CHEMISTRY 101

Notes for Pure Chemistry so you won't die for O' levels
All information accurate as of 2020

All credits to Qin Guan for his notes (and me for making the docs)

TABLE OF CONTENTS

INSTRUCTIONS FOR USE

EXPERIMENTAL TECHNIQUES AND MEASUREMENTS
PURIFICATION AND SEPARATION
KINETIC PARTICLE THEORY
ELEMENTS, COMPOUNDS, AND MIXTURES
ATOMIC STRUCTURE
IONIC BONDING
COVALENT AND METALLIC BONDING
CHEMICAL EQUATIONS
STOICHIOMETRY AND THE MOLE CONCEPT
ACIDS AND BASES
SALTS
QUALITATIVE ANALYSIS
SPEED OF REACTION
ENERGY FROM CHEMICALS
AMMONIA
AIR AND THE ENVIRONMENT
REDOX
METALS
THE PERIODIC TABLE
ELECTROLYSIS

INSTRUCTIONS FOR USE

Hello people! Welcome to Chemistry 101!

Feel free to **print, download, or make a copy** of these notes! Here is the link to the original **Google Docs**:

https://docs.google.com/document/d/1ODpZq-XEX3ZVCgag6TEU-2IOEGtURe9odrX6c9w4zzU/edit?usp=sharing

This set of notes was made in 2020, do note the following topics that have not been included in this set of notes:

FUELS AND CRUDE OIL

ALKANES AND ALKENES

ALCOHOLS, CARBOXYLIC ACIDS AND ESTERS

MACROMOLECULES

Hopefully people will benefit from this and do well for chemistry. This set of notes will help you to break everything down into smaller concepts so that it's easier for digestion and understanding. :)

Hopefully less people will die for O' levels Chemistry (Pure) after using this set of notes.

I would like to fully credit myself for making the google docs and doing notes for all the S4 chapters. I would also like to credit my friend Qin Guan, for writing all notes for all the S3 chapters and allowing me to convert it into a google docs.

Jun Ci

(The human who did the google docs)

EXPERIMENTAL TECHNIQUES AND MEASUREMENTS

PHYSICAL QUANTITIES

Physical Quantity	Unit	Conversion	Measurement
Mass	Kg	1kg = 1000g	Beam balance, Electronic balance
Volume	cm^3	$1cm^3$ = 0.001 liters	Measuring cylinder, Burette
	dm^3	$1dm^3 = 1000cm^3$ = 1 liter	
Time	Seconds (s)	1h = 60min= 3600s	Stopwatch
	Minutes (min)		
	Hours (h)		
Temperature	Celsius (ºC)	-	Thermometer
	Kelvin (K)	Celsius + 273K	

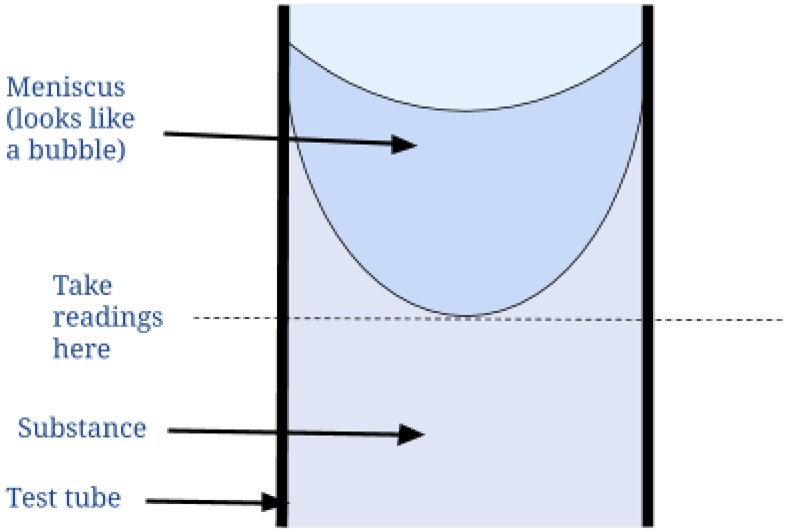
MEASUREMENT APPARATUS

Apparatus used	How it is used
Electronic Balance	● Values are given at 2 d.p. ● Mini glass enclosure around the balance to prevent wind from affecting measurements
Measuring Cylinder	● Approximate measurements ● $1cm^3$ to $1000cm^3$ or more
Burette	● Accurate measurements to $0.05cm^3$ ● Measure up to $50cm^3$
Volumetric flask	● Very accurate

	- Used to measure fixed volumes - Used to prepare standard solutions of exact concentration
Syringe	- Small syringe → small volume of liquids - Large syringe → large volume of gases

TAKING MEASUREMENTS

Take the reading of a liquid at the bottom of the meniscus.



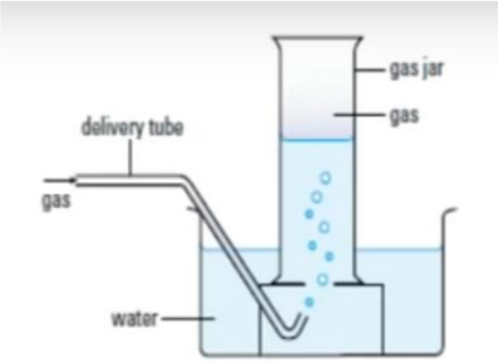
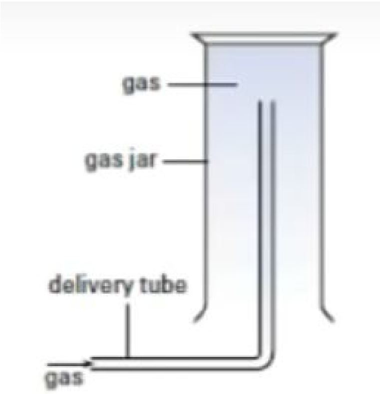
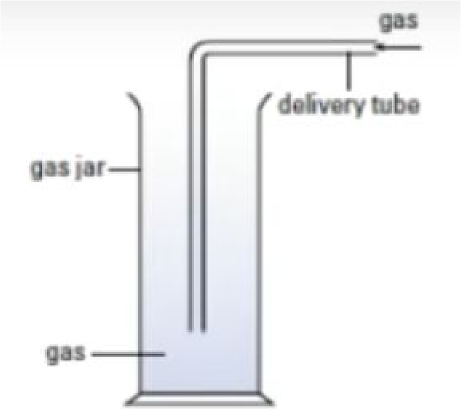
WHY DO WE NEED TO COLLECT GASES?

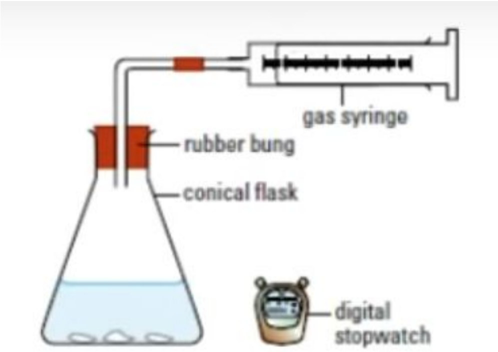
Gases allow us to determine the rate and progress of a reaction.

METHODS OF COLLECTING GASES

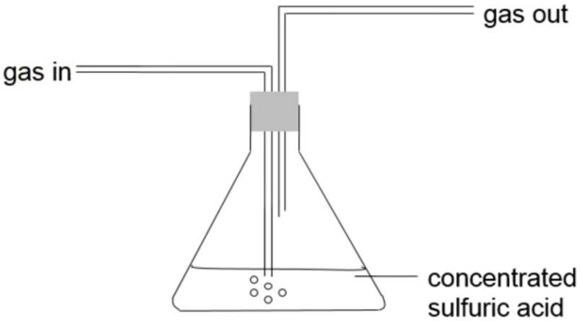
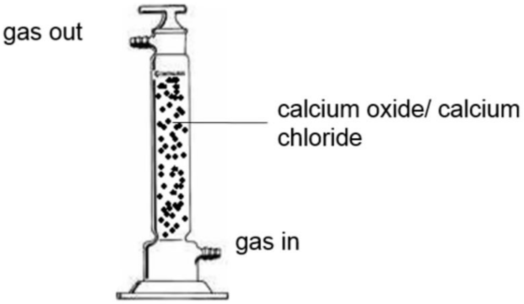
Considerations to be made:

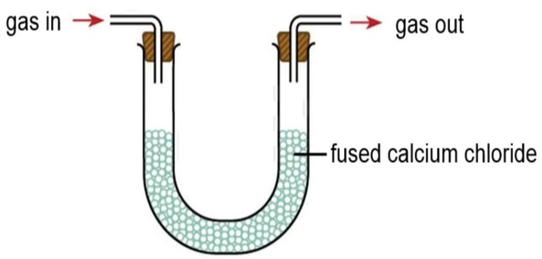
- Density (compared to the air)
- Solubility (in water)

Methods	Diagram	When to be used
Displacement of water	 	For gases which are insoluble in water ● E.g. H_2, O_2, CO_2
Upward delivery	 	For gases which are denser in the air (and soluble in water) ● E.g. Cl_2
Downward delivery	 	For gases which are less dense than air (and soluble in water) ● E.g. NH_3, H_2

Gas syringe		When there is a need to monitor the rate of evolution in a gas

HOW TO COLLECT DRY SAMPLES OF GAS

Properties of gas	Diagram
Acidic (e.g. Concentrated sulfuric acid (H_2SO_4))	
Alkaline (e.g. Calcium Oxide (CaO))	

Neutral (e.g. Fused Calcium Chloride)	

PROPERTIES OF COMMON GASES

Gas	Formula	Colour	Density (compared to air)	Solubility in water
Ammonia	NH_3	Colourless	Less dense	Extremely soluble
Carbon Dioxide	CO_2	Colourless	Denser	Slightly soluble
Chlorine	Cl_2	Greenish-Yellow	Denser	Soluble
Hydrogen	H_2	Colourless	Less dense	Insoluble
Hydrogen Chloride	HCl	Colourless	Denser	Very soluble
Oxygen	O_2	Colourless	Slightly denser	Slightly soluble
Carbon Monoxide	CO	Colourless	Slightly denser	Slightly soluble
Nitrogen	N_2	Colourless	Slightly denser	Insoluble
Sulfur dioxide	SO_2	Brownish	Denser	Very soluble
Helium	He	Colourless	Less dense	Insoluble

PURIFICATION AND SEPARATION

WHAT IS A PURE SUBSTANCE?

- A single substance that is not mixed with anything else
- Fixed melting and boiling point
- Only produces a single spot on a chromatogram

WHAT IS A MIXTURE?

- Consists of 2 or more substances that are mixed (not chemically combined together, can be separated)
- An impure substance can be a mixture of more than 2 substances (e.g. water + sand + silt)
- Melts and boils over a range of temperatures
    - The boiling point becomes higher
    - The melting point becomes lower

PURITY IS IMPORTANT BECAUSE IT AFFECTS...

- The extent of a reaction
    - How far the reaction proceeds
- The efficiency of a reaction
- The efficacy of a chemical
    - **Efficacy**: the ability to produce the intended result
- Price of a material

Separation techniques

DECANTING

Separates an insoluble solid from a liquid by pouring off the liquid from the container carefully.

FILTRATION

Used to separate an insoluble solid from a liquid.

Apparatus:

- Beaker
- Filter funnel
- Filter paper

Key terms:

- **Residue**: The solid left in the filter paper
- **Filtrate**: The liquid collected in the beaker

EVAPORATION TO DRYNESS

Separates dissolved solids from a solution.

Note: Some solids are not heat stable and will decompose when heated, even after evaporation, and some impurities will still be present in the solution

Apparatus: Evaporating dish

CRYSTALLISATION

Separates a dissolved solid from a solution, forming pure crystals of the substance. The filtrate has to be saturated before the flame is turned off. Rinse the crystals with cold distilled water (to remove impurities).

The solubility of the solute decreases, leaving the extra solute that cannot remain dissolved to be separated as crystals.

Apparatus:

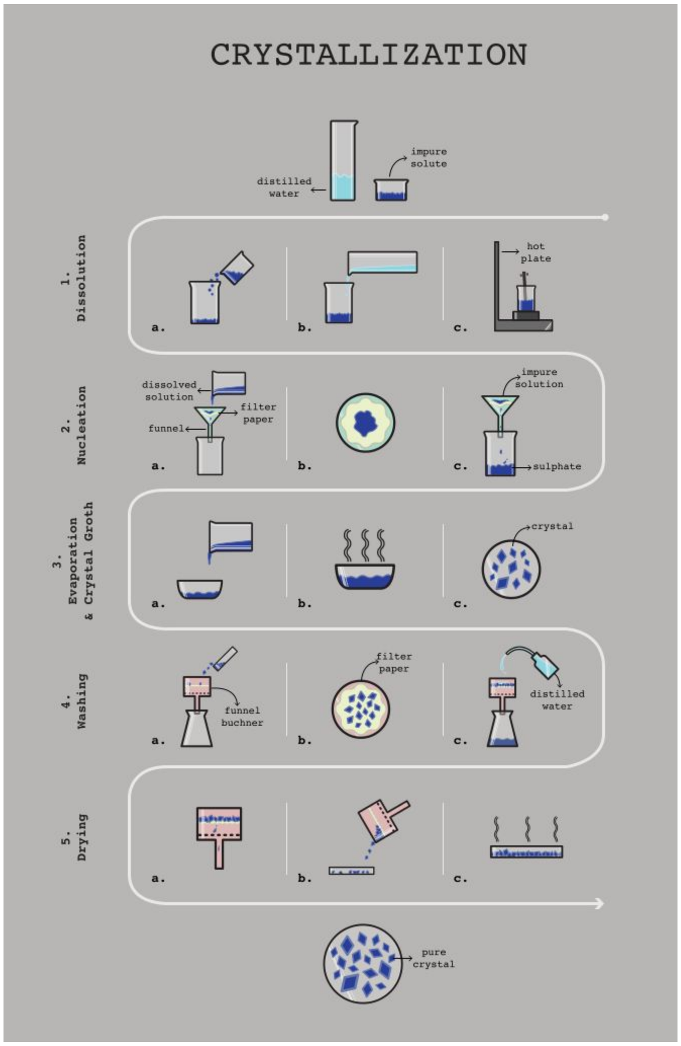
- Evaporating dish
- Filter paper
- Bunsen burner
- Beaker
- Filter funnel
- Stirring rod

Key terms:

- **Filtrate**: Liquid collected in the beaker after filtration
- **Solvent**: Liquid in mixture along with solvent
- **Crystals**: Pure solids obtained
- **Impure solid**: Solid in mixture along with solvent
- **Solute**: The substance dissolved in a given solution

Example of substances that can crystalise:

- The crystallisation of Copper (II) sulfate solution



SUBLIMATION

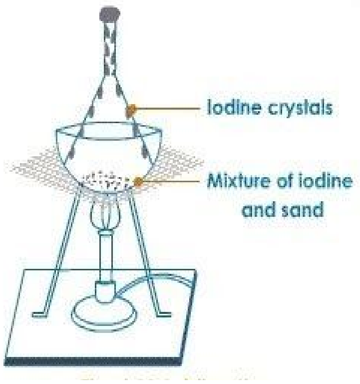
Used to separate a mixture of solids, one of which sublimes.

Key terms:

- **Sublime**: The change of state from the solid state to the gaseous state

Example of substances that can sublime:

- Iodine found in a mixture of iodine and sand will sublime when heated. Ammonium compounds can also sublime.



SIMPLE DISTILLATION

Separates a pure liquid from a solution.

