NaOH, and thus, pH is 13 due to the excess NaOH [1].

Key Qns 8.3.3 *(Indicators)*

Explain why methyl orange would not be suitable to use when titrating ethanoic acid with dilute sodium hydroxide.

When ethanoic acid is titrated with dilute sodium hydroxide, neutralisation occurs between pH 7.5 and pH 11. Methyl orange will not be able to detect this change in pH as it changes colour at between pH 3 and pH 5 before the neutralisation reaction occurs.

8.4 | Types of Oxides (Acids and Bases)

What Is An Oxide?

Definition 8.4.1 *(Oxides)*

A binary compound of oxygen and another element.

Table 8.4.1. Types of Oxides

Metal oxides		Nonmetal oxides	
Basic oxides	**Amphoteric oxides**	**Acidic oxides**	**Neutral oxides**
Most metal oxides	Zinc oxide, aluminium oxide, lead(II) oxide	Most nonmetal oxides	Carbon monoxide, water, nitric oxide
Basic properties (Refer: 8.2) **Water-soluble basic oxides will dissolve in water to form alkalis.**	Both basic and acidic properties Reacts like a base when mixed with acid, reacts like an acid when mixed with base. **Insoluble in water.**	Acidic properties (Refer: 8.1) **Water-soluble acidic oxides will dissolve in water to form acid solutions.**	Generally do not react with acids and bases. **Exception:** Water reacts with water-soluble metal oxides to form metal hydroxide solution. Water reacts with acidic oxides to form aqueous solution of acids.

Basic Oxides

Definition 8.4.2 *(Basic Oxides)*

A compound of oxygen and a metal. This compound reacts with acids to form salt and water.

General word equation

$$metal\ (s)\ +\ oxygen\ (g) \rightarrow basic\ oxide\ (s)$$

General neutralisation word equation

$$basic\ oxide\ +\ dilute\ acid \rightarrow salt\ +\ water$$
$$basic\ oxide\ +\ acidic\ oxide \rightarrow salt$$

Amphoteric Oxides

Definition 8.4.3 *(Amphoteric Oxides)*

A compound of oxygen and a metal. This compound reacts with acids and bases to form salt and water.

General word equation

$$metal\ (s)\ +\ O_2(g) \rightarrow amphoteric\ metal\ oxide$$

Reaction as a base

$$amphoteric\ oxide\ +\ dilute\ acid \rightarrow salt\ +\ water$$
$$amphoteric\ oxide\ +\ acidic\ oxide \rightarrow salt$$

Reaction as an acid

$$amphoteric\ oxide\ +\ metal\ hydroxide \rightarrow salt\ +\ water$$
$$amphoteric\ oxide\ +\ basic\ oxide \rightarrow salt$$

Example

Zinc oxide reacts with aqueous sodium hydroxide to produce aqueous sodium zincate and water

$$ZnO(s)\ +\ 2NaOH(aq) \rightarrow Na_2ZnO_2(aq)\ +\ H_2O(l)$$

Amphoteric Oxides

Definition 8.4.4 *(Acidic Oxides)*

A compound of oxygen and a nonmetal. This compound reacts with bases to form salt and water.

General word equation

$$nonmetal\ (g)\ +\ oxygen\ (g) \rightarrow acidic\ nonmetal\ oxide\ (s)$$

General neutralisation word equation

$$acidic\ oxide\ +\ metal\ hydroxide \rightarrow salt\ +\ water$$
$$acidic\ oxide\ +\ basic\ oxide \rightarrow salt$$
$$acidic\ oxide\ +\ amphoteric\ oxide \rightarrow salt$$

Neutral Oxides

Definition 8.4.5 *(Neutral Oxides)*

A compound of oxygen and a non-metal. This compound does not show properties of acids and bases.

General word equation

$$nonmetal\ (g)\ +\ oxygen\ (g) \rightarrow neutral\ nonmetal\ oxide\ (s)$$

> ⓘ **Note**
> Water reacts with water-soluble metal oxides to form metal hydroxide solution.
> Water reacts with acidic oxides to form aqueous solution of acids.

9.1 | Solubility of Salts (Salts)

What are salts?

Definition 9.1.1 *(Salts)*

A salt is an ionic compound that consists of a cation and an anion.

> ⓘ **Note**
> For anion, it can be any anion **EXCEPT** oxide and hydroxide ions.

Solubility Rules of Salts

- All group 1 metal salts and ammonium salts are **soluble in water**
- All nitrates are **soluble in water**
- Most chlorides are **soluble in water** except **lead(II) chloride** and **silver chloride**
- Most sulfates are **soluble in water** except **lead(II) sulfate** and **calcium sulfate** and **barium sulfate**
- Most carbonates are **insoluble in water** except **group 1 metal carbonates** and **ammonium carbonate**
- All **lead(II)** and **silver** halides are insoluble in water

9.2 | Preparation of Salts (Salts)

Table 9.2.1. Main methods to prepare salts

Reaction of acids with excess insoluble substances (metals, insoluble carbonates and bases)	Most soluble salts except group 1 metal salts and ammonium salts.
Titration	Soluble salts of all group 1 metals and ammonium salt.
Precipitation	Insoluble salts.

Reaction of acids with excess insoluble substances

Most soluble salts except group 1 metal salts and ammonium salts.

- Reaction of acids and reactive metals (Cannot use copper, silver, gold and platinum)
- Reaction of acids and insoluble bases
- Reaction of acids and insoluble carbonates

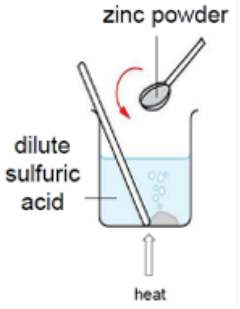
(Refer: *8.1*)

Describe how you would prepare a pure dry sample of zinc sulfate from the reaction of an acid with a metal, using all the separation techniques you would use.

Why add excess insoluble substance?

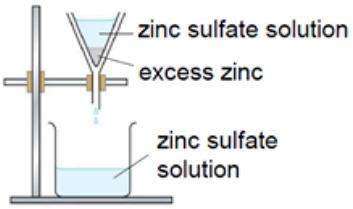
To ensure that all dilute acid has reacted completely. Filtrate obtained after filtration of reaction mixture is salt solution that is not contaminated with unreacted dilute acid.

1. Add excess Zn powder into dilute H_2SO_4 acid. Warm the mixture. With constant stirring, add zinc powder until no more reaction occurs with the acid (no more zinc powder could dissolved in the acid).