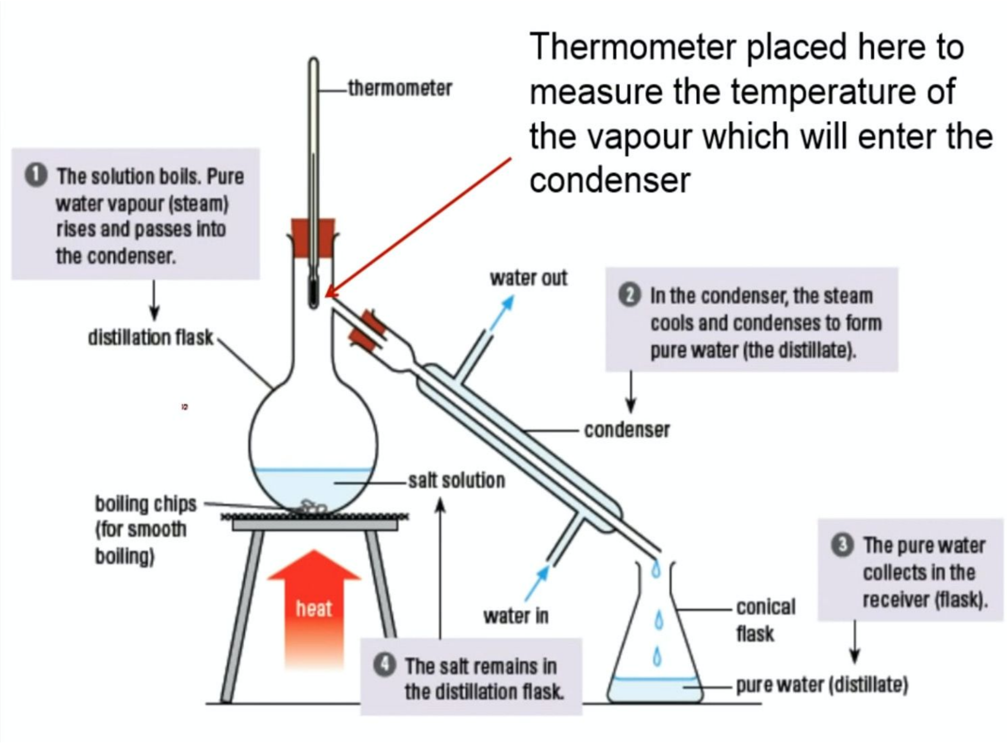
Note: A thermometer is placed near the exit as we want to measure the temperature of the vapour. Water enters from the bottom of the condenser and comes out from the top as the coolest region should be at the end so that more vapour can be condensed.

Apparatus:

- Distillation flask
- Boiling chips
- Thermometer
- Liebig condenser
- Conical flask

Key terms:

- Salt solution
- Pure water
- Distillate: Pure liquid collected in the receiver (flask)



Thermometer placed here to measure the temperature of the vapour which will enter the condenser

① The solution boils. Pure water vapour (steam) rises and passes into the condenser.

② In the condenser, the steam cools and condenses to form pure water (the distillate).

③ The pure water collects in the receiver (flask).

④ The salt remains in the distillation flask.

FRACTIONAL DISTILLATION

Separates mixtures of miscible liquids with different boiling points. The fractionating column has convolutions or glass beads packed inside to provide a larger surface area to volume ratio for cooling of the vapour, which allows for better separation.

Note: ***Evaporation happens at any temperature.*** *Hence, both substances will evaporate when heated.*

- *When mixture rises to the fractionating column, the one with the lower boiling point will go up the column, the one with a higher boiling point will condense in the fractionating column.*
- *A thermometer is placed near the exit to measure the temperature of the vapour.*

Apparatus:

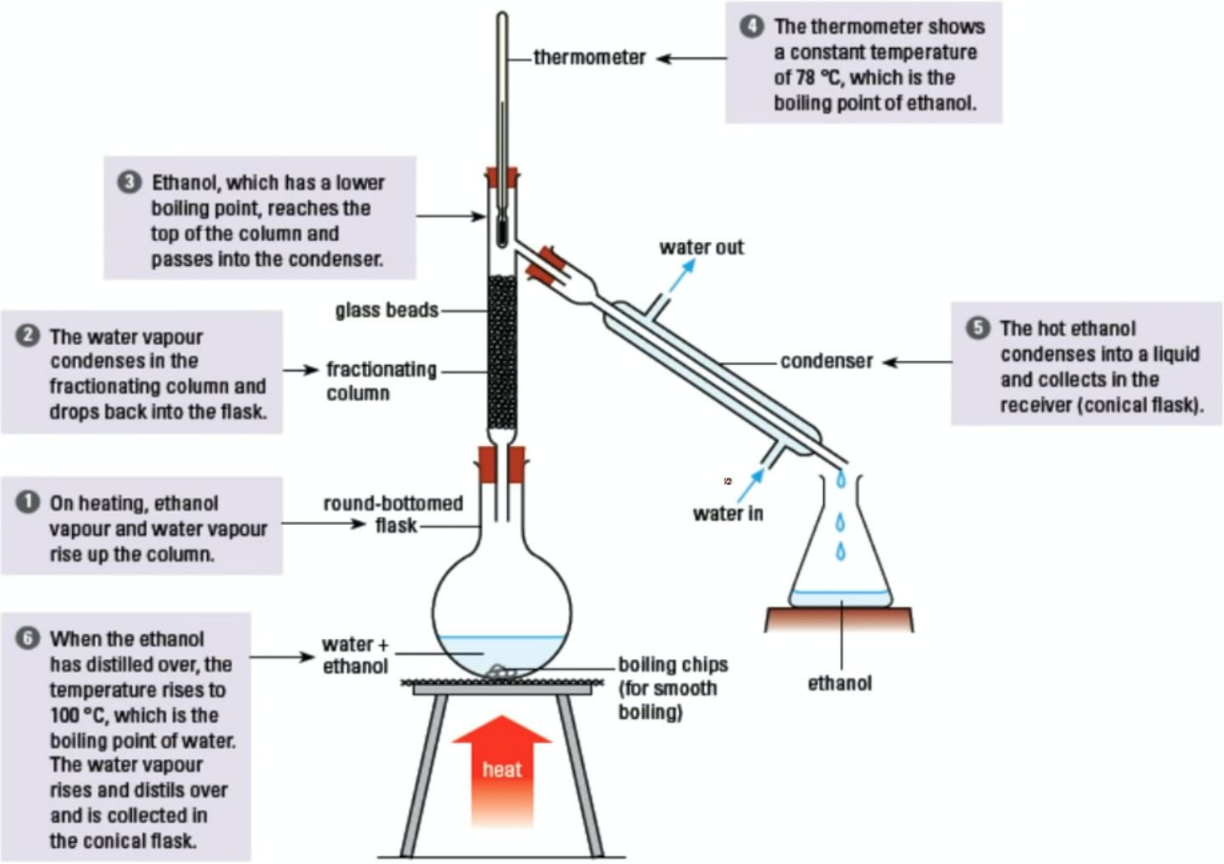
- Distillation flask
- Boiling chips
- Thermometer
- Liebig condenser
- Conical flask
- Fractionating column (with glass beads)

Key terms:

- **Miscible**: The capability of being mixed to form a solution in any ratio

Example of substances used in fractional distillation:

- Water and ethanol, during production of vodka.



① On heating, ethanol vapour and water vapour rise up the column.

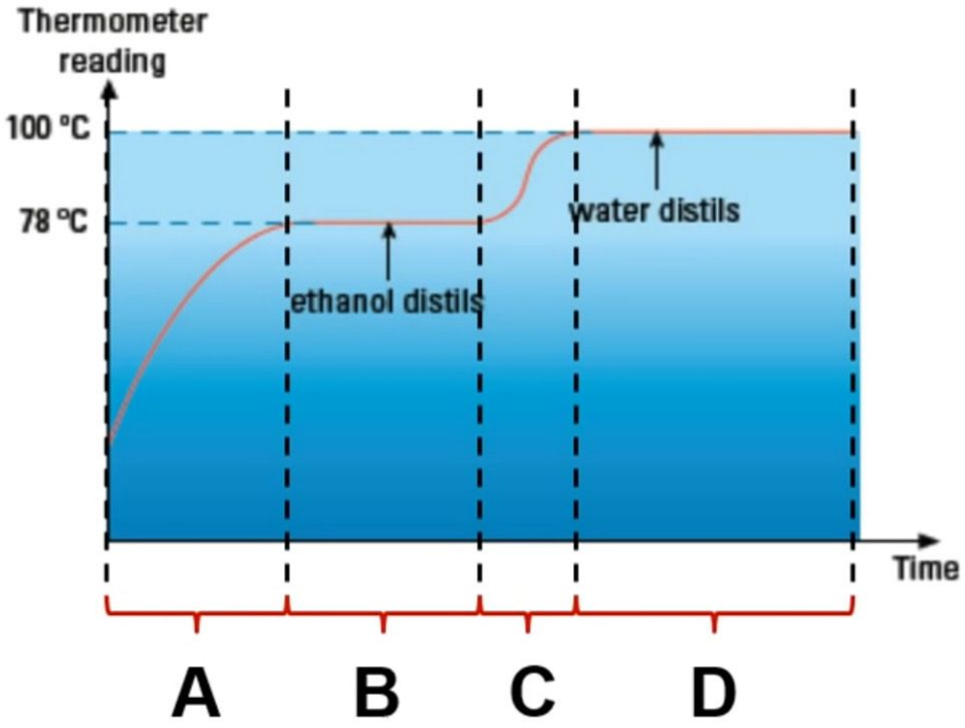
② The water vapour condenses in the fractionating column and drops back into the flask.

③ Ethanol, which has a lower boiling point, reaches the top of the column and passes into the condenser.

④ The thermometer shows a constant temperature of 78 °C, which is the boiling point of ethanol.

⑤ The hot ethanol condenses into a liquid and collects in the receiver (conical flask).

⑥ When the ethanol has distilled over, the temperature rises to 100 °C, which is the boiling point of water. The water vapour rises and distils over and is collected in the conical flask.



Segment A

- The temperature rises as the ethanol-water mixture is heated

Segment B

- The temperature reaches the boiling point of ethanol
- The temperature remains at the boiling point of ethanol as ethanol boils off
- Any water which evaporates off is cooled and condensed by the condenser

Segment C

- All ethanol distilled off and temperature rises

Segment D

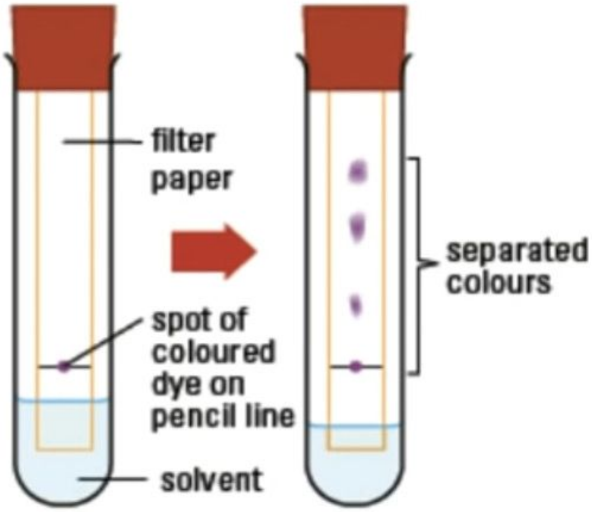
- The temperature of the mixture reaches the boiling point of water
- The temperature remains at the boiling point of water as the water boils off
- The temperature remains constant at B and D as the thermal energy gained is used to boil off the liquid

CHROMATOGRAPHY

Used to separate and identify mixtures.

Steps:

1. Dissolve dye in a suitable solvent
2. A drop of the dissolved dye is placed on the pencil line near the bottom of the paper
3. A line is drawn to determine the starting position of the drop dye to determine the Rf value
4. Solvent travels up the filter paper and dye in the spot dissolved and travel up the paper at different speeds
5. This separates the dye into its various components (i.e. Less soluble = moves up slower, more soluble = moves up faster)
6. Solvent from reaches nearly the top of the filter paper
7. The filter paper is removed



Real-life applications of chromatography:

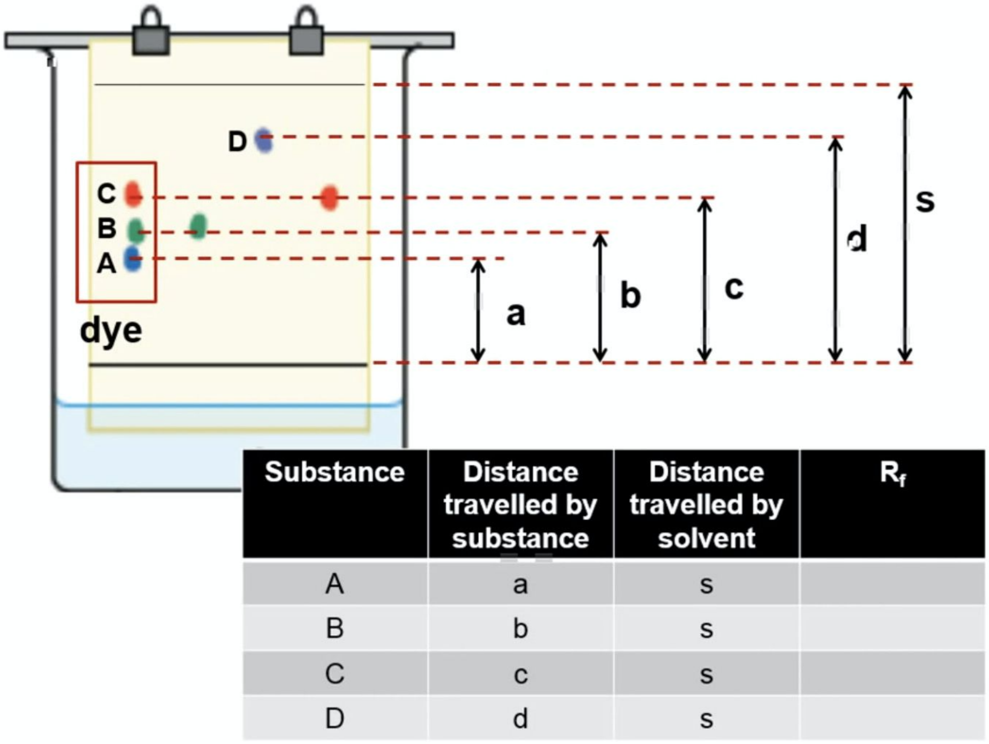
- Separate and identify components present in coloured substances and food dyes
- Separate and identify substances in drugs
- Separate and identify substances in blood and urine to check if athletes used drugs

Advantages of chromatography:

- Only a small sample is used
- Sample can be recovered

Interpretation of data in chromatography:

- Identical substances travel the same distance up paper
- Distance travelled depends on the solubility of the substance in the solvent. The more soluble the substance, the further it travels up
- Compare the components in the dye with standards used
- **Rf value** = $\dfrac{\textit{Distance travelled by a substance}}{\textit{Distance travelled by the solvent}}$



Substance	Distance travelled by substance	Distance travelled by solvent	R_f
A	a	s	
B	b	s	
C	c	s	
D	d	s	

KINETIC PARTICLE THEORY

MATTER

Matter is anything that has mass and takes up space. All matter has different states.

KINETIC PARTICLE THEORY STATES THAT…

- Particles are discrete and too small to be seen directly
- There are spaces between particles of matter and this space is different between solids, liquids and gasses
- Particles are in constant random motion, and they move at different speeds in solids, liquids, and gases

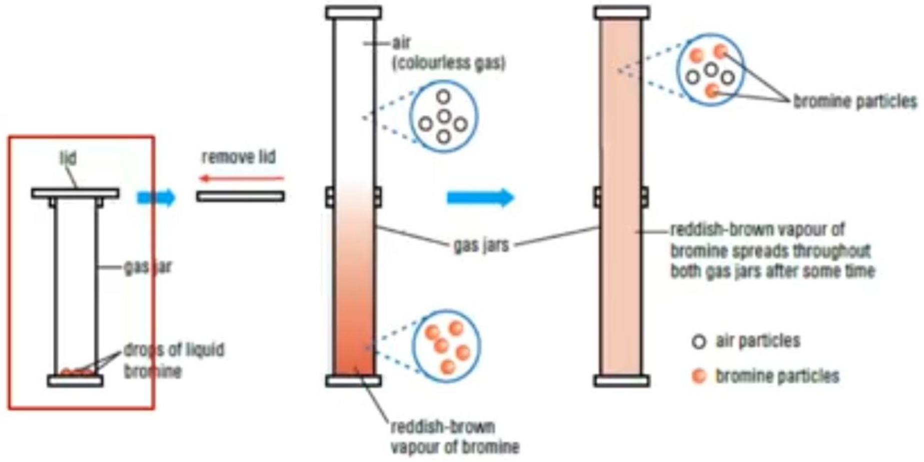
DIFFUSION

Diffusion: the spreading and mixing of particles in liquids and gasses from a region of higher concentration to a region of lower concentration down a diffusion gradient.

Examples of diffusion:

In gases

1. A few drops of liquid bromine are placed in a gas jar
2. Bromine will evaporate and diffuse upwards until both jars are evenly mixed



In water

1. Crystal of potassium manganate (VII) dissolves to form a purple solution
2. Potassium manganate diffuses and spreads out through the water

WHAT AFFECTS DIFFUSION

Temperature

Particles move faster at higher temperatures as there's more kinetic energy, as particles spread around spaces quicker. Hence, faster diffusion

Mass of particle

The greater the molecular mass, the slower the particle moves, hence the slower rate of diffusion.

PARTICLES

Particles in a **solid**	Particles in a **liquid**	Particles in a **gas**
• Packed closely together in an orderly arrangement • Very little empty space between particles • Cannot be compressed • Vibrate but cannot move freely about their fixed positions	• Packed closely together, but not in an orderly arrangement • Very little empty space between particles, although more than solids • Cannot be compressed • Vibrate and can move freely throughout the liquid	• Far apart and in a random arrangement • Large spaces between particles • Can be compressed easily • Vibrate and are free to move throughout the container

MACROSCOPIC VS MICROSCOPIC PROPERTIES

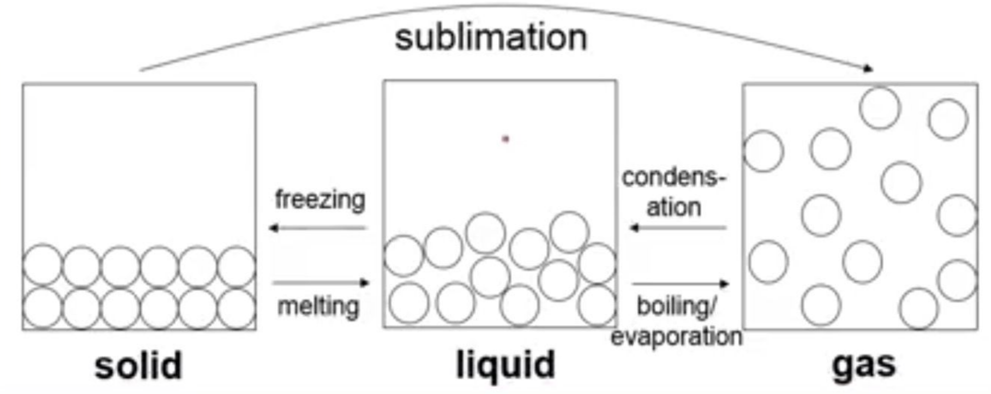
Macroscopic: Matter

- May be coloured
- May be hot or cold
- Expands when heated

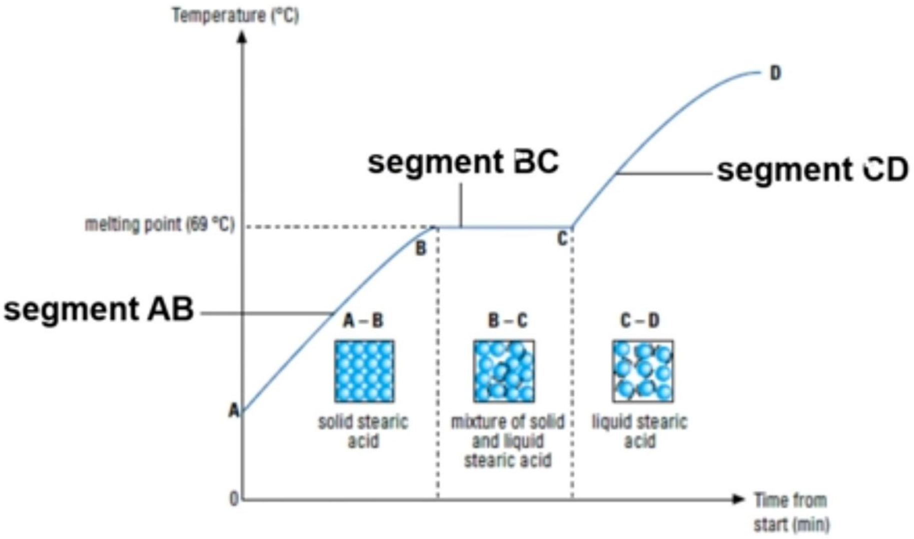
Microscopic: Particles

- Are not coloured
- Do not get hot or cold
- Do not expand

Change in states of matter



MELTING



At segment **AB**:

- The temperature of solid increases until it reaches the melting point at 69°C
- Solid begins to melt

At segment **BC**:

- Heat energy is used to overcome forces of attraction holding the particles together instead of being used to raise the temperature

- Temperature does not rise and remains constant during melting
- Solid-liquid mixture

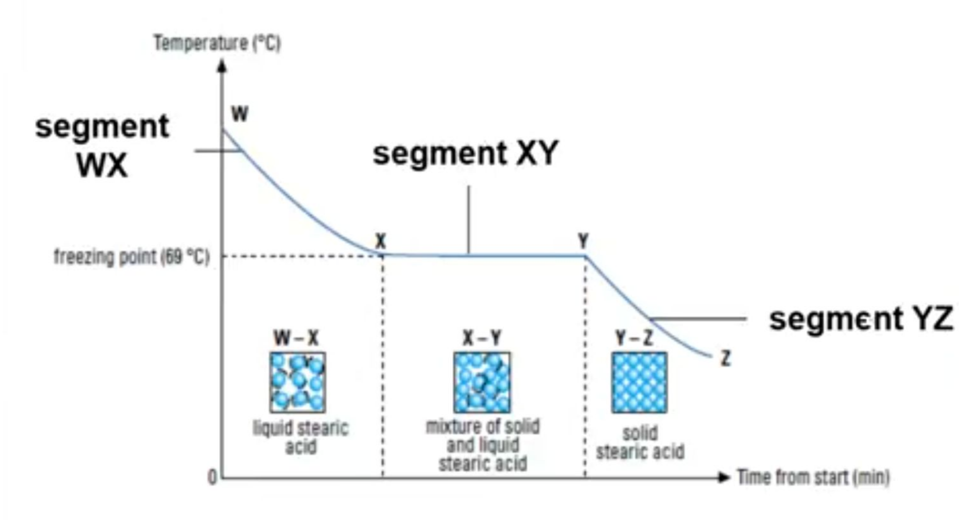
At segment **CD**:

- All solid has melted
- The temperature of the liquid increases now as more heat is absorbed

Explanation:

Particles absorb energy, hence they vibrate faster, which allows them to overcome forces of attraction, hence letting particles move about freely.

FREEZING



At segment **WX**:

- The temperature of the liquid decreases until it reaches freezing point
- The liquid begins to freeze

At segment **XY**:

- Heat energy is released from the liquid when the movement of the particles slow down and take up fixed positions in the solid
- Temperature does not decrease further and remains constant

- Solid-liquid mixture

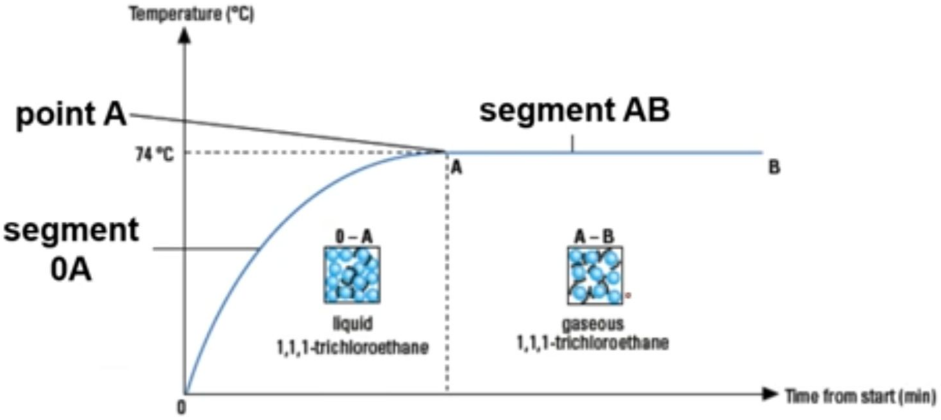
At segment **YZ**:

- Liquid changed into a solid
- Temperature decreases as the cooling continues

Explanation:

The particles lose energy, hence they vibrate slower. This causes the particles to be overcome by the forces of attraction, thus the particles cannot move about freely.

BOILING



Segment **0A**:

- The temperature of liquid increases until it reaches boiling point
- The liquid begins to boil

Segment **AB**:

- Heat energy is absorbed by the particles of the liquid and used to overcome the forces holding them together
- Temperature does not change and remains constant
- Particles form a gas and move about freely

Explanation:

The particles absorb energy, making them vibrate faster. This causes the particles to be overcome by the forces of attraction, hence they are able to move about freely and further apart.

CONDENSATION

The temperature at which a gas condenses to form a liquid is the same as the boiling point.

EVAPORATION

The change in state from liquid to gas without boiling. Evaporation occurs below the boiling point of a liquid, and the particles escape from the surface of the liquid to form a gas.

SUBLIMATION

The change in state from solid to gas, without melting to form a liquid.

ELEMENTS, COMPOUNDS, AND MIXTURES

ELEMENTS VS MOLECULES VS COMPOUNDS VS MIXTURES

Elements	**Molecules**	**Compounds**	**Mixtures**
A substance that cannot be broken down into simpler substances by chemical methods and electrolysis.	2 or more atoms chemically combined together. This term is only used for **covalent substances**.	Substances containing 2 or more elements chemically combined together.	Consists of 2 or more substances that are not chemically combined together **Compound ≠ mixture**

General rules when **naming** compounds:

General rule	**Examples**
If a compound contains both a metal and non metal, the metal is usually written first in the formula.	Sodium chloride (NaCl) is written with sodium first
Compounds with only 2 elements often end in -ide	Sodium chloride Magnesium chloride
Compounds that contains OH^- ion is named a hydroxide	Sodium hydroxide Calcium hydroxide Copper (II) hydroxide
Compound with an ion containing oxygen usually ends with -ate	Calcium carbonate Silver nitrate Lead (II) sulfate
The number of atoms in a formula is written as subscript	H_2O $CuSO_4$
Roman numerals are used to denote isotopes	Iron (II) hydroxide, $Fe(OH)_2$ Iron (III) hydroxide, $Fe(OH)_3$

For covalent molecules, for some compounds, we usually name the entire compound by counting the number of atoms	Carbon Dioxide (CO_2) Carbon monoxide (CO)
For others, we use their common names	Water (H_2O) Ammonia (NH_3)

COMPOUNDS VS MIXTURES

	Compound	**Mixture**
Composition	Fixed composition by mass	Variable composition by mass
Melting and boiling points	Fixed melting and boiling points	Variable melting and boiling points
Properties	Physical and chemical properties of a compound are different from its elements	Does not have its own properties, but rather have properties of its components
Separation	Cannot be separated into 2 or more substances by physical means; a chemical reaction is required to separate the elements	Can be separated into its components by physical means without a chemical reaction

BALANCING CHEMICAL EQUATIONS

Aqueous sodium hydroxide reacts with dilute hydrochloric acid to give aqueous sodium chloride and water.

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

LHS	**RHS**	**Outcome**
1 Na	1 Na	Balanced

1 O	1 O	Balanced
2 H	2 H	Balanced
1 Cl	1 Cl	Balanced

ATOMIC STRUCTURE

Atoms

ATOM

Smallest component of an element having the chemical properties of the element.

- **Proton number** = **Electron number** in an *electrically neutral atom*
- Almost all the mass of an atom is contained in its nucleus as the mass of electrons is negligible compared to the mass of the protons and neutrons

ISOTOPES

Isotopes: atoms of the same element (or same number of protons) with different numbers of neutrons.

- Different physical properties but similar chemical properties
- Chemical properties depends on the number of electrons

	Number of protons	**Mass number**	**Number of neutrons**
Chlorine-35	17	35	18
Chlorine-37	17	37	20

ELECTRONIC STRUCTURE OF AN ATOM

Electronic structure: The arrangement of electrons in an atom in the different shells.

Valence shell electrons: The shell which is the furthest away from the nucleus is called the outer shell or valence shell.

- *Chemical properties* of the atom are determined by the valence shell electrons

ELECTRONIC CONFIGURATION

The electronic configuration is the numerical expression of the number of electrons in each shell in an atom.

- Chlorine-35: 2, 8, 7
- Carbon-12: 2, 4

PERIODIC TABLE

All atoms in the same **group** (1 column) have the same number of valence electrons.

All groups in the same **period** (1 row) have the same number of electron shells

FORMATION OF IONS

Atoms lose or gain valence electrons to attain a stable electronic structure of a noble gas. This structure is called an **ion**.

- A lithium atom has 3 electrons. It loses its 1 valence electron to achieve stable duplet configuration. This forms a lithium ion which has 2 electrons.

The noble gases are group 0 in the periodic table. They have a **full valence shell**, which makes them unreactive and stable.

Metals tend to form **cations** by losing their valence electron(s) to attain noble gas configuration. **Non metals** tend to form **anions** by gaining valence electron(s) to attain noble gas configuration.

However, noble gases are monatomic as they are unreactive and do not readily form bonds with other atoms.

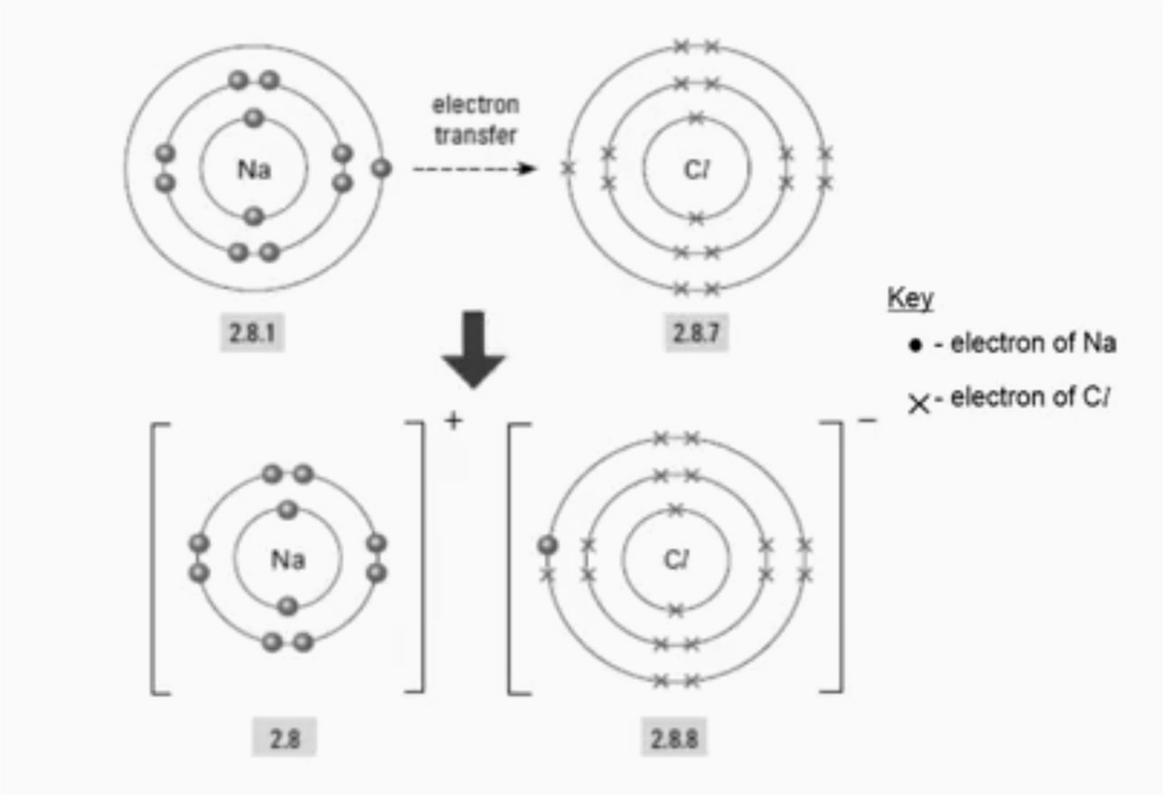
IONIC BONDING

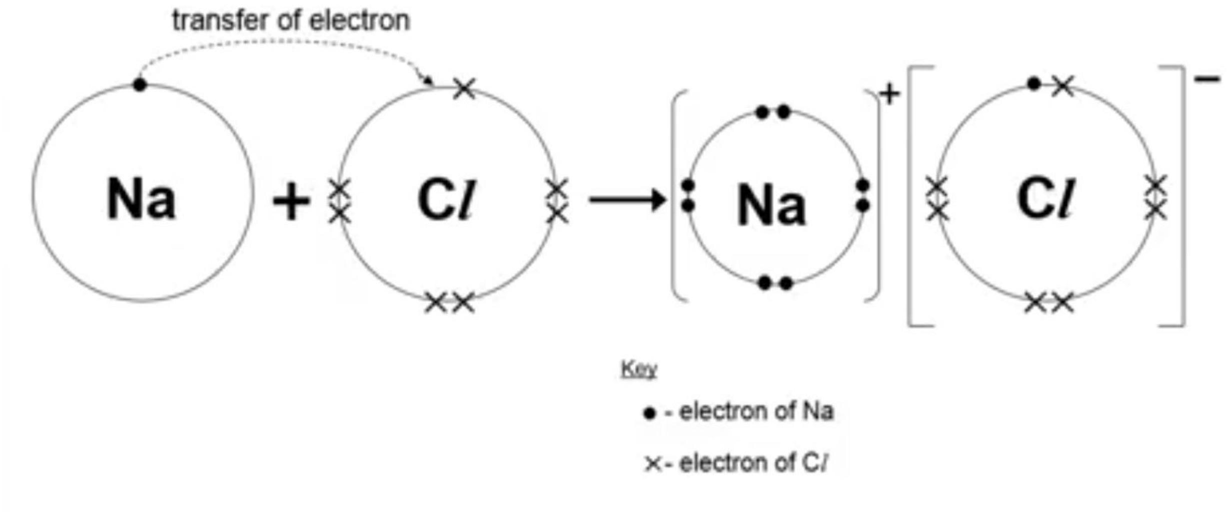
IONIC BOND

Ionic bonds are the **strong electrostatic forces of attraction between oppositely charged ions**.