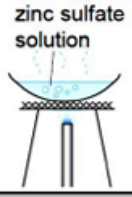
separation technique: **Filtration**

2. Filter to remove excess (unreacted) Zn powder. Collect the filtrate, $ZnSO_4$ solution.

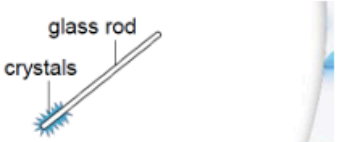


Why heat filtrate?

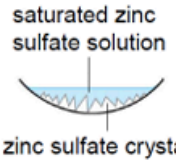
Heat to concentrate filtrate to obtain a saturated salt solution. On heating filtrate, water in filtrate will evaporate on heating.

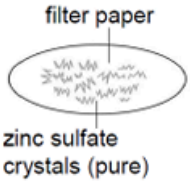


3. Heat filtrate to obtain a saturated $ZnSO_4$ solution.



Test for saturation of solution. The $ZnSO_4$ solution is saturated when crystals form on glass rod.





purification technique: **Crystallisation**

4. Leave the saturated $ZnSO_4$ solution to cool and crystallise.

5. Filter to obtain residue which is the $ZnSO_4$ crystals.
6. Rinse $ZnSO_4$ crystals with cold, distilled water to remove impurities. Dry crystals between sheets of filter paper.

Titration

Soluble salts of all group 1 metals and ammonium salt.

- Reaction of acids and alkali
- Reaction of acids and soluble carbonates

Let's Investigate 9B

Aim

To prepare sodium nitrate from the titration of aqueous sodium hydroxide with dilute nitric acid

Procedure

This investigation involves two parts:

(a) titration to determine the volumes of reactants required

(b) preparation of a pure sample of sodium nitrate

(a) Titration

1 Fill a burette with dilute nitric acid. Note the initial burette reading (V_1 cm³).
2 Pipette 25.0 cm³ of aqueous sodium hydroxide into a conical flask.
3 Add one to two drops of methyl orange (indicator) to the aqueous sodium hydroxide. The solution turns yellow (Figure 9.8).
4 While swirling the conical flask, add dilute nitric acid from the burette slowly until the solution just turns orange permanently. This is the endpoint. Record the final burette reading (V_2 cm³).
5 Find the volume of dilute nitric acid required for complete neutralisation, which is ($V_2 - V_1$) cm³.

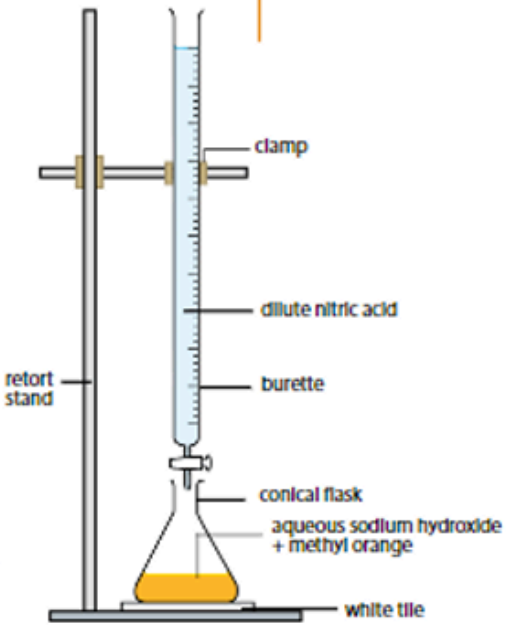


Figure 9.8 Titration of aqueous sodium hydroxide with dilute nitric acid

(b) Preparation of a pure sample of sodium nitrate

1 Pipette 25.0 cm³ of aqueous sodium hydroxide into a conical flask. Then, add ($V_2 - V_1$) cm³ of dilute nitric acid from the burette. Do not add the indicator as it will make the salt impure.
2 Pour the solution into an evaporating dish.
3 Heat the solution until it is saturated.
4 Allow the saturated solution to cool so that the salt can crystallise.
5 Filter to collect the crystals. Then, wash the crystals with a little cold distilled water to remove impurities. Dry the crystals between a few sheets of filter paper.

Precipitation (Double-Displacement Reactions)

Preparation of insoluble salts.

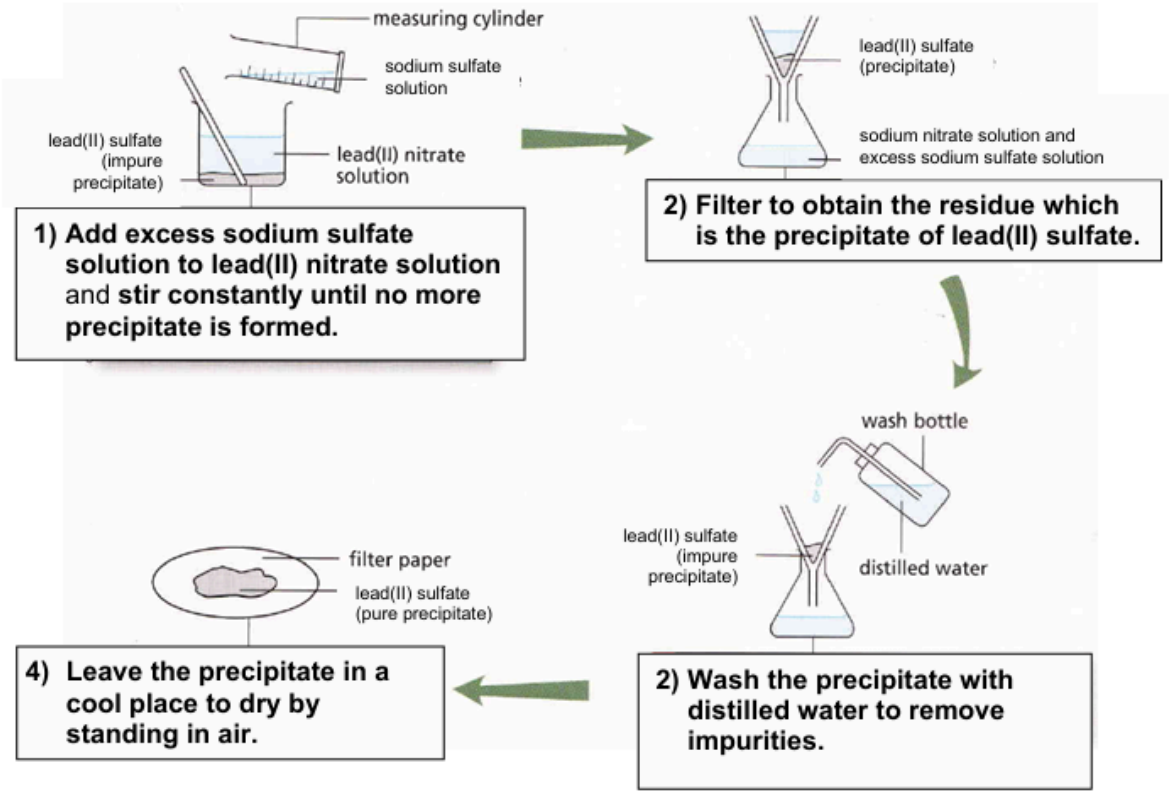
- Reactions must be in aqueous solutions (both reactants must be soluble in water)
- Reaction produces two products
    - Solid ionic compound (precipitate)
    - The other product is made up of *spectator ions*

Example

Precipitation of barium sulfate

$$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + NaNO_3(aq)$$

Describe how you would prepare a pure dry sample of lead(II) sulfate from the *reaction of 2 aqueous solutions of salts*, using all the separation techniques you would use.