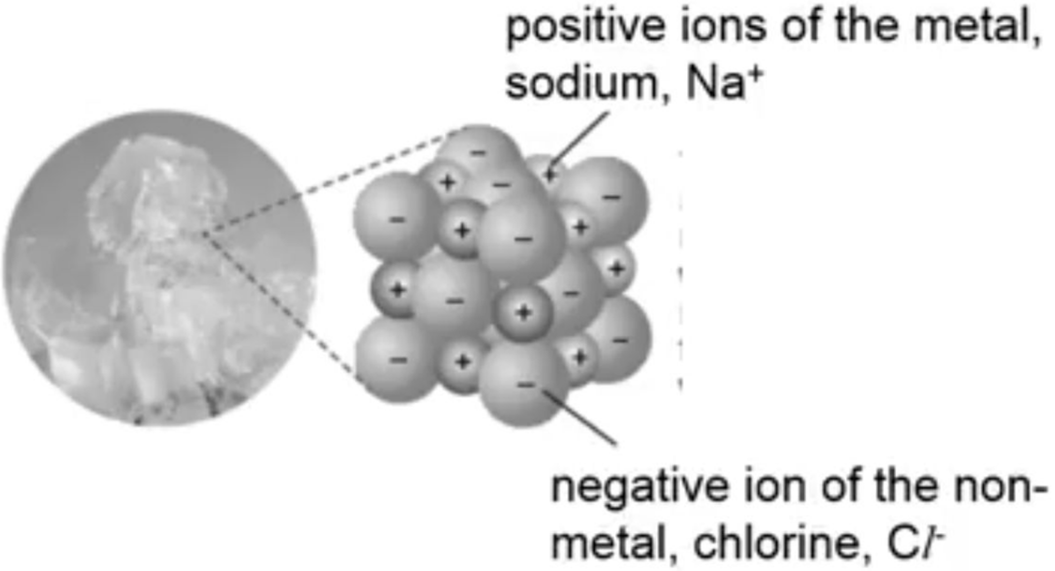
- Results when electrons are transferred from atoms of one element to atoms of another element
- Leads to ionic compounds having a high melting and boiling point.
- Orderly arrangement of ions
- Usually formed between metallic elements (cations) and non-metallic elements (anions)

DOT AND CROSS DIAGRAM





STRUCTURE OF IONIC COMPOUNDS



- Positive ions (usually metals) and negative ions (usually non-metals) are arranged in an **orderly manner**
  - If it is not in an orderly manner, the layer will split apart as the attractive

forces are directional

- All ionic compounds have a **giant ionic lattice structure**

PHYSICAL PROPERTIES OF IONIC COMPOUNDS

Ionic compounds are usually crystalline solids that are **hard but brittle**

- **Hard**: large amounts of energy is required to overcome the strong electrostatic forces of attraction between the cations and anions (oppositely charged ions)
- **Brittle**: When the layers of ions slide, it brings ions of like charges next to each other, causing these ions of like charges to repel each other. Thus, the ionic lattice shatters.

Ionic compounds have **high melting and boiling points.**

- Large amounts of energy must be supplied to overcome the strong forces of attraction between the cations and anions (oppositely charged ions)

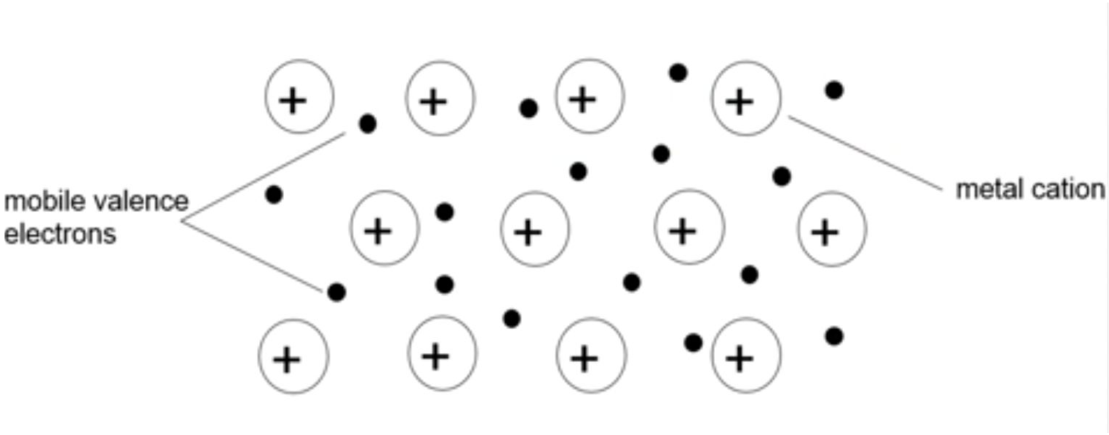
Conducts electricity in molten and aqueous states but not in solid state.

- **Solid**: The ions are held in fixed positions in a giant ionic lattice crystal structure by strong electrostatic forces of attraction, and thus the ions are in fixed positions, leading to the absence of mobile charge carriers.
- **Molten/Aqueous**: The ions are mobile (delocalised) and will migrate to the electrodes when a potential difference is applied

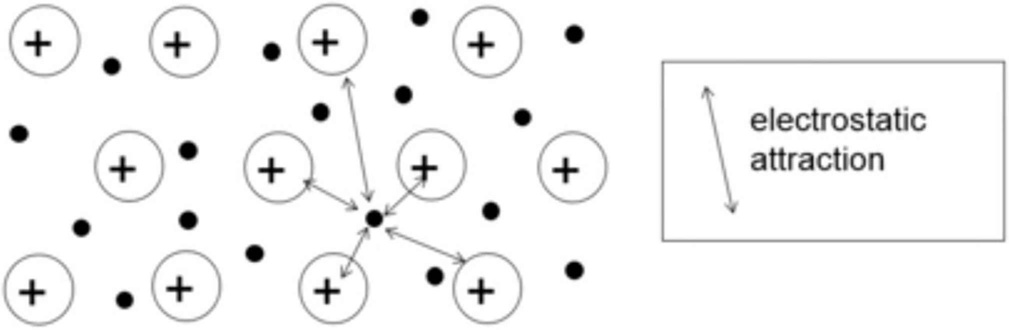
Ionic compounds are generally soluble in water and polar solvents (e.g. Ethanol), insoluble in non-polar solvents (made of organic compounds).

- *They are soluble due to the formation of ion-dipole interactions*

Metallic bonding



- A metallic bond is the strong electrostatic forces of attraction **between the positively charged metal ions and negatively charged mobile valence electrons**
  - Lattice of metal cations are surrounded by a 'sea' of mobile (delocalised) valence electrons
  - When the cations are in the structure of the metal, they lose some electrons into the 'sea' of mobile (delocalised) valence electrons
- Metal atoms are electropositive in nature
  - As a result, they tend to lose one or more electrons easily to become cations (positive)
- Metallic bonds are **non-directional**



Metallic bonds are **non-directional**.

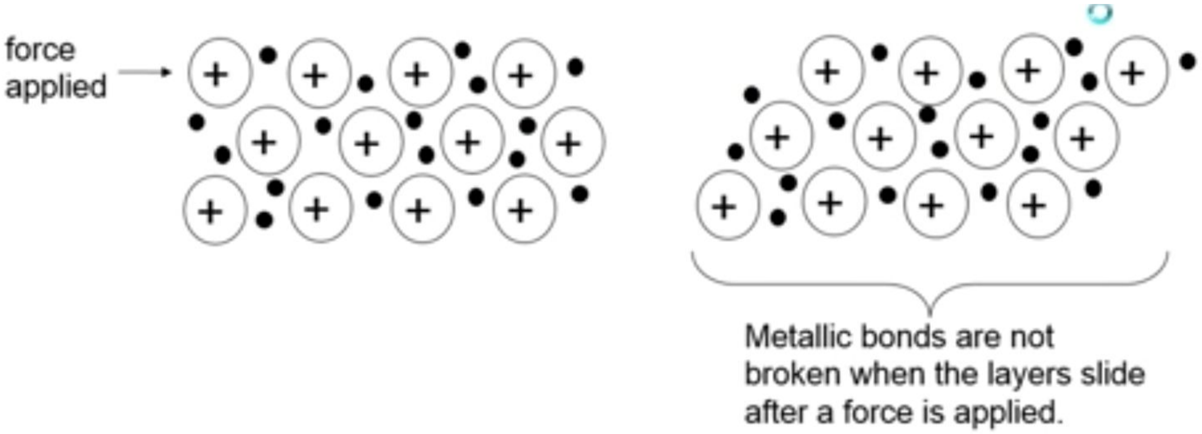
PHYSICAL PROPERTIES OF METALLIC BONDS

Typical metals are **lustrous (shiny)** and **sonorous (producing a ringing sound when hit).**

Metals are of **high density**.

- Closely packed, regular arrangement of metal cations surrounded by delocalised electrons
- Metals have high density as they have a high mass in a given volume

Metals are **malleable** and **ductile**.



Due to the non-directional nature of the bonds:

- Layers of cations slide over one another easily
- Allows layers of metal cations slide over one another without strong metallic bonds

Metals have a **high boiling and melting point**

- A large amount of energy is required to overcome the strong electrostatic forces of attraction between the negatively charged mobile (delocalised) electrons

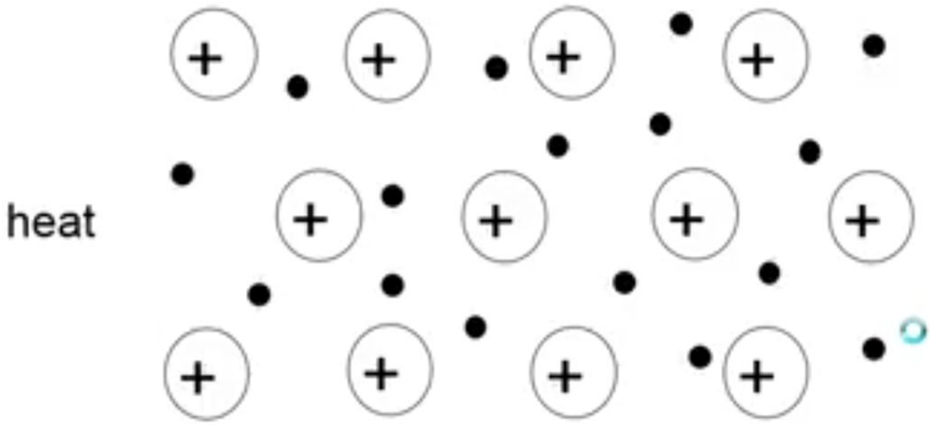
Metals are **insoluble**.

Metals are **good conductors of electricity**.

- Presence of mobile (delocalised) valence electrons which can migrate freely

through the metallic structure when a potential difference is applied. These mobile (delocalised) valence electrons are able to work as mobile charge carriers.
- Note that the metallic ions are in fixed positions, and are not used to conduct electricity

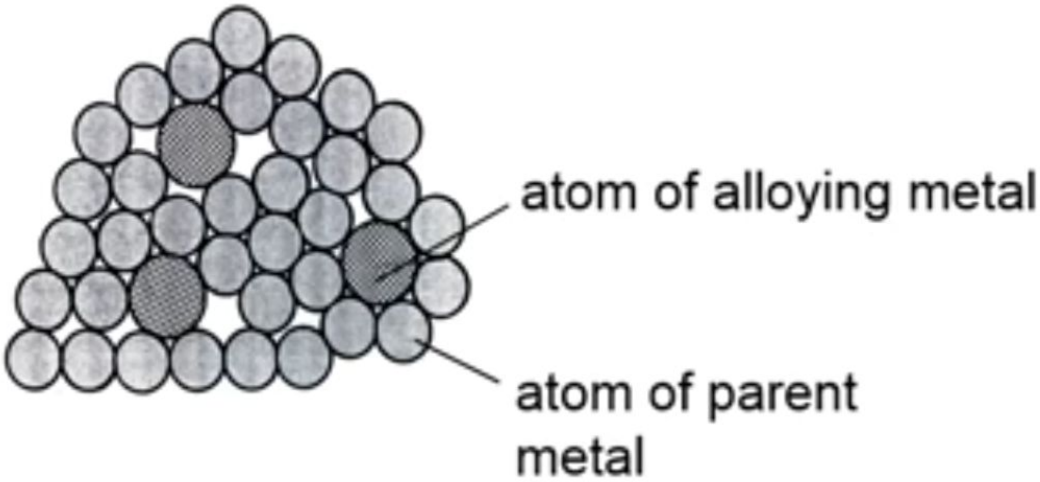
Metals have **good thermal conductivity**.



- The mobile (delocalised) valence electrons at the heated end gain energy and move more rapidly
- Mobile (delocalised) electrons will gain kinetic energy, vibrate faster, collide with surrounding electrons, and move through the 'sea'
- There is a lot of mobile (delocalised) electrons, thus heat is transferred very fast

ALLOYS

Alloy: A mixture of a metal and small amounts of other metals/ non-metals



Different atoms of metals/non-metals have different sizes.

- When a force is exerted, the layer of atoms cannot slide over each other as easily due to the 'obstruction' by the other atoms of different sizes
- This disrupts the orderly arrangement of the layers
- Alloys are hence **less malleable, less ductile, and harder**

Common Alloys	Parent Metal	Alloying Elements	Special properties	Uses
Amalgam	Hg	Sn	Durable	Tooth fillings
Brass	Cu	Zn	Resistant to corrosion	Electric contracts, musical instrument
Bronze	Cu	Sn	Hard and resistant to corrosion	Medals, bells, statues
Coinage metal	Cu	Ni	Strong and durable	Making coins
Constantan	Cu	Ni	Electrical conductivity	Thermocouple
Duralumin	Al	Cu, Mg, Mn	Strong and light	Aircraft bodies

Pewter	Sn	Sb, Cu	Shiny and beautiful	Ornaments, souvenirs
Solder	Pb	Sn	Lower melting point	Joining metals
Stainless steel	Fe	Cr, Ni	Strong and good corrosion/chemical resistance	Cutlery
Steel	Fe	C	Hard and strong	Building structures, bridges

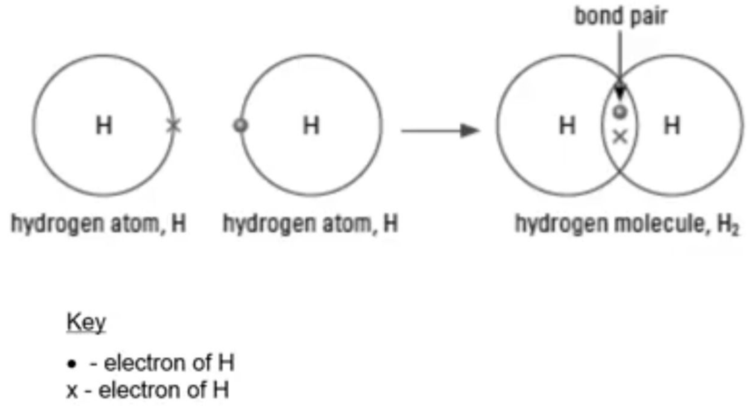
Covalent bonding

A covalent bond is the focus of attraction between the shared electrons and the 2 positive nuclei of the atoms.

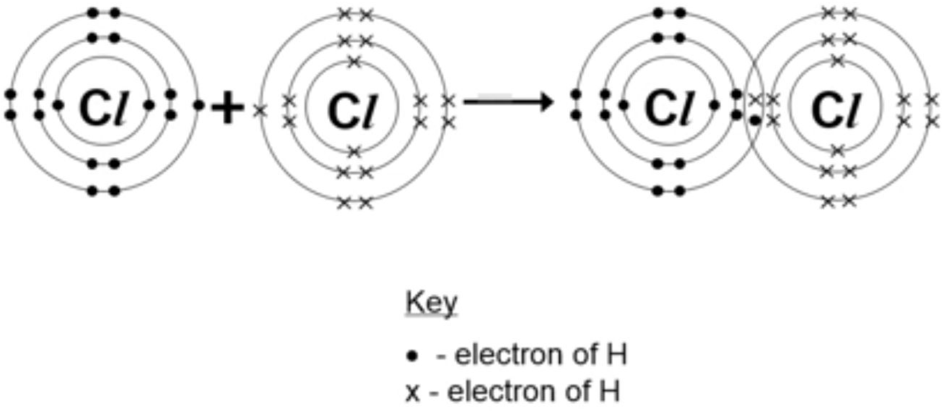
- Bonds are between non metals

DOT AND CROSS DIAGRAM

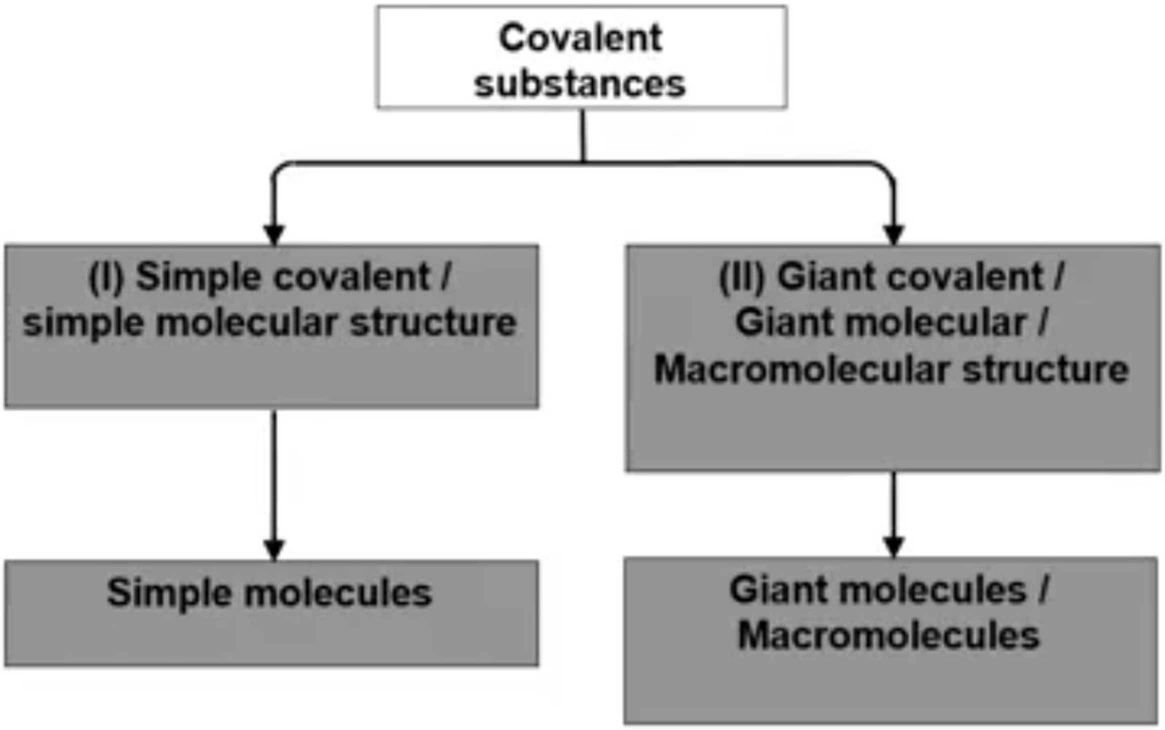
Hydrogen, H_2



Chlorine, Cl_2



STRUCTURE OF COVALENT SUBSTANCES



Substance with simple covalent structure	Substance with giant covalent structure
Soft A small amount of energy is required to overcome the weak intermolecular forces holding the discrete molecules together	**Strong** A large amount of energy is required to overcome the strong covalent bonds between atoms
Low melting and boiling points A small amount of energy is required to overcome the weak intermolecular forces holding the discrete molecules together	**High melting and boiling points** A large amount of energy is required to overcome the strong covalent bonds between atoms

Soluble in solvents of similar polarity Soluble in organic solvents (e.g. oil, hydrocarbons), insoluble in water. ● Non polar substances are soluble in non polar solvents ○ Hexane (non polar substance) dissolves well in octane (non polar solvents) ● Polar substances soluble in polar solvents ○ Water (polar substance) dissolves well in ammonia (polar solvent)	**Insoluble in all solvents (both organic solvents and water)**
Non conductor of electricity Absence of mobile charge carriers	**Non conductor of electricity (except graphite)** No mobile charge carriers present

Van der Waals forces of attraction: Collective intermolecular forces between simple molecules.

SIMPLE MOLECULES

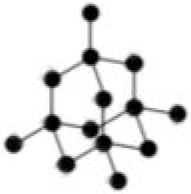
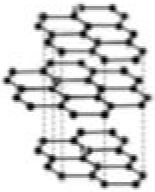
Simple molecules have a **simple molecular structure**, which consists of small, discrete covalent molecules.

- Molecules of elements = H_2, O_2, Diatomic molecules, etc.
- Molecules of covalent compounds = HCl, CO_2, CH_4
- The weak intermolecular forces are easily overcome. Hence substances with simple covalent structures have low melting and boiling points

GIANT STRUCTURES

All the atoms in substances with giant covalent structure are held tightly by tight covalent bonds

- Does not form a giant molecules, just exists as a giant covalent structure, molecules are simply simple and discrete

Giant covalent substances made up of **carbon atoms**	 Diamond	 Graphite
Structure	• Giant covalent structure • Each carbon atom is joined to 4 other carbon atoms by strong covalent bonds to form a tetrahedral arrangement (ie. Three-dimensional covalent network structure)	• Giant covalent structure • Each carbon atom is joined to 3 other carbon atoms by strong covalent bonds to form a network of hexagons • The hexagons are arranged in layers thus resulting in a layered structure • Weak van der waals forces hold the layers together • The one free electron from graphite gives graphite its special properties: good electrical conductor
Properties	Hardest natural substance known and has abrasive character **Why?** All the carbon atoms are held	Soft and slippery **Why?** Layers of carbon can easily slide

	together by very strong covalent bonds into a giant molecular structure. A large amount of energy is needed to break these very strong covalent bonds	over one another, since the weak van der waals forces of attraction between the hexagonal layers can be overcome by a small amount of energy
	High melting point (3500ºC) High boiling point (4800ºC) **Why?** All the carbon atoms are joined by very strong covalent bonds forming a giant covalent structure. A large amount of energy is needed to overcome the strong electrostatic forces of attraction between all carbon atoms	High melting point Able to conduct electricity **Why?** Each carbon atom only makes use of 3 electrons for bonding, bonds to three other carbon atoms and has one outermost electron not used to form covalent bonds. The electrons not used in bond formation are mobile (delocalized) between the layers. They can move along the layers and thus can conduct electricity
Uses	● Jewelry ● In saws and drills for cutting solids	● Pencil lead ● Lubricant

Both the structures of diamond and graphite consist of only atoms. They are carbon in different forms but the same physical state and are known are allotropes.

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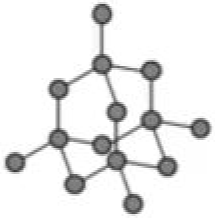
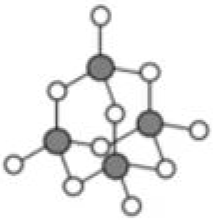
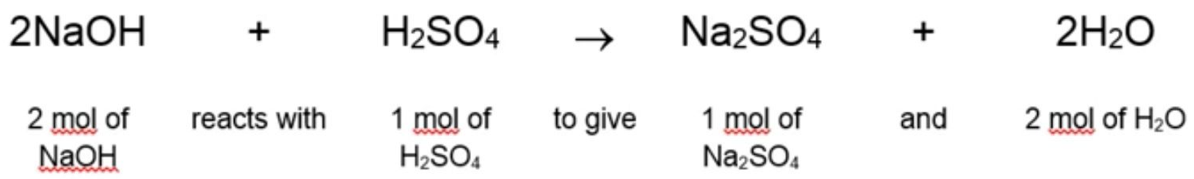
Giant covalent substances made up of **silicon atoms and its oxygen atoms**		
	Silicon	Silicon dioxide/ Silicon (IV) oxide
Structure	● Giant covalent structure ● Each silicon atom is joined to 4 other silicon atoms ● By strong covalent bonds to form a tetrahedral arrangement (ie. Three-dimensional covalent network structure)	Each silicon atom is joined to 4 oxygen atoms and each oxygen atom is joined to 2 silicon atoms by strong covalent bonds to form a tetrahedral arrangement
Properties	Able to conduct electricity at high temperatures (with the insertion of 'impurity' such as phosphorus atoms, which can take in >8 electrons, hence, this allows silicon to conduct electricity) ● Semiconductor	Hard and has abrasive character (unable to conduct electricity since all electrons are used in bonding)
	High melting point (1410ºC) High boiling point (2360ºC)	High melting point (1710˚C)
Uses	● Chips ● Microp-electronics	● Sand (Glass) ● Quartz

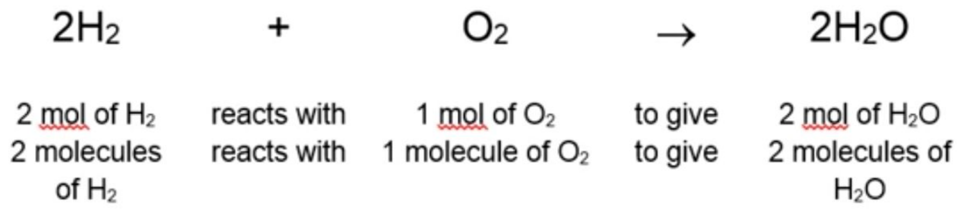
<u>Do note that the structure is covalent, not the molecule</u>.

CHEMICAL EQUATIONS

READING CHEMICAL EQUATIONS



$2NaOH$	+	H_2SO_4	→	Na_2SO_4	+	$2H_2O$
2 mol of NaOH	reacts with	1 mol of H_2SO_4	to give	1 mol of Na_2SO_4	and	2 mol of H_2O



$2H_2$	+	O_2	→	$2H_2O$
2 mol of H_2 2 molecules of H_2	reacts with reacts with	1 mol of O_2 1 molecule of O_2	to give to give	2 mol of H_2O 2 molecules of H_2O

For chemical calculations, we always work with the number of moles (or amount).

CALCULATION FROM EQUATIONS

Calculate the mass of water produced during the complete combustion of 100g of hydrogen gas.

1. Write balanced equation

$2H_2 + O_2 \rightarrow 2H_2O$

2. Work with whatever numbers you have, in this case, H_2 gas. The chemical equation shows molar equivalents, we need to convert the mass of H_2 gas to the number of moles of H_2 gas.

Mass (g) / number of moles (mol) = molar mass

100/2 = 50 mol of H_2 gas

3. Use the balanced chemical equation to determine the number of moles of water produced, hence determining the mole ratio.

2 mol of H2 will form 2 mol of H2O

Water = 50 mol

∴50*18 = 900g

CALCULATION FROM EQUATION - VOLUME

The volume of gas produced in a reaction is proportional to the amount of moles. This **only applies for gases**.

- To find the **number of moles, divide the volume (in dm^3)by 24**
- This applies throughout the equation, so there's no need to divide at all
- Cannot use volumes to determine when the product (reactant) is not in gaseous state
- 1 dm^3 = 1000 cm^3

CALCULATION FROM EQUATION - LIMITING REAGENTS

The number of moles of a product formed during a chemical reaction is determined by the number of moles of the reactant which is **not in excess**.

- This reactant is known as the **limiting reagent**

150g of magnesium is combusted in a container of 50000cm^3 of oxygen (at R.T.P) to form magnesium oxide. Calculate the mass of magnesium oxide formed.

1. Write balanced equation

2Mg + O2 → 2MgO

2. Determine limiting reagent

150/24 = 6.25 mol of Mg

50000/1000/24 = 2.0833 mol of O_2 gas

2 mol of Mg = O_2

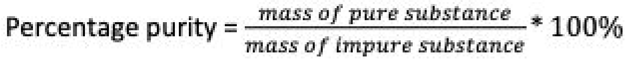
4.16 mol of Mg = 2.08 mol of O_2

∴Mg is in excess, O_2 is limiting reagent

3. 4.16 * (24.3 + 16) = 167.648g of MgO

PERCENTAGE PURITY

Percentage purity =



$$\text{Percentage purity} = \frac{mass\ of\ pure\ substance}{mass\ of\ impure\ substance} * 100\%$$

Do not count the number of moles of an impure substance.

100g of calcium carbonate was reacted with excess dilute sulfuric acid to form 20g of carbon dioxide. What is the percentage purity of the sample of calcium carbonate?

1. Write balanced equation

$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2$

2. Find number of moles

20/12+16*2 = 0.4545 mol of CO_2

By mol ratio, there is 0.4545 mol of calcium carbonate.

3. Determine mass of initial product

0.454 * (40+12+16*3) = 45.5g (Pure $CaCO_3$)

4. Calculate percentage purity

45.5/100 * 100 = 45.5%

PERCENTAGE YIELD

Percentage yield = $\frac{actual\ mass\ of\ pure\ product}{theoretical\ mass\ of\ pure\ product} * 100\%$

= $\frac{actual\ number\ of\ moles\ of\ pure\ product}{theoretical\ number\ of\ moles\ of\ pure\ product} * 100\%$

100g of calcium carbonate was reacted with excess dilute sulfuric acid to form 30g of carbon dioxide. What is the percentage yield of calcium dioxide?

1. Write balanced equation

$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2$

2. Find number of moles

100 / 40.1+12+16*3 = 0.999 mol of $CaCO_3$

By mole ratio, there is 0.999 mol of CO_2

0.999 * (12+16*2) = 43.956g of CO_2

3. Find percentage yield

Percentage yield = 30/43.96 * 100% = 68.2439%

DETERMINING THE EQUATION FROM EMPIRICAL DATA

We can use the mass of reactants in an experiment to determine the chemical equation of a reaction.

In an experiment, 6.2g of phosphorus (P_4) with 4.8g of oxygen (O_2)to produce an oxide of phosphorus. Determine the equation of this reaction.

1. Convert all values to moles as we cannot work with mass in a chemical equation.

6.2/31*4 = 0.05 mol of phosphorus

4.8/16*2 = 0.15 mol of oxygen

2. Compare the mole ratio to determine the equation

0.05 mol of phosphorus + 0.150 mol of oxygen → PO

0.05/0.05 = 1

0.150/0.05 = 3

∴Equation is $P_4 + 3O_2 \rightarrow P_4O_6$

CALCULATIONS INVOLVING TITRATIONS

Titration allows us to find the concentration of a solution by reacting it with a solution of known concentration.

- Calculations are similar to calculations involving concentrations

If 25 cm^3 of aqueous NaOH is required to completely neutralise 20 cm^3 of 0.5 mol/dm^3 of H_2SO_4, what is the concentration of aqueous NaOH?

1. Write balanced equation

$2NaOH + H2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

2. Find moles

20/1000 * 0.5 = 0.01 mol of H_2SO_4

By mole ratio, 2 mol of NaOH = 1 mol of H_2SO_4

0.01*2 = 0.02 mol of NaOH

3. Find concentration

0.02 / 25 = 0.0008 mol/dm^3

STOICHIOMETRY AND THE MOLE CONCEPT

RELATIVE ATOMIC MASS

The relative atomic mass (A_r) of an element is an average of one atom of the element when compared to 1/12 of the mass of an atom of carbon-12. Essentially, the mass number of an atom (Proton + Neutron number). There are no units.