14.1 | Periodic Trends (Periodic Table)

Periodic Trends

Down a group	Across a period (left to right)
Increase atomic number	Increase atomic number
Increase number of e^- shells	Number of e^- shells stays the same
Decreasing electrostatic forces of attraction between nucleus and valence electrons	Increasing electrostatic forces of attraction between nucleus and valence electrons
Increasing atomic radius	Decreasing atomic radius
Increase ease of losing electrons for metals	Decrease ease of losing electrons for metals
Decrease ease of gaining electrons for nonmetals (decreased electronegativity)	Increase ease of gaining electrons for nonmetals (increased electronegativity)

14.2 | Properties of Group 1, 17, 18 elements (Periodic Table)

Physical Properties of Group 1 (Alkali) Metals

- Shiny, silvery solids
- Soft and can be cut easily
- Good electrical conductors
- Good thermal conductors
- Malleable and ductile
- Lower densities (compared to transition metals)
    - Li, Na and K can float on water
- Low melting and boiling points

Trends In Physical Properties

Melting points and boiling points of alkali metals **decreases** down the group

- Down Group 1, there is an increase in the number of electron shells. Hence, the valence electrons are further away from the nucleus. Thus, the electrostatic forces of attraction between the nucleus and the valence electron are weaker.
- As a result, down Group 1, the sea of valence electrons would be delocalised more readily and move about more freely. Lesser heat will be required to overcome the weaker electrostatic forces of attraction between the nucleus and the sea of delocalised negatively charged valence electrons. Hence, melting and boiling points decrease.

Density of alkali metals **increases** down the group

- This is due to increase in atomic mass (more protons and neutrons)

Reactivity

The alkali metals are **very reactive**. Thus, they must be **stored in oil** to prevent them from reacting with oxygen in air and water.

Key Qns 14.2.1 *(Reactivity)*

Using atomic structure, explain why alkali metals are highly reactive.

Atoms of these elements have 1 valence electron. When group 1 elements react, their atoms lose their valence electrons very readily to form positively charged ions which have stable electronic configuration. Thus these elements are highly reactive.

Table 14.2.1. The chemical properties of alkali metals

React with water to form metal hydroxide solution and hydrogen gas

alkali metal (s) + water (l) → metal hydroxide (aq) + hydrogen gas (g)

Alkali metals react vigorously with water to form aqueous solutions, this is an exothermic reaction and highly flammable hydrogen gas is liberated. The heat released may ignite hydrogen gas, resulting in fires.

Metal	Observations (when placed in water)
Lithium	Reacts **quickly** with water Lithium floats on water No flame is seen
Sodium	Reacts **violently** with water Sodium melts to form a silvery white globule The molten sodium darts around the water surface It catches fire A yellow flame is seen
Potassium	Reacts **explosively** with water Potassium melts It catches fire A lilac flame is seen.

React with dilute acid

alkali metal (s) + dilute acid (aq) → salt (aq) + hydrogen gas (g)

Displacement reaction

alkali metal (s) + compound of less reactive metal (aq) → *less reactive metal (s) + compound of alkali metal (aq)*

Group 1 metals react with compounds of less reactive metals to form less reactive metal and compound of group 1 metal.

	Reasoning Group 1 metals are very reactive as they lose their single valence electron more readily than other metals forming positive ions. Hence, they have *low ionisation energy.* Thus, they can displace less reactive metals from their compounds.
React with oxygen to form metal oxides	*alkali metal (s) + oxygen (g) → metal oxide (s)* **Metal** / **Observations (when burned in oxygen)** Lithium — Burns with a red flame to give a **white solid**. Sodium — Burns with a bright yellow flame to produce a **white solid**. Potassium — Burns violently with a lilac flame to produce a **white solid**.

Trends In Reactivity

Reactivity increases down group 1.

- Down the group, atomic size increases due to increasing number of electron shells, thus electrostatic forces of attraction between the negatively charged valence electrons and the positively charged nucleus decreases.
- Hence, it is easier for bigger metal atoms to lose the valence electron to form a positively charged ion which has the stable electronic configuration of a noble gas.

Group 17 Elements (Halogens)

They are elements which exist as diatomic molecules and are nonmetals.

Physical Properties of Halogens

- Poor conductors of electricity
- Low melting and boiling points
- Coloured substances
- Some halogens are soluble in water while others are not

Trends In Physical Properties

Down group 17

- Melting and boiling points increases
- Colours become darker (increased colour intensity)
- Solubility of halogens in water decreases

Table 14.2.2. The physical properties of halogens

Halogens have low melting and boiling points	*(Refer to bonding and structure)* ***Use bonding and structure to explain why halogens have low melting and boiling points.*** • Halogens are covalent substances which exist as simple covalent molecules. • The molecules are held together by weak intermolecular forces of attraction. • A small amount of thermal energy is required to overcome these weak forces to melt and boil the halogens. Thus, halogens have low melting and boiling points. ***Down the group, the melting/ boiling points of halogens increase.*** • Down Group 17, there is an increase in the number of electron shells. Hence, the halogen molecules are larger in size. • Thus, there are stronger intermolecular forces of attraction between the larger halogen molecules. • Hence, a larger amount of thermal energy is required to overcome these stronger forces. Thus, going down Group 17, the melting and boiling point increases.
Different physical states at r.t.p and different colours	<table><tr><th>Metal</th><th>Physical state (@r.t.p.)</th><th>Colour</th></tr><tr><td>Fluorine</td><td>Gas</td><td>Pale yellow</td></tr><tr><td>Chlorine</td><td>Gas</td><td>Yellow-green</td></tr><tr><td>Bromine</td><td>Liquid</td><td>Red-brown</td></tr><tr><td>Iodine</td><td>Solid</td><td>Purple-black</td></tr><tr><td>Astatine</td><td>Solid</td><td>Black</td></tr></table> The elements become **darker** down group 17.