Calculating A_r of naturally occurring chlorine, made up of 75% of chlorine-35 and 25% chlorine-37.

(37* 25%) + (35*75%) = 35.5

RELATIVE MOLECULAR MASS

Relative **molecular** mass (M_r) of a molecule is the average mass of one molecule of a substance when compared to 1/12 of the mass of one atom of carbon-12. Relative **formula** mass (M_r) is the average mass of one formula unit of an ionic compound when compared to 1/12 of the mass of one atom of carbon-12.

To put it simply in English, relative molecular mass is the combined atomic mass of all atoms in a molecule. It does not have any units.

PERCENTAGE

Percentage composition of an element = $\frac{Ar\ of\ the\ element}{Mr\ of\ the\ compound} * 100\%$

Percentage mass of an element in a compound =

$$\frac{Ar\ of\ the\ element}{Mr\ of\ the\ compound} * mass\ of\ sample$$

THE MOLE CONCEPT

The mole is used to describe/measure the amount of particles (atom, ion, electron, molecule).

- It is a SI unit
- Unit is **mol**
    - 1 mol = $6*10^{23}$ particles
    - The **Avogadro's constant** is $6*10^{23}$ of particles in 1 mole of any substance

Number of moles = $\frac{\text{number of particles}}{6 * 10^{23}}$

Molar mass (M) is the mass of one mole of any substance.

- Units are in **g/mol**
- E.g. Molar mass of H_2O = 1* 2 + 16 = 18 g/mol

Number of moles = $\frac{\text{mass (in g)}}{\text{molar mass}}$

MOLAR GAS VOLUME

1 mole of any gas occupies a volume of **24 dm^3** at R.T.P.

- Avogadro's law states that equal volume of gasses under the same conditions of temperature and pressure will contain the same number of particles
    - E.g. $100cm^3$ of H_2 (g) will contain the same number of particles as $100cm^3$ of H_2O
- 1 dm^3 = 1000 cm^3

Number of moles = $\frac{volume\ of\ gas\ (dm3)}{24}$

CONCENTRATION

The concentration of a solution tells us the number of moles of solutes per unit volume of solution.

- Units can be in **g/dm^3** or **mol/dm^3**

Find the concentration of an aqueous solution in mol/dm^3 of NaOH if 50g of NaOH is dissolved in 1500cm^3 of water.

1. Number of moles of NaOH = 50/(23+16+1) = 1.25 mol
2. 1500 cm^3 = 1.5 dm^3
3. Concentration of NaOH = 1.25/1.5 = 0.8333 mol/dm^3

ACIDS AND BASES

Acids

ACID

Acid: A substance which dissociates/ionises to produce Hydrogen ions (H^+) when it is dissolved in water.

- $HX\ (aq) \rightarrow H^+\ (aq) + X^-\ (aq)$

Acids do not behave like an acid unless **water** is present. **Water** is required for the formation of the hydrogen ions.

TYPES OF ACIDS

Mineral acids (made from chemical compounds from the earth)	**Chemical formula**
Hydrochloric acid	HCl
Nitric Acid	HNO_3
Sulfuric Acid	H_2SO_4
Carbonic Acid	H_2CO_3
Phosphoric Acid	H_3PO_4
Organic Acids (from plants and animals)	-

PROPERTIES OF ACIDS

Sour taste

pH < 7 at 25ºC (R.T.P.)

- Turns blue litmus paper red
- Turns universal indicator orange/red

Reacts with **metals** to give a **salt** and **hydrogen gas**

- $Mg (s) + 2HCl (aq) \rightarrow MgCl_2 (aq) + H_2 (g)$
- The hydrogen will extinguish a lighted splint with a “pop” sound
- **Exception**: When dilute HCl is reacted with lead, a later of insoluble Lead (II) chloride is formed which surrounds the lead, hence the reaction cannot continue.

Reacts with **carbonates** and **hydrogen carbonates** to give a **salt**, **water**, and **carbon dioxide gas**

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

Reacts with **metal oxides and hydroxides** to give a **salt** and **water**

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

Can conduct electricity due to the presence of **mobile charge carriers**

- HCl forms $H^+ (aq) + Cl^- (aq)$ when it dissociates in water
- These ions act as mobile charge carriers

PURE ACIDS

Pure acids (in the absence of water) consists of s**mall covalent molecules.** In the presence of water, the **molecules dissolve and forms ions**. This process is known as **ionisation/dissociation**.

BASICITY OF AN ACID

Basicity: Maximum number of hydrogen ions in one molecule of acid that can be replaced by a metal.

Type of acid	**Examples**
Monobasic	HCl, HNO_3
Dibasic	H_2SO_4, H_2SO_3
Tribasic	H_3PO_4

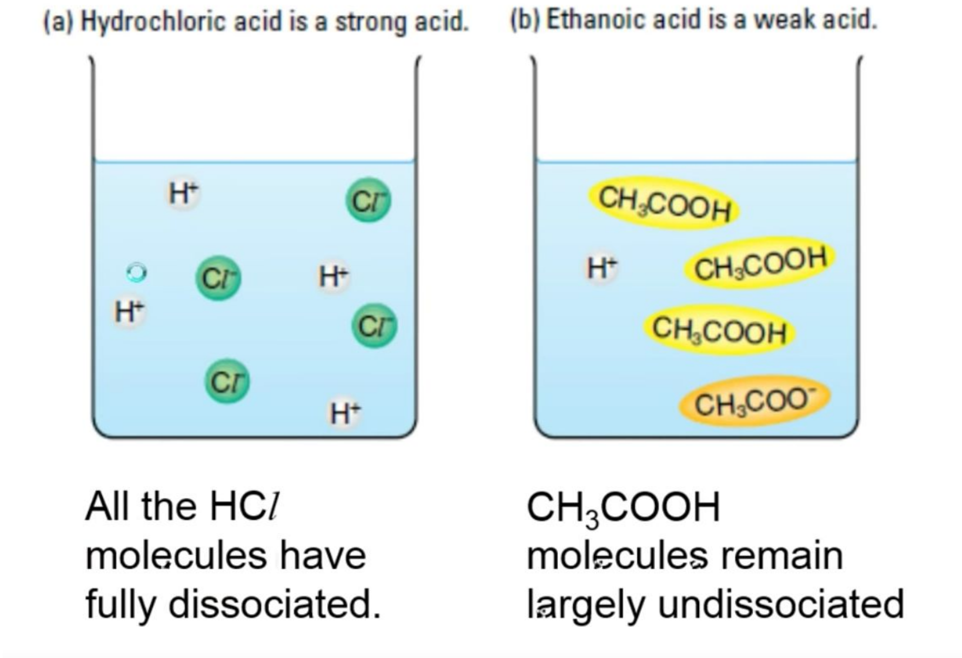
STRENGTH OF AN ACID

Strong acids are fully dissociated in water.

- $HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$
    - 1 mol HCl, 1 mol H^+, 1 mol Cl^-
- Examples of strong acids: **HCl, HNO_3, H_2SO_4**
- *Note: Strong acids react more vigorously than weak acids in the same reactions.*

Weak acids are partially dissociated in water.

- $CH_3COOH\ (aq) \rightleftharpoons CH_3COO^-\ (aq) + H^+\ (aq)$
    - 1 mol CH_3COOH, <1 mol CH_3COO^- , <1 mol H^+
- Examples of weak acids: **CH_3COOH** (Ethanoic acid)



CONCENTRATION OF AN ACID

Concentration: Refers to the amount of acid per unit volume of solution (mol/dm^3)

Alkalis

ALKALI

Alkali: An alkali is a soluble base which dissociates/ionizes to produce hydroxide (OH^-) ions in water.

- $Na_2O + H_2O \rightarrow 2NaOH$
- $NaOH \rightarrow Na^+ + OH^-$

PROPERTIES OF ALKALIS

Bitter taste

Feel **slippery/soapy**

- Alkali can dissolve the thin layer of oil on your skin surface which makes you feel like there's less friction when you're rubbing your fingers together.

pH > 7 (R.T.P.)

- Turns red litmus paper blue
- Turns universal indicator blue/violet

Can conduct electricity due to the presence of **mobile charge carriers** (ions)

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

Reacts with **acids** to form **salt** and **water**.

- $KOH\ (aq) + HCl\ (aq) \rightarrow KCl\ (aq) + H_2O\ (l)$
- <u>Neutralisation</u> reaction

Reacts with **ammonium compounds** on **heating**, to give a **salt**, **water** and **ammonia gas**.

- $OH^{-} + NH_4^{+} \rightleftharpoons NH_3 + H_2O$
- Ammonia (NH_3) ≠ Ammonium ion (NH_4^{+})
- Test for ammonia gas with damp red litmus paper

Forms **insoluble metal hydroxide precipitate** when added to solutions of salts of many metals

- Precipitate reaction
- $CuSO_4$ (aq) + 2NaOH (aq) → $Cu(OH)_2$ (s) + Na_2SO4 (aq)

TYPES OF BASES

Type of base	**Example(s)**
Monobasic/ Monoprotic	NaOH, KOH
Dibasic/ Diprotic	$Ca(OH)_2$, $Mg(OH)_2$
Tribasic/ Triprotic	$Al(OH)_3$

STRENGTH OF AN ALKALI

Strong alkalis undergo complete dissociation (ionisation) to form hydroxide ions (OH^{-}) in water.

- NaOH (aq) → Na^{+} (aq) + OH^{-} (aq)

Weak alkalis undergo partial dissociation (ionisation) to form hydroxide ions (OH^{-}) in water.

- NH_3 (aq) + H_2O (l) $\rightleftharpoons$ NH_4^{+} (aq) + OH^{-} (aq)

NEUTRALISATION

All neutralisation reactions have the same ionic equation.

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

Indicators and the pH scale

INDICATORS

An indicator is a substance that has different colours in acidic and alkaline solutions.

Indicator	Colour in acids	Colour at change	Colour in alkali
Phenolphthalein	Colourless	Colourless (pH 7)	Pink
Methyl orange	Red	Orange (pH 3.7)	Yellow
Screened methyl orange	Red	Grey (pH 7)	Green
Litmus	Red	Maroon (pH 6.5)	Blue
Universal Indicator	Red/orange, depending on pH	Green (pH 7)	Blue/violet, depending on pH

PH AND PLANT GROWTH

The pH of the soil will affect the way plants absorb nutrients from the soil.

- **Acidic** soil can be treated with bases such as quicklime (calcium oxide, CaO) or slaked lime (Calcium hydroxide, $Ca(OH)_2$.
- **Alkaline** soil can be treated with ammonium sulfate ($(NH_4)_2SO_4$)

Oxides

NEUTRAL OXIDES

Neutral oxides are **non-metallic oxides** that exhibit neither basic or acidic properties. They **do not react** with acids or bases.

- Water (H_2O)
- Carbon monoxide (CO)
- Nitrogen monoxide (NO)

ACIDIC OXIDES

Acidic oxides are **oxides of non-metals**. Most of them **dissolve in water** and form **acids**. They also react with **alkalis to form salt and water** only.

- Carbon dioxide (CO_2) produces carbonic acid (H_2CO_3) in water
- Nitrogen dioxide (NO_2) produces nitric acid (HNO_3) in water
- Sulfur dioxide (SO_2) produces sulfurous acid (H_2SO_3) in water
- Sulfur trioxide (SO_3) produces sulfuric acid (H_2SO_4) in water
- Phosphorus (V) oxide (P_4O_{10}) produces phosphoric acid (H_3PO_4) in water
- Silicon dioxide (SiO_2) is insoluble in water

BASIC OXIDES

Basic oxides are **oxides of metals**. They are **mostly insoluble in water**, while those that **dissolve in water** form **alkalis**. They react with **acids to form salt and water only**.

- Sodium oxide (Na_2O) produces sodium hydroxide ($NaOH$) in water
- Potassium oxide (K_2O) produces potassium hydroxide (KOH) in water
- Calcium oxide (CaO) produces calcium hydroxide ($Ca(OH)_2$) in water
- Copper (II) oxide (CuO) is insoluble in water
- Magnesium oxide (MgO) is insoluble in water

AMPHOTERIC OXIDES

Amphoteric oxides are **metallic oxides** (ZAP) which can behave as **both an acidic and basic oxide**.

- **Zinc** oxide (ZnO)
- **Aluminium** oxide (Al_2O_3)
- **Lead (II)** oxide (PbO)

Writing ionic equations

An **ionic equation** is an equation involving ions in an aqueous solution. You will need to know the chemical equation.

STEPS

1. Write chemical equation (balanced)
    a. $NaOH (aq) + HCl (aq) \rightarrow NaCl (aq) + H2O (l)$
2. Dissociate into ions
    a. $Na+ (aq) + OH- (aq) + H+ (aq) + Cl- (aq) \rightarrow Na+ (aq) + Cl- (aq) + H2O (l)$
    b. H2O cannot dissociate thus it remains the same
3. Cancel out the like ions on both sides
    a. $Na+ (aq) + OH- (aq) + H+ (aq) + Cl- (aq) \rightarrow Na+ (aq) + Cl- (aq) + H2O (l)$
    b. $OH- (aq) + H+ (aq) \rightarrow H2O (l)$

Ions that do not contribute to the reaction are called spectator ions

SALTS

SOLUBILITY OF COMPOUNDS

Compounds	**Remarks**
Nitrates (NO_3^-)	All are soluble
Chlorides (Cl^-)	All are soluble except **$AgCl$, $PbCl_2$, $HgCl_2$**
Sulfates (SO_4^{3-})	All are soluble except **$PbSO_4$, $BaSO_4$, $CaSO_4$**
Carbonates (CO_3^{2-})	All are insoluble except **ammonium and group I salts**
Hydroxides (OH^-)	All are insoluble except **ammonium and group I salts**
Iodides (I^-)	All are soluble except **$AgCl$, $PbCl_2$, $HgCl_2$**
Bromides (Br^-)	All are soluble except **$AgCl$, $PbCl_2$, $HgCl_2$**

PREPARATION OF INSOLUBLE SALTS

Insoluble salts are prepared by **precipitation**.

- One of the reactants must be a soluble salt

Salt	**Solubility**	**Steps**
Barium sulfate	Insoluble	1. Identify solubility of salt 2. Identify acid to use 3. Identify other reactant 4. Add excess BaCl2 with dilute H2SO4 until no more precipitate forms 5. Filter the mixture to obtain precipitate 6. Wash with distilled water 7. Press between sheets of filter paper to dry

Reasons for each of the steps:

1. Barium sulfate BaSO4 is insoluble in water
2. BaSO4 contains sulfate anion -> use sulfuric acid (H2SO4)
3. Replace H+ ion in sulfuric acid with Ba2+. The salt we want is insoluble, thus we need to ensure that at the end of the reaction, there is only one insoluble compound, hence we can choose Barium Chloride (BaCl2) or Barium Nitrate

Barium Chloride

$BaCl_2$ (aq) + H_2SO_4 (aq) → $BaSO_4$ (aq) + HCl (aq)

Ba^{2+} (aq) + SO_4^{2-} (aq) → $BaSO_4$ (aq)

Barium Nitrate

$Ba(NO,)_2$ (aq) + H_2SO_4 (aq) → $BaSO_4$ (aq) + HNO_3 (aq)

PREPARATION OF SOLUBLE SALTS

Soluble salts can be prepared by **titration** or by **adding reactants in excess**.

The **adding reactants in excess** method is used when one of the reactants is an insoluble carbonate, oxide, or metal.