Reactivity

Definition 14.2.1 *(Reactivity of Non-Metals)*

Tendency of an atom of non-metal to gain one or more electron(s) to form its negative ion.

Hence, atoms of very reactive non-metals such as halogens have a high tendency to **gain electrons** to form negative ions (halide ions). Thus, halogens are **powerful oxidising agents.**

Table 14.2.3. The chemical properties of halogens

Halogen displacement reaction	*more reactive halogen* + *compound of less reactive halogen* → *less reactive halogen* + *compound of more reactive halogen* The more reactive halogen has a greater tendency to form negative ions compared to a less reactive halogen. Thus, a more reactive halogen displaces a less reactive halogen from its halide compound to form a halide compound of the more reactive halogen. **Example** Displacement reaction involving bromine and iodine. $$Br_2(aq) + 2KI(aq) \rightarrow 2KBr(aq) + I_2(aq)$$ ***What will be observed when the displacement reaction of bromine with aqueous potassium iodide? Explain your reasoning.*** • Colourless solution turns dark brown. • Bromine is a more reactive halogen than iodine. • Thus a displacement reaction will take place. Bromine will displace iodine from a colourless solution of aqueous potassium iodide to form a colourless solution of potassium bromide and dark brown solution of iodine. A black solid deposit is formed. The black deposit is solid iodine as iodine is sparingly soluble in water.
React with most metals for form halides	*halogen + metal → halide*
React with hydrogen gas to form hydrogen halides	*halogen + hydrogen gas → hydrogen halide*

Trends In Chemical Properties

Reactivity of halogens decreases down group 17.

- Down the group, atomic size increases due to increasing number of electron shells, thus electrostatic forces of attraction between the negatively charged valence electrons and the positively charged nucleus decreases.
- Hence, it is harder for bigger halogen atoms to attract one more valence electron to form a negatively charged ion which has the stable electronic configuration of a noble gas.

Group 18 Elements (Noble Gases)
They exist as monoatomic non-metals.

Physical Properties

- Monatomic nonmetals
- Low melting and boiling points
- Colourless gas at room temperature
- Insoluble in water

Table 14.2.4. Uses of noble gases

Argon	Argon is used to fill light bulbs because it will not react with the white-hot tungsten filament.
Helium	Helium is used to fill weather or advertisement balloons.
Neon	Neon is extensively used in advertising strip lights (neon lights).

14.3 | Transition Elements (Periodic Table)

Transition Elements — Where Are They Found?
They are found in the "d-block" section of the periodic table.

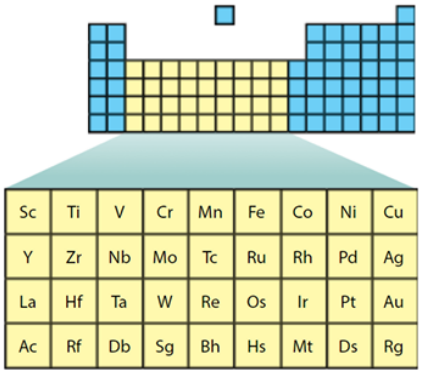


Fig 14.3.1. Where transition elements are found.

Table 14.3.1. The physical properties of transition metals

Shiny (lustrous) solids	**Exception:** Copper is a red-brown solid.
Good conductors of electricity	(Refer to metallic bonding) ● Transition elements have a giant metallic lattice structure. ● In this structure, positively charged transition metal ions are surrounded by a sea of negatively charged delocalised electrons. ● When a potential difference is applied, these delocalised electrons are able to move from the negative terminal through the metal wires and the metal to the positive terminal of the battery.
Good thermal conductors	(Refer to metallic bonding) ● Transition elements have a giant metallic lattice structure. In this structure, positively charged transition metal ions are surrounded by a sea of negatively charged delocalised electrons. ● The delocalised electrons are able to move throughout the giant metallic lattice structure and hence carry the heat through the metal.
Malleable and Ductile	(Refer to metallic bonding) ● Pure metal is made up of metal atoms which have the same atomic size. ● When force is applied to the metal, the layers of metal atoms are able to slide over one another easily.
Harder and stronger than group 1 metals	● Transition metals can lose **more than 2 valence electrons** to form positively charged ions. ● With more delocalised negatively charged electrons and a higher positively charged ion, the **electrostatic forces of attraction** between the positively charged ions and the sea of delocalised negatively charged electrons are **stronger**. Thus, transition metals have stronger metallic bonds than those in Group 1 and Group 2. ● Hence, more force must be applied to overcome the stronger metallic bonds to displace the transition metal ions from their fixed positions compared to Group 1 and 2 elements.
High densities and high melting points	***Explain why transition elements have high densities?*** ● Transition elements have a giant metallic lattice structure. ● In this structure, the transition element atoms are closely packed together in a regular arrangement due to strong electrostatic forces of attraction between the positively charged metal ions and the negatively charged delocalised electrons. ● At r.t.p, where transition elements are solids, there are a lot of atoms occupying the space per unit volume. ● Since transition elements have high relative atomic mass, density of these elements will be high.

Explain why transition metals have higher melting points than other metals?

- Transition metals can lose more than 2 valence electrons to form positively charged ions.
- With more delocalised negatively charged electrons and higher positively charged ions, the electrostatic forces of attraction between the positively charged ions and the sea of delocalised negatively charged electrons are stronger. Thus, transition metals have stronger metallic bonds than those in Group 1 and Group 2.
- Hence, more heat is required to overcome the stronger metallic bonds. Thus transition elements have higher melting points than other metals.

Table 14.3.2. The chemical properties of transition metals

They form coloured compounds

esCompounds of transition metals	Colour and physical state @r.t.p
Iron(II) hydroxide	Dark-green solid
Aqueous solution of iron(II) compounds	Green solution
Iron(III) hydroxide	Red-brown solid
Iron(III) oxide	Red-brown solid
Aqueous solution of iron(III) compounds	Yellow solution
Copper(II) carbonate	Green powdery solid
Copper(II) hydroxide	Light blue solid
Copper(II) oxide	Black powdery solid
Other copper(II) compounds	Blue solid/solution
Silver nitrate	White solid/Colourless solution
Silver chloride	White solid
Manganese(IV) oxide	Black powdery solid
Potassium manganate(VII)	Purple solid
Acidified potassium manganate(VII)	Purple solution
Manganese(II) compounds	White solid/Colourless solution

	The colour of a hydrated compound may be different from that of the anhydrous compound. Anhydrous copper(II) sulfate, $CuSO_4$ — white Copper(II) sulfate pentahydrate, $CuSO_4.5H_2O$ — blue Anhydrous cobalt(II) chloride, $CoCl_2$ — blue Cobalt(II) chloride hexahydrate, $CoCl_2.6H_2O$ — pink
Form ions with different oxidation states	**Application** Oxygen is able to bind to the haemoglobin molecules in the lungs and is released as blood passes through the body. This is possible due to the iron in the haemoglobin having different oxidation states of $+2$ and $+3$.
Transition elements and their compounds are good catalysts	**Applications** • Iron is used as a catalyst in Haber Process which manufactures ammonia • Catalytic converters contain catalysts namely platinum & rhodium which help to remove air pollutants in vehicle exhaust gases. • Nickel is used as a catalyst in the manufacture of margarine.

12.1 | Oxidation and Reduction

What Is Redox?

Definition 12.1.1 *(Redox)*

A redox reaction is a chemical reaction where both oxidation and reduction takes place simultaneously.

Oxidation	Reduction
Gain of oxygen	Loss of oxygen
Loss of hydrogen	Gain of hydrogen
Loss of electrons	Gain of electrons
Increase in oxidation state	Decrease in oxidation state

Gain and Loss of Oxygen

Definition 12.1.2 *(Oxidation)*

Oxidation is a reaction in which oxygen combines with an element or compound to form another substance.

Definition 12.1.3 *(Reduction)*

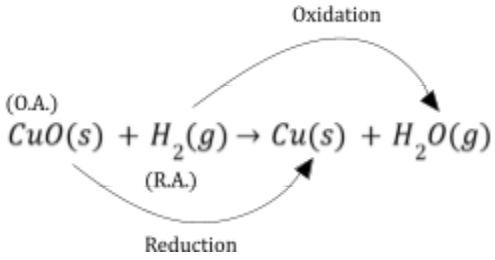
Reduction is a reaction in which oxygen is removed from a compound to form another substance.

Definition 12.1.4 *(Oxidising Agents)*

A substance that causes another substance to undergo oxidation. An oxidising agent loses oxygen to another substance, causing the other reactant to undergo oxidation while itself undergoes reduction.

Definition 12.1.5 *(Reducing Agents)*

A substance that causes another substance to undergo reduction. A reducing agent gains oxygen from another substance, causing the other reactant to be reduced while itself is oxidised.



Explain, in terms of oxygen, why the reaction between copper(II) oxide and hydrogen gas is a redox reaction.

- H_2 has gained oxygen from CuO to form H_2O. H_2 has undergone **oxidation**.
- CuO has lost oxygen to H_2 to form Cu. CuO has undergone **reduction**.
- Since oxidation and reduction occur at the same time, this is a redox reaction.

Identify the oxidising and reducing agents in this reaction. Explain your answer in terms of oxygen.

- CuO is an **oxidising agent**. CuO **caused** H_2 to be oxidised to H_2O by losing oxygen to H_2.
- H_2 is a **reducing agent**. H_2 **caused** CuO to be reduced to Cu by gaining oxygen to CuO.

Gain and Loss of Hydrogen

Definition 12.1.6 *(Oxidation)*

Oxidation is a reaction in which hydrogen is removed from a compound to form another substance.

Definition 12.1.7 *(Reduction)*

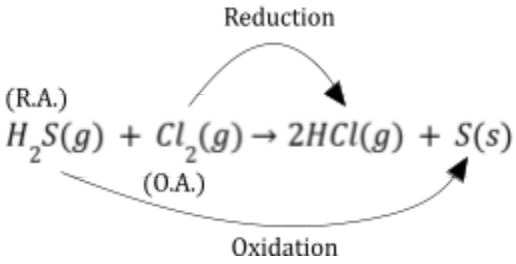
Reduction is a reaction in which hydrogen combines with an element or compound to form another substance.

Definition 12.1.8 *(Oxidising Agents)*

A substance that causes another substance to undergo oxidation. An oxidising agent gains hydrogen from another substance causing the other reactant to undergo oxidation while itself undergoes reduction.

Definition 12.1.9 *(Reducing Agents)*

A substance that causes another substance to undergo reduction. A reducing agent loses hydrogen to another substance causing the other reactant to be reduced while itself is oxidised.



Explain, in terms of hydrogen, why this is a redox reaction.

- H_2S lost hydrogen to Cl_2 to form S. H_2S has undergone **oxidation**.
- Cl_2 gained hydrogen from H_2S to form HCl. Cl_2 has undergone **reduction**.
- Since oxidation and reduction occur at the same time, this is a redox reaction.

Identify the oxidising and reducing agents in this reaction. Explain your reasoning in terms of hydrogen.

- Cl_2 is an **oxidising agent**. Cl_2 **caused** H_2S to be oxidised to S by gaining hydrogen from H_2S.
- H_2S is a **reducing agent**. H_2S **caused** Cl_2 to be reduced to HCl by losing hydrogen to Cl_2.

Gain and Loss of Electrons

Definition 12.1.10 *(Oxidation)*

Oxidation is a reaction in which a reactant loses electrons to form a substance.

Definition 12.1.11 *(Reduction)*

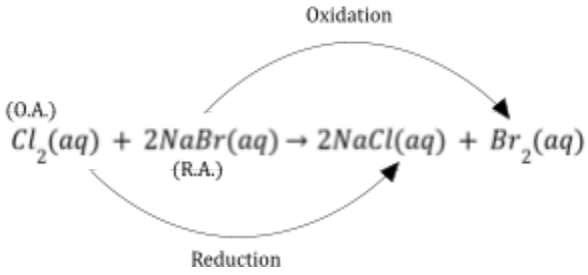
Reduction is a reaction in which a reactant gains electrons to form a substance.

Definition 12.1.12 *(Oxidising Agents)*

A substance that causes another substance to undergo oxidation. An oxidising agent gains electrons from another substance causing the other reactant to undergo oxidation while itself undergoes reduction.

Definition 12.1.13 *(Reducing Agents)*

A substance that causes another substance to undergo reduction. A reducing agent loses electrons to another substance causing the other reactant to be reduced while itself is oxidised.



Explain why this is a redox reaction in terms of electron transfer:

$2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$	Br^- ions in $NaBr$ lost electrons to Cl_2 to form Br_2. Thus Br^- is oxidised. Hence $NaBr$ is **oxidised**.
$Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$	Cl_2 gained electrons from Br^- ions in $NaBr$ to form Cl^- ions in $NaCl$. Hence Cl_2 is **reduced**.
Since oxidation and reduction occur at the same time, this is a redox reaction.	