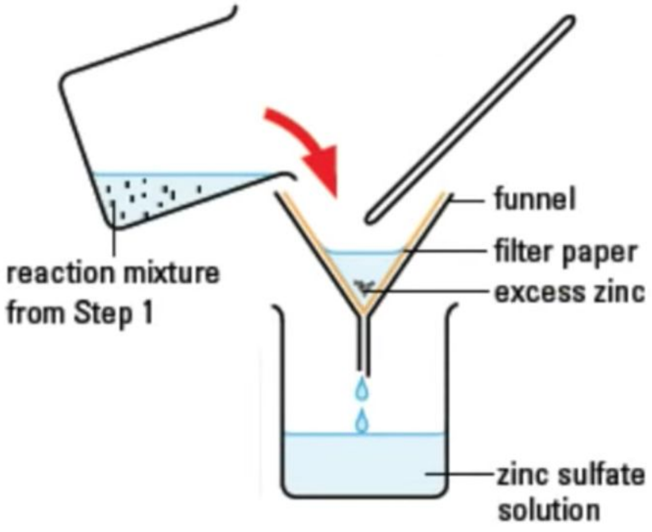
- Only moderately reactive metals (zinc, magnesium) are suitable for this method

Salt	**Solubility**	**Steps**
Zinc Sulfate ($ZnSO_4$)	Soluble (excess)	1. Identify solubility of salt 2. Identify method of preparation 3. Identify other reactant 4. Add excess $ZnSO_4$ with warm dilute H_2SO_4 5. Remove excess unreacted Zn by filtration, $ZnSO_4$ is obtained in the filtrate 6. Heat filtrate (indicate the reagent for the filtrate) to

		evaporate off excess water 7. Filter out crystals from the solution 8. Wash crystals with small volume of cold distilled water 9. Press dry between sheets of filter paper

Reasons for each of the stepa:

1. $ZnSO_4$ is soluble in water
2. $ZnSO_4$ contains Zn^{2+} which can be derived from Zn metal. Zinc is a moderately reactive metal. SO_4^{2-} can be derived from H_2SO_4. Hence an acid-metal reaction is used
3. Excess zinc powder is added to ensure that all H_2SO_4 is reacted. Warm H_2SO_4 is used so as to speed up reaction



4. Heat to saturation, leave solution to cool and crystalize
5. To wash off impurities
6. To dry crystals

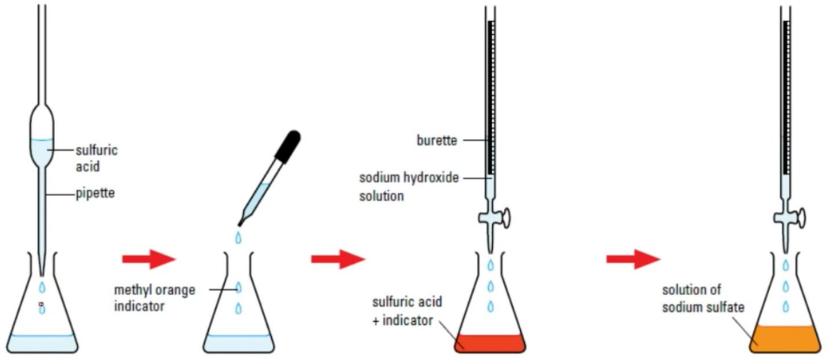
Equation

$Zn\ (s) + H_2SO_4\ (aq) \rightarrow ZnSO_4\ (aq) + H_2\ (g)$

The **titration** method is used when both reactants are aqueous, such as these reactions:

- Acid with soluble metal carbonate
- Acid with alkali

Salt	**Solubility**	**Steps**
Sodium sulfate (Na_2SO_4)	Soluble (titration)	1. Identify solubility of salt 2. Identify method of preparation ***Prepare for PTSD ahead (important)*** 3. Using a pipette, transfer 25.0 cm^3 of dilute H_2SO_4 (aq) into clean conical flask 4. Add 2-3 drops of methyl orange (pH indicator). a. 1. Red in acid b. 2. Orange in Alkali 5. Fill burette (the long one with a valve at the bottom) with NaOH (aq). Release alkali into the conical flask slowly. a. Take note of initial Volume Reading b. Swirl flask 6. Continue until ONE drop of the alkali turns the methyl orange indicator from red to orange. a. This is the end point 7. Read the volume of acid used in burette a. Calculate how much acid is used i. Final volume - Initial volume (usually start reading at 0.0cm3) = volume used 8. Repeat steps 3 to 7 but without methyl orange indicator. 9. Resulting solution will contain Na_2SO_4+ H_2O 10. Solution is heated to evaporate off excess water (or until saturated) 11. Cool solution for crystals to form 12. Filter solution and collect crystals of Na_2SO_4 13. Wash solid crystals with cold deionised water (to remove water-soluble impurities) 14. Press dry between sheets of filter paper and leave to dry

		If you didn’t read that: 

Reasons for each of the steps:

1. Na_2SO_4 is soluble in water
2. Na_2SO_4 contains sulfate anion, hence we use sulfuric acid. The sodium (Na^+) ion cannot be obtained by reacting Na with H_2SO_4, because it will explode like the Note 7 and then you'll die but then you won't need to come for the O’ levels and your parents can claim insurance so maybe it’s a good thing…?
    a. So we use a reactant containing Na^+
        i. NaOH (aq), alkali
            1. $NaOH$ (aq) + H_2SO_4 → Na_2SO_4 (aq) + H_2O (l)
3. Which goes back to the point above, if we use excess of either NaOH (aq) or H2SO4, we are unable to determine when both reactants have reacted completely and will be left with the excess reactant. Both aqueous substances cannot be separated.

Equation

$2NaOH$ (aq) + H_2SO_4 (aq) → Na_2SO_4 (aq) + H_2O (l)

OH^- (aq) + H^+ (aq) → H_2O (l) (Neutralisation reaction)

QUALITATIVE ANALYSIS

TESTING CATIONS

Sodium hydroxide (NaOH) and aqueous ammonia ($NH_{4,}$)is used to identify the positive ions (cations) and negative ions (anions) in a given solution.

Using sodium hydroxide to find cations:

Ion	**Added dropwise**	**Added in excess**
Ca^{2+}	**White** precipitate of $Ca(OH)_2$ is formed Ionic Equation: Ca^{2+} (aq) + $2OH^-$ (aq) → $Ca(OH)_2$ (s)	White precipitate **remains**
Zn^{2+}	**White** precipitate of $Zn(OH)_2$ is formed Ionic Equation: Zn^{2+} (aq) + $2OH^-$ (aq) → $Zn(OH)_2$ (s)	White precipitate **dissolves to form colourless solution**
Al^{3+}	**White** precipitate of $Al(OH)_3$ is formed Ionic Equation: Al^{3+} (aq) + $2OH^-$ (aq) → $Al(OH)_3$ (s)	White precipitate **dissolves to form colourless solution**
Pb^{2+}	**White** precipitate of $Pb(OH)_3$ is formed Ionic Equation: Pb^{2+} (aq) + $2OH^-$ (aq) → $Pb(OH)_2$ (s)	White precipitate **dissolves to form colourless solution**
Fe^{2+}	**Green** precipitate of $Fe(OH)_2$ is formed Ionic Equation” Fe^{2+} (aq) + $2OH^-$ (aq) → $Fe(OH)_2$ (s)	Green precipitate **remains**
Fe^{3+}	**Reddish brown** precipitate of $Fe(OH)_3$ is formed (formed on standing of $Fe(OH)_2$)	Reddish brown precipitate **remains**

	Ionic Equation Fe^{3+} (aq) + $3OH^-$ (aq) → $Fe(OH)_3$ (s)	
Cu^{2+}	**Blue** precipitate of Cu(OH)2 is formed Ionic Equation Cu^{2+} (aq) + $2OH^-$ (aq) → $Cu(OH)_2$ (s)	Blue precipitate **remains**
NH_4^-	**No precipitate formed** Ammonia gas (NH_3) is released when heated	-

Using **aqueous ammonia** to find cation:

Ion	**Added dropwise**	**Added in excess**
Ca^{2+}	**No precipitate is formed**	
Zn^{2+}	**White** precipitate of $Zn(OH)_2$ is formed Ionic Equation: Zn^{2+} (aq) + $2OH^-$ (aq) → $Zn(OH)_2$ (s)	White precipitate **dissolves** With excess ammonia, zinc hydroxide reacts to form a soluble complex ion, tetra-amine (organic chemistry) zinc ion, which is colorless $Zn(OH)_2$ (s) + $4NH_3$ (aq) → $Zn(NH_3)_4^{2+}$ (aq) + $2OH^-$ (aq)
Al^{3+}	**White** precipitate of $Al(OH)_3$ is formed Ionic Equation: Al^{3+} (aq) + $2OH^-$ (aq) → $Al(OH)_3$ (s)	White precipitate **remains**
Pb^{2+}	**White** precipitate of $Pb(OH)_3$ is formed Ionic Equation: Pb^{2+} (aq) + $2OH^-$ (aq) → $Pb(OH)_2$ (s)	White precipitate **remains**

Fe^{2+}	**Green** precipitate of $Fe(OH)_2$ is formed Ionic Equation” Fe^{2+} (aq) + $2OH^-$ (aq) → $Fe(OH)_2$ (s)	Green precipitate **remains**
Fe^{3+}	**Reddish brown** precipitate of $Fe(OH)_3$ is formed (formed on standing of $Fe(OH)_2$) Ionic Equation Fe^{3+} (aq) + $3OH^-$ (aq) → $Fe(OH)_3$ (s)	Reddish brown precipitate **remains**
Cu^{2+}	**Blue** precipitate of Cu(OH)2 is formed Ionic Equation Cu^{2+} (aq) + $2OH^-$ (aq) → $Cu(OH)_2$ (s)	Blue precipitate **dissolved to form dark blue solution**

Note: Pb^{2+} and Al^{3+} cannot be differentiated by using NaOH and NH3, use ions (specifically potassium iodide or potassium chloride) to distinguish. Expected observations?

TESTING ANIONS

When determining the identity of cations:

- Test for gas produced (NO_3^-, CO_3^{2-})
- Observe colour of precipitate (Cl^-, I^-, SO_4^{2-})

Ion	**Reagents used**	**Observations**
NO_3^-	1. Add aqueous NaOH 2. Add aluminum powder/foil 3. Warm gently 4. Use damp red litmus paper to test for alkaline gas (ammonia)	Damp red litmus paper turns blue in pungent, colourless alkaline ammonia (NH_3) gas released

CO_3^{2-}	1. Add any acid a. Nitric acid to remove alkalis (all soluble in nitrate) 2. Use limewater to test for gas released (carbon dioxide)	Effervescence is observed in solution and white precipitate is formed in limewater when colourless, odourless carbon dioxide (CO_2) produced is bubbled into it
CL^{-}	1. Add dilute nitric acid a. To remove any alkalis/impurities 2. Add aqueous silver nitrate solution	White precipitate of AgCl is formed Ionic equation Ag+ (aq) + Cl- (aq) → AgCl (s)
I^{-}	1. Add dilute nitric acid 2. Add aqueous Lead (II) Nitrate solution or Silver Nitrate solution	Yellow precipitate of Lead(II) Iodide is formed or yellow precipitate of silver iodide Ionic equation $2I^{-}$ (aq) + Pb^{2-} (aq) → PbI_2 (s) or I^{-} (aq) + Ag^{+} (aq) → AgI (s)
SO_4^{2-}	1. Add dilute nitric acid 2. Add aqueous barium nitrate solution	White precipitate of Barium Sulfate is formed Ionic equation SO_4^{2-} (aq) + Ba^{2+} (aq) → $BaSO_4$ (s)

TESTING GASES

Gas	Testing agent	Observations
NH_3	Damp red litmus paper at mouth of test tube	**Colorless, pungent gas is present. Effervescence is observed. Damp** Red litmus paper turns blue Alkaline: Red turn Blue
CO_2	Limewater	White precipitate is formed and dissolves in excess CO_2 **Effervescence is observed, Colorless, odorless gas is present. CO_2 is present.**

Cl_2	Damp blue litmus paper	**Effervescence is observed. Yellowish-green, pungent gas** turns Litmus paper is bleached white
H_2	Lighted splint	**Effervescence is observed. Colourless odourless gas is present.** Lighted splint is extinguished with a pop sound
O_2	Glowing splint	**Colorless, odorless gas is present.** Effervescence is observed. Glowing splint relights
SO_2	Place a drop of acidified potassium manganate (VII) on a piece of filter paper, then place on gas	**Colorless, pungent gas is present. Effervescence is observed.** Acidified potassium manganate (VII) turns from purple to colorless

SPEED OF REACTION

MEASURING RATE OF REACTION

There are 3 main ways we can measure the rate of a reaction:

- Measuring the amount of product formed
- Measuring the amount of reactant used
- Observing changes which occur during a reaction

COLLISION THEORY

For 2 or more particles to react and form a product, they have to meet and collide with:

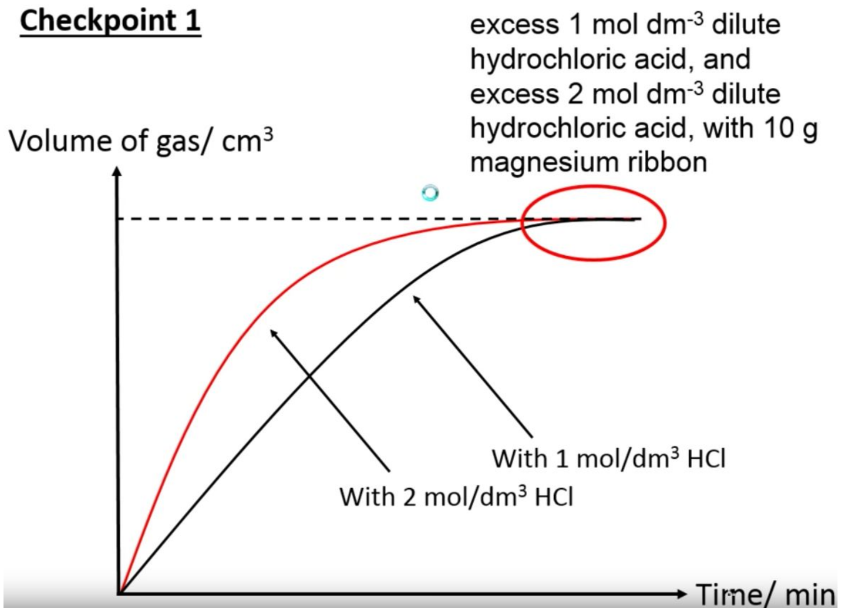
- **Sufficient energy to overcome activation energy**
    - The more energy provided to the reaction mixture, the more likely it is for the particle to collide with energy equals to or greater than the activation energy
- **Correct orientation**
    - The more collisions per unit time, the faster the reaction will occur

FACTORS THAT AFFECT THE SPEED OF A REACTION

Concentration of particles

- THe higher the concentration, the more reactant particles per unit volume
- The frequency of collisions increase
- Frequency of effective collisions increase
- The speed of the chemical reaction increases

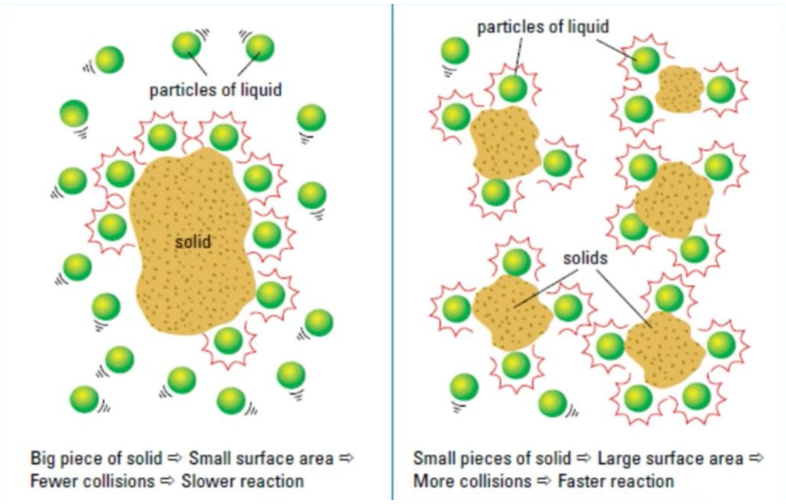
If there is 1 mol/dm^3 of HCl vs 2 mol/dm^3 of HCl as an excess reagent reacting with magnesium (limiting reagent), no matter how much HCl we add, the same amount of H_2 gas will be produced.