Identify the oxidising and reducing agents in this reaction. Explain your reasoning in terms of electron transfer.

$2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$	Cl_2 is the **oxidising agent**. Cl_2 **caused** Br^- ions in $NaBr$ to be oxidised to Br_2 by gaining electrons from Br^- ions.
$Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$	$NaBr$ is the **reducing agent**. $NaBr$ **caused** Cl_2 to be reduced to Cl^- ions in $NaCl$ by losing electrons to Cl_2.

Oxidation State

Definition 12.1.14 *(Oxidation State)*

Oxidation state is the charge that an atom or an element would have if it existed as an ion in a compound.

Rule	Oxidation state
All elements	0
Group 1 and 2 metals	+1
	+2
Fluorine	-1
Hydrogen (usually)	+1
Oxygen (usually)	-2
Elements in compounds	Compounds have a no net charge, hence the oxidation states of these elements can be calculated from the rules above.

Definition 12.1.15 *(Oxidation)*

Oxidation is a reaction in which a reactant loses electrons to form a substance. When the reactant loses electrons, its oxidation state increases.

Definition 12.1.16 *(Reduction)*

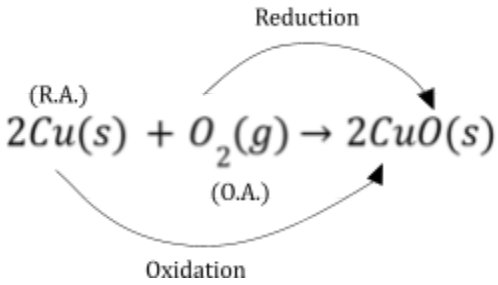
Reduction is a reaction in which a reactant gains electrons to form a substance. When the reactant gains electrons, its oxidation state decreases.

Definition 12.1.17 *(Oxidising Agent)*

A substance that causes another substance to undergo oxidation. An oxidising agent causes the oxidation state of another reactant to increase.

Definition 12.1.18 *(Reducing Agent)*

A substance that causes another substance to undergo reduction. A reducing agent causes the oxidation state of another reactant to decrease.



Explain, in terms of oxidation states, why the reaction between copper and oxygen is a redox reaction.

- The oxidation state of copper increases from 0 in copper metal to +2 in copper(II) oxide. Copper metal is oxidised.
- The oxidation state of oxygen decreases from 0 in oxygen gas to -2 in copper(II) oxide. Oxygen gas is reduced.
- Since oxidation and reduction occur at the same time, it is a redox reaction.

Identify the oxidising and reducing agents in this reaction. Explain your reasoning in terms of oxidation states.

- O_2 is an oxidising agent. O_2 caused Cu to be oxidised to CuO by increasing the oxidation state of Cu from 0 to +2 in CuO.
- Cu is a reducing agent. Cu caused O_2 to be reduced to CuO by decreasing the oxidation state of oxygen from 0 in O_2 to -2 in CuO.

Table 12.1.1. Redox vs Non-Redox Reactions

Examples of Redox Reactions	Example of Non-Redox Reactions
Reacting copper with oxygen to form copper(II) oxide	Neutralisation
Respiration	Precipitation
Photosynthesis	

Table 12.1.2. Oxidising and Reducing Agents

Examples of Oxidising Agents	Example of Reducing Agents
Halogens	Reactive metals
Oxygen gas in air	Hydrogen gas
Hydrogen peroxide	Carbon
Acidified potassium manganate(VII)	Carbon monoxide
	Hydrogen peroxide
	Potassium iodide
	Sulfur dioxide

Testing For Oxidising and Reducing Agents

Table 12.1.3. Test for Oxidising and Reducing Agents

	Test for oxidising agents	Test for reducing agents
Reagent used	potassium iodide, *KI*	acidified potassium manganate(VII)
Initial appearance of reagent	colourless solution	Purple solution
Procedure	Add a few drops of aqueous potassium iodide to a solution containing an oxidising agent.	**Solutions** Add a few drops of acidified potassium manganate(VII) solution to a solution containing a reducing agent.

		Gases Place a piece of filter paper soaked with acidified aqueous potassium manganate (VII) at the mouth of the test tube.
Observations	**Colourless** aqueous potassium iodide turns **brown**.	**Purple** acidified potassium manganate(VII) solution is **decolourised**.
Explanation	$2I^-(aq) \rightarrow I_2(aq) + 2e^-$ When KI(aq) is added to an oxidising agent, the oxidising agent caused colourless solution of I^- ions in $KI(aq)$ to be oxidised to brown $I_2(aq)$ by increasing the oxidation state of iodine from -1 in KI to 0 in I_2.	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$ When acidified $KMnO_4$ is added to a reducing agent, the reducing agent caused purple solution of MnO_4^- to be reduced to a colourless solution of Mn^{2+} ions by decreasing the oxidation state of manganese from +7 in $KMnO_4$ to +2 in Mn^{2+}.

> ⓘ **Note**
> Starch-iodide paper (starch+potassium iodide) can also be used to test for the presence of oxidising agents. Oxidising agents change the colour of moist starch-iodide paper from white to blue.

11.1 | Qualitative Analysis

Tests For Cations

Procedure for test

Aqueous sodium hydroxide, $NaOH(aq)$	**Aqueous ammonia,** $NH_3(aq)$
1) Add a few drops of aqueous sodium hydroxide to an aqueous solution of the unknown substance. 2) Add aqueous sodium hydroxide till in excess	1) Add a few drops of aqueous ammonia to an aqueous solution of the unknown substance. 2) Add aqueous ammonia till in excess

> ⓘ **Note**
>
> The alkalis provide the OH^- ions that react with cations to form insoluble metal hydroxide.
>
> $$M^{n+}(aq) + nOH^-(aq) \rightarrow M(OH)_n$$

Table 11.1.1. Observations for cation tests

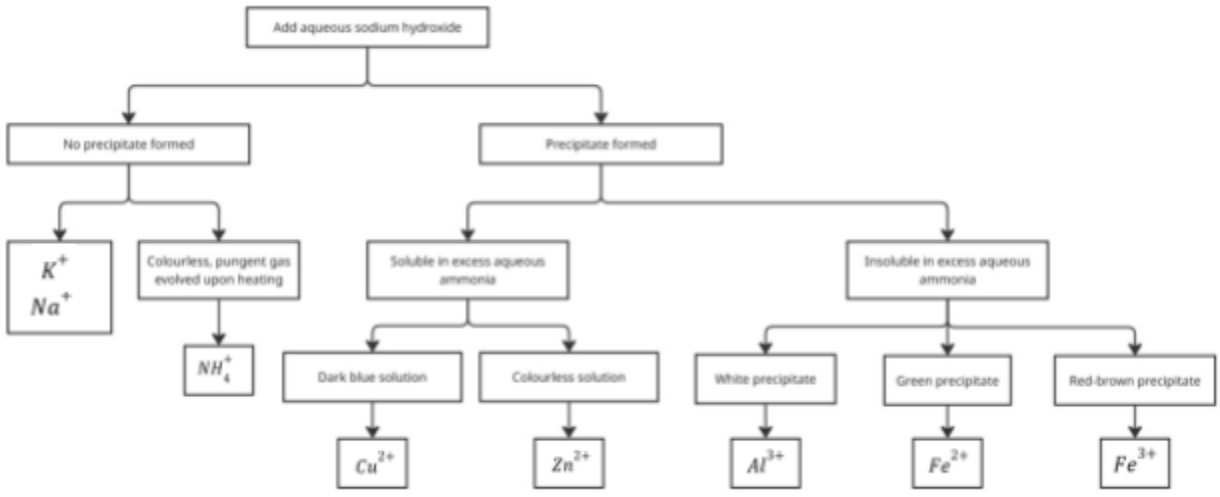
Cation	Aqueous sodium hydroxide, $NaOH(aq)$	Aqueous ammonia, $NH_3(aq)$
Aluminium, Al^{3+}	**White** precipitate formed upon addition of $NaOH(aq)$**.**	**White** precipitate formed upon addition of $NH_3(aq)$.
	White precipitate **dissolves** in excess aqueous sodium hydroxide to form a **colourless** solution.	**White** precipitate is **insoluble** in excess aqueous ammonia.
Ammonium, NH_4^+	No observable change upon addition of $NaOH$. The solution remains colourless. On warming, a colourless and pungent gas evolved. Gas turned moist red litmus paper blue. Gas is **ammonia**.	No observable change upon addition of $NH_3(aq)$.
Calcium, Ca^{2+}	**White** precipitate formed upon addition of NaOH(aq).	No observable change upon addition of $NH_3(aq)$.
	White precipitate is **insoluble** in excess aqueous sodium hydroxide	
Copper(II), Cu^{2+}	**Light-blue** precipitate formed upon addition of NaOH(aq).	**Light-blue** precipitate formed upon addition of $NH_3(aq)$.
	Light blue precipitate **insoluble** in excess aqueous sodium hydroxide.	**Light blue** precipitate **dissolves** in excess aqueous ammonia to give a **dark blue** solution.
Iron(II), Fe^{2+}	**Green** precipitate formed upon addition of $NaOH(aq)$	**Green** precipitate formed upon addition of $NH_3(aq)$.
	Green precipitate is **insoluble** in excess aqueous sodium hydroxide	Green precipitate **insoluble** in excess aqueous ammonia.
	On standing, **green** precipitate turned **red brown**. The green precipitate of $Fe(OH)_2$ is oxidised in air to form red-brown precipitate of $Fe(OH)_3$.	

Iron(III), Fe^{3+}	**Red-brown** precipitate formed upon addition of $NaOH(aq)$.	**Red-brown** precipitate formed upon addition of $NH_3(aq)$.
	Red-brown precipitate is **insoluble** in excess aqueous sodium hydroxide	Red-brown precipitate is **insoluble** in excess aqueous ammonia
Zinc, Zn^{2+}	**White** precipitate formed upon addition of NaOH(aq).	**White** precipitate formed upon addition of NaOH(aq).
	White precipitate **dissolves** in excess aqueous sodium hydroxide to form a colourless solution	White precipitate **dissolves** in excess aqueous ammonia to form a colourless solution

Colours of Common Metal Hydroxides

Metal hydroxide	Colour of precipitate
aluminium hydroxide	white
calcium hydroxide	white
copper(II) hydroxide	light-blue
iron(II) hydroxide	green
iron(III) hydroxide	red-brown
zinc hydroxide	white

Cation Test (Mind Map)



Tests For Anions

Table 11.1.2. Observations for anion tests

Anion	Reagents	Procedure	Observations
Carbonate, CO_3^{2-}	Any dilute acid	Add any dilute acid. Bubble the gas evolved into limewater.	Effervescence is observed upon addition of dilute acid. Colourless and odourless gas evolved. Gas formed white precipitate in limewater. Gas is **carbon dioxide**.
Chloride, Cl^-	Dilute nitric acid & aqueous silver nitrate	Add dilute nitric acid then add aqueous silver nitrate.	No observable change upon addition of nitric acid. White precipitate formed upon addition of aqueous silver nitrate.
	Aqueous ammonia	Add aqueous ammonia to the white precipitate of silver chloride till excess	White precipitate dissolves in aqueous ammonia to give a colourless solution.
Iodide, I^-	Add dilute nitric acid & aqueous silver nitrate	Add dilute nitric acid then add aqueous silver nitrate	No observable change upon addition of nitric acid. Yellow precipitate formed upon addition of aqueous silver nitrate.
Nitrate, NO_3^-	Aqueous sodium hydroxide & aluminium foil	Add aqueous sodium hydroxide then a piece of aluminium foil. Warm carefully.	Upon addition of NaOH(aq), aluminium foil and with warming, effervescence is observed. Colourless and pungent gas evolved. Gas turned moist red litmus paper turned blue. Gas is ammonia.
Sulfate, SO_4^{2-}	Dilute nitric acid and aqueous barium nitrate	Add dilute nitric acid then add aqueous barium nitrate	No observable change upon addition of nitric acid. White precipitate formed upon addition of aqueous barium nitrate.

Test For Gases

Colour and odour of common gases

Gas	Colour	Odour
Ammonia, NH_3	colourless	**pungent**
Carbon dioxide, CO_2	colourless	odourless
Chlorine, Cl_2	pale yellowish green	**pungent**
Hydrogen, H_2	colourless	odourless
Oxygen, O_2	colourless	odourless
Sulfur dioxide, SO_2	colourless	**pungent**
Hydrogen chloride, HCl	colourless	**pungent**
Nitrogen dioxide, NO_2	brown	**pungent**

Tests For Gases

Table 11.1.3. Observations for tests

Gas	Identification test	Observations
Ammonia, NH_3	Place a moist red litmus paper near the mouth of the test tube containing the gas	Upon warming the reaction mixture, colourless and pungent gas evolved. Gas turned moist red litmus paper blue. Gas is **ammonia**.
		Proof that ammonia is an alkaline gas Colourless and pungent NH_3 gas turned moist red litmus paper blue. Moist blue litmus paper remained blue.
Carbon dioxide, CO_2	Bubble the gas evolved through limewater	Colourless and odourless gas.evolved. Gas formed a white precipitate in limewater. Gas is **carbon dioxide.** When excess CO_2 is bubbled into limewater, the white precipitate dissolves to form colourless solution.
		$CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$ $CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$

<table>
<tr><td rowspan="2">Chlorine, Cl_2</td><td rowspan="2">Place a moist blue litmus paper near the mouth of the test tube containing the gas</td><td>Pale yellowish green and pungent gas evolved. Gas turned moist blue litmus paper red then white. Gas is <u>chlorine.</u></td></tr>
<tr><td>$Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HClO(aq)$

Production of chlorine
$NaClO(aq) + 2HCl(aq) \rightarrow Cl_2(g) + NaCl(aq) + H_2O(g)$</td></tr>
<tr><td>Hydrogen, H_2</td><td>Place a lighted splint near the mouth of the test tube containing the gas</td><td>Colourless and odourless gas evolved. Gas evolved extinguished lighted splint with a ‘pop’ sound. Gas is <u>hydrogen</u>.</td></tr>
<tr><td rowspan="2">Oxygen, O_2</td><td rowspan="2">Insert a glowing splint into the test tube containing the gas</td><td>Colourless and odourless gas evolved. Gas evolved relighted glowing splint. Gas is <u>oxygen.</u></td></tr>
<tr><td>Production of oxygen
$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$
$2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + NO_2(g) + O_2(g)$</td></tr>
<tr><td rowspan="3">Sulfur dioxide, SO_2</td><td rowspan="3">Place a strip of filter paper soaked in acidified potassium manganate (VII) solution near the mouth of the test tube containing the gas

OR

Bubble gas through a test tube containing acidified potassium manganate (VII) solution</td><td>ⓘ Note
Sulfur dioxide is an acidic oxide. (Refer: 8.4 - Types of Oxides)

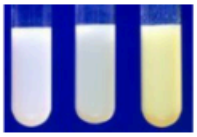
Colourless and pungent gas evolved. Gas evolved turned purple acidified potassium manganate (VII) colourless. Gas is <u>sulfur dioxide</u>.</td></tr>
<tr><td>MnO_4^- in purple acidified potassium manganate(VII) oxidises SO_2 to SO_4^{2-} by increasing the oxidation state of sulfur from +4 in SO_2 to +6 in SO_4^{2-}.

SO_2, the reducing agent caused purple solution of MnO_4^- to be reduced to a colourless solution of Mn^{2+} by decreasing the oxidation state of manganese from +7 in MnO_4^- to +2 in Mn^{2+}.</td></tr>
<tr><td>$5SO_2(g) + 2MnO_4^-(aq) + 16H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 8H_2O(l)$
purple solution (+7)+4(-2)= -1 +1 colourless solution +2 2(+1)+(-2)=0</td></tr>
</table>

Appearance of Common Compounds

- Group 1, 2, 13 and ammonium ions are white solids and form colourless solutions
- Zn^{2+} compounds are a white solid when cold and a yellow powdery solid when hot. They form colourless solutions when dissolved in water.

▪ **Appearance (colour and physical state) of Common Compounds**

Compounds containing these ions	In solid form	In aqueous form (If these compounds dissolve in water)
Ions of Group 1, 2, 13 metals (e.g. Na^+, Ca^{2+}, Al^{3+}) and Ammonium ion (NH_4^+)	White solid (crystals or powder) For e.g. Na_2O 	Colourless solution: For e.g. $NaNO_3$, K_2SO_4, $AlCl_3$ and $(NH_4)_2CO_3$ 
Cu^{2+} copper(II) ion	Black powdery solid: CuO Green powdery solid: $CuCO_3$ Blue solid: $CuSO_4.5H_2O$ 	Blue solution: $Cu(NO_3)_2$ and $CuSO_4$ Blue - green solution: $CuCl_2$ 
Fe^{2+} iron(II) ion	Green solid (green precipitate) For e.g. $Fe(OH)_2$ 	Pale green solution For e.g. $Fe(NO_3)_2$, $FeSO_4$ and $FeCl_2$ 
Fe^{3+} iron(III) ion	Red-brown solid For e.g. $Fe(OH)_3$ 	Yellow/Orange solution For e.g. $Fe(NO_3)_3$, $Fe_2(SO_4)_3$ and $FeCl_3$ 
Ag^+ silver ion	White solid: $AgCl$ Cream solid: AgBr Yellow solid: AgI 	Colourless solution: $AgNO_3$ 
Zn^{2+} zinc ion	ZnO: yellow powdery solid when hot and white solid when cold. 	Colourless solution: For e.g. $Zn(NO_3)_2$, $ZnSO_4$ and $ZnCl_2$ 
MnO_4^- manganate(VII) ions	Purple solid: $KMnO_4$ 	Purple solution: $KMnO_4$ 

NOTES FOR QUALITATIVE ANALYSIS

Test for anions

anion	test	test result
carbonate (CO_3^{2-})	add dilute acid	effervescence, carbon dioxide produced
chloride (Cl^-) [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	white ppt.
iodide (I^-) [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	yellow ppt.
nitrate (NO_3^-) [in solution]	add aqueous sodium hydroxide, then aluminium foil; warm carefully	ammonia produced
sulfate (SO_4^{2-}) [in solution]	acidify with dilute nitric acid, then add aqueous barium nitrate	white ppt.

Test for aqueous cations

cation	effect of aqueous sodium hydroxide	effect of aqueous ammonia
aluminium (Al^{3+})	white ppt., soluble in excess giving a colourless solution	white ppt., insoluble in excess
ammonium (NH_4^+)	ammonia produced on warming	–
calcium (Ca^{2+})	white ppt., insoluble in excess	no ppt.
copper(II) (Cu^{2+})	light blue ppt., insoluble in excess	light blue ppt., soluble in excess giving a dark blue solution
iron(II) (Fe^{2+})	green ppt., insoluble in excess	green ppt., insoluble in excess
iron(III) (Fe^{3+})	red-brown ppt., insoluble in excess	red-brown ppt., insoluble in excess
zinc (Zn^{2+})	white ppt., soluble in excess giving a colourless solution	white ppt., soluble in excess giving a colourless solution

Test for gases

gas	test and test result
ammonia (NH_3)	turns damp red litmus paper blue
carbon dioxide (CO_2)	gives white ppt. with limewater (ppt. dissolves with excess CO_2)
chlorine (Cl_2)	bleaches damp litmus paper
hydrogen (H_2)	'pops' with a lighted splint
oxygen (O_2)	relights a glowing splint
sulfur dioxide (SO_2)	turns aqueous acidified potassium manganate(VII) from purple to colourless