Note: steeper gradient for higher concentration

Size of reacting particles

- The larger the reacting particles, the larger the surface area to volume ratio
- More surface area is exposed for particles to collide
- Frequency of collision increases
- Frequency of <u>effective</u> collisions increases
- The speed of the chemical reaction increases



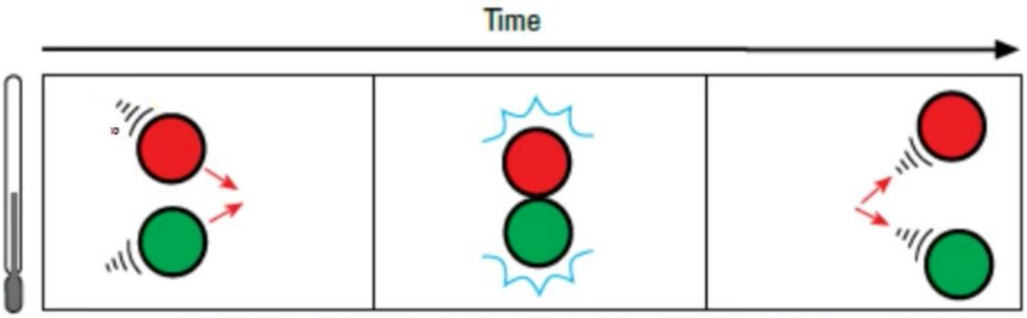
Pressure (Gases)

- When the pressure increases, reactant particles are closer to each other
- There are more particles per unit volume
- Frequency of collision increases
- Frequency of successful collision increases
- The speed of the chemical reaction increases

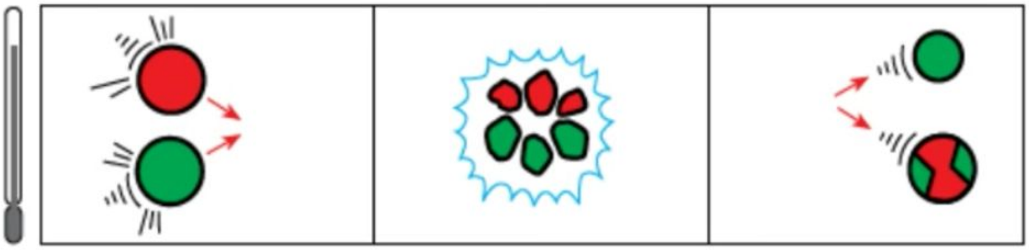
For example, the Haber process (where ammonia gas is involved) occurs at 450°C and 200 atm.

Temperature

- When the temperature increases, particles gain kinetic energy and move faster
- Increased frequency of collision with sufficient energy to overcome activation energy
- Frequency of successful collisions increases
- The speed of the chemical reaction increases



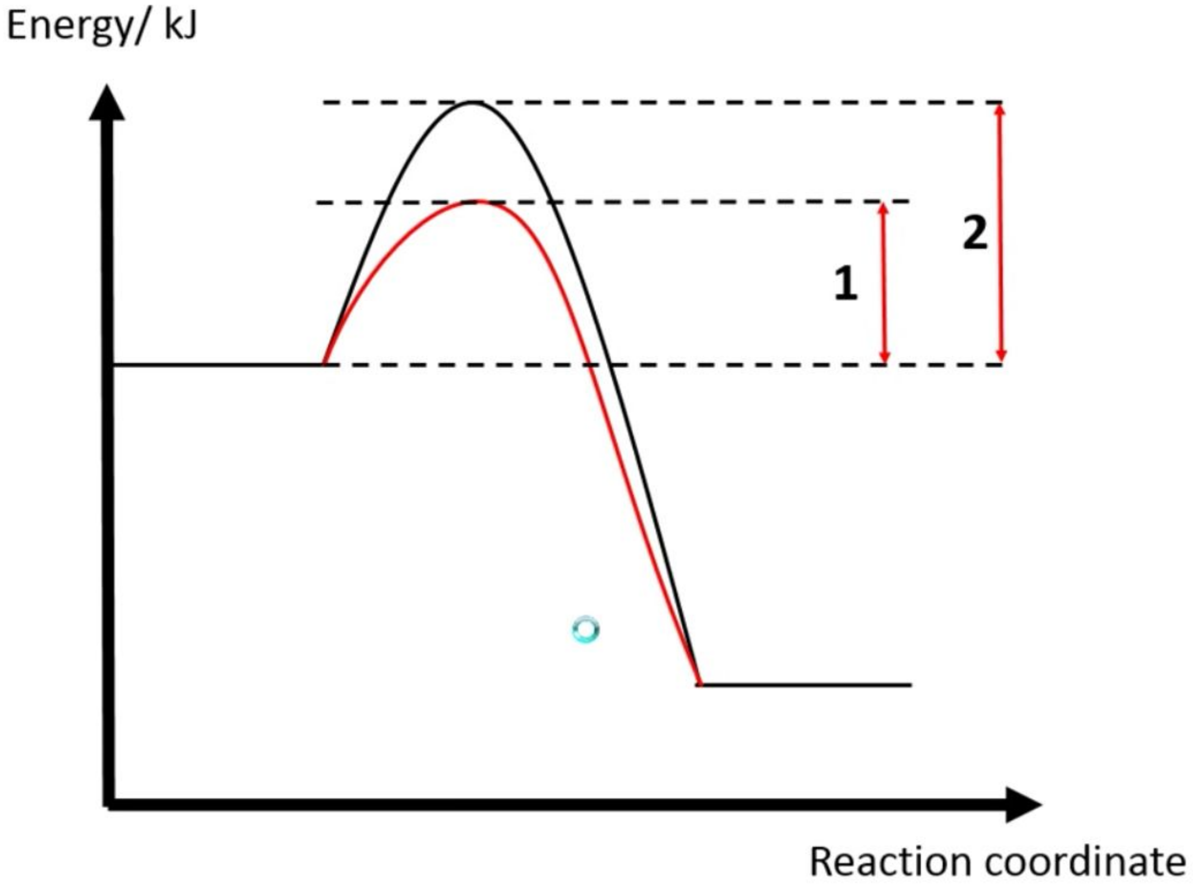
Particles bounce apart at low temperatures without reacting.



High-speed collisions at higher temperatures more often produce a reaction.

Presence of catalyst

A catalyst is a substance which increases the speed of a chemical reaction and remains chemically unchanged at the end of the reaction.



Note: X-axis should be "progression of reaction". (1) is the catalysed reaction, (2) is the uncatalysed reaction

- Catalysts provide an alternate pathway for the reaction with a lower activation energy
- More particles can collide with enough energy to overcome the activation energy
- Frequency of successful collisions increases
- The speed of the chemical reaction increases
- Catalysts DO NOT lower the activation energy of a reaction

Characteristics of a catalyst:

- Only a small amount is needed as it remains chemically unchanged

- It is not used up during the reaction
- Selective and different reactions require different catalysts
- Catalysts increase speed of reaction but not yield of reaction
- Catalysts can be poisoned by impurities and no longer work or become less effective
- Catalysts are usually transition metals or their compounds

ENERGY FROM CHEMICALS

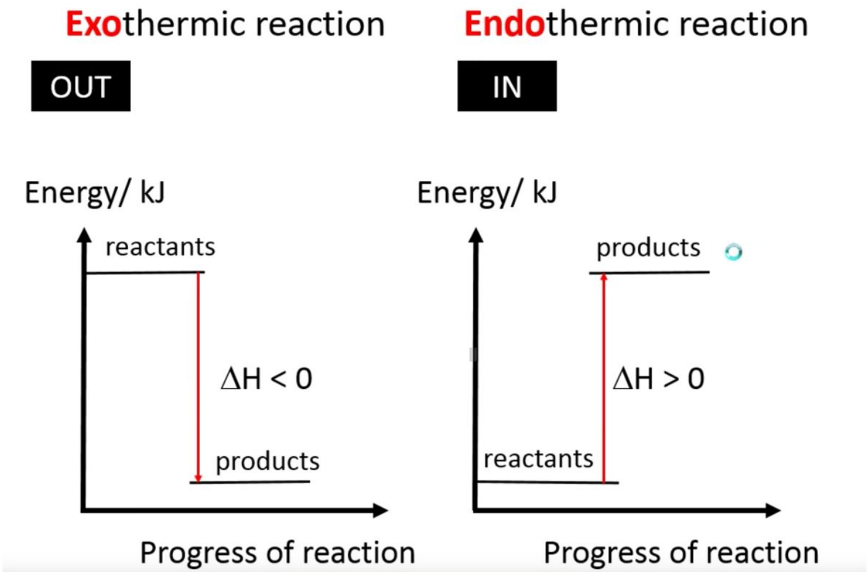
TYPES OF ENERGY CHANGES

Exothermic

- Energy is released out into the surroundings
- Products will have a lower energy level than reactants
    - Leads to a decrease in temperature

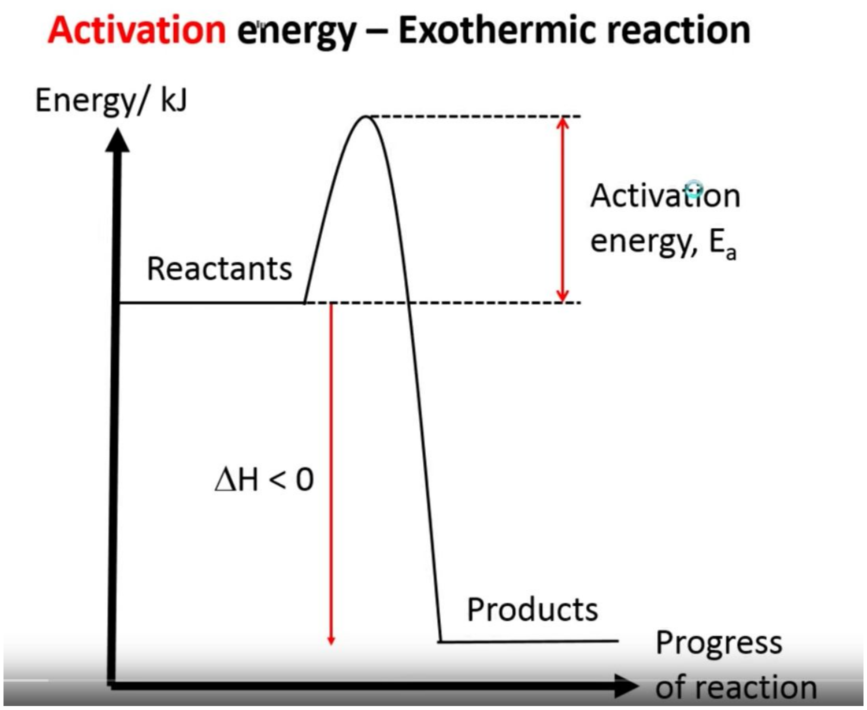
Endothermic

- Energy is absorbed from the surroundings
- Products will have a higher energy level than the reactants
    - Increase in temperature

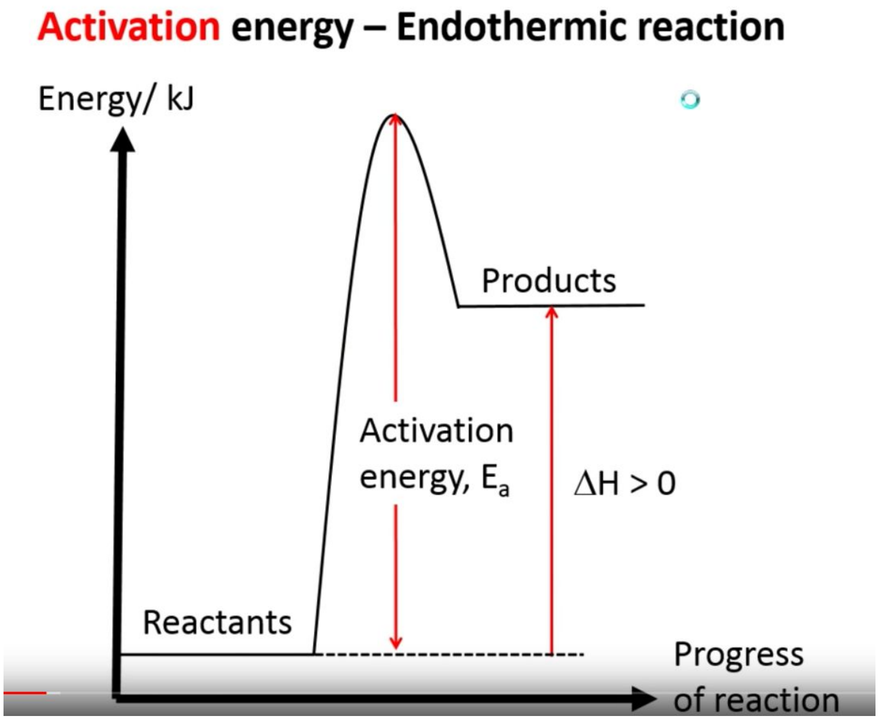


ACTIVATION ENERGY

For all reactions, bonds in the reactant particle must be broken before new bonds in the product can be formed.



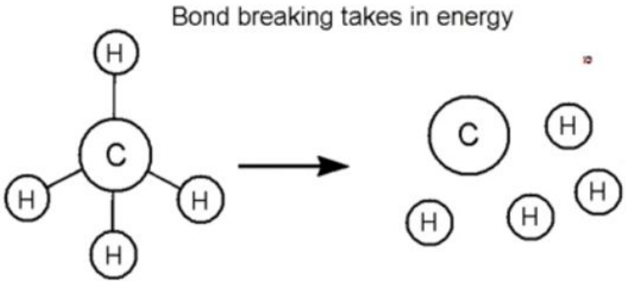
Activation energy – Exothermic reaction



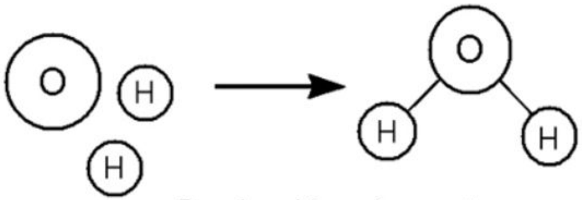
Activation energy – Endothermic reaction

BOND BREAKING AND FORMING

Bond breaking is when energy is absorbed to break bonds between the particles.



Bond forming is when energy is released when bonds are formed between particles. This may occur in the form of heat energy being released.



Bond making gives out energy

ENTHALPY CHANGE

Overall enthalpy change = total energy absorbed for bond breaking - total energy released during bond forming

The reaction is exothermic if there is less energy absorbed (during bond breaking) than released (during bond forming)

The reaction is endothermic if there is more energy absorbed (during bond breaking) than released (during bond forming).

When finding enthalpy change, determining the number of bonds to be broken, the number of bonds to be formed, and calculate from the table of bond energy (given).

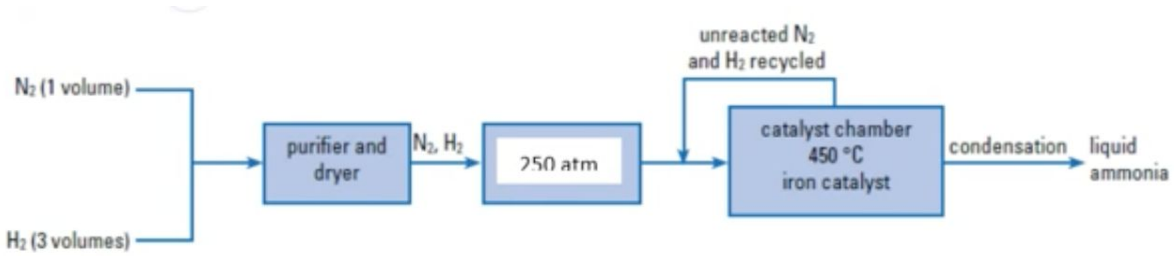
AMMONIA

REVERSIBLE REACTIONS

Reversible reactions (as implied in the name) are chemical reactions that can be reversed.

- Both backward and forward reactions take place at the same time
    - How much product depends on the rate of the forward reaction
    - How much reactant depends on rate of backward reaction
- Note that these reactions **do not go to completion**
    - Once there's a product, a backward reaction happens
- The amount of reactants and products in a reversible reaction can be changed by altering the conditions of the reaction
    - Temperature
    - Pressure
    - The use of a catalyst

HABER PROCESS



Conditions: **450˚C, 200atm, finely divided iron catalyst**

Equation: N_2 (g) + $3H_2$ (g) ⇋ $2NH_3$ (g)

- N_2: H_2 = 1:3 ratio of reactants
- Exothermic reaction
- Reversible reaction
- Only 10-15% of reactants are converted to ammonia

Nitrogen is obtained from the rational distillation of liquified air.

- Air has about 70% nitrogen

Hydrogen is obtained from the cracking of crude oil or steam reforming water

- Steam reforming: $CH_{4,}$ (g0 + H_2O (l) → CO (g) + $3H_2$ (g)
  - 700˚C to 1100˚C
  - Nickel catalyst

Why high pressure?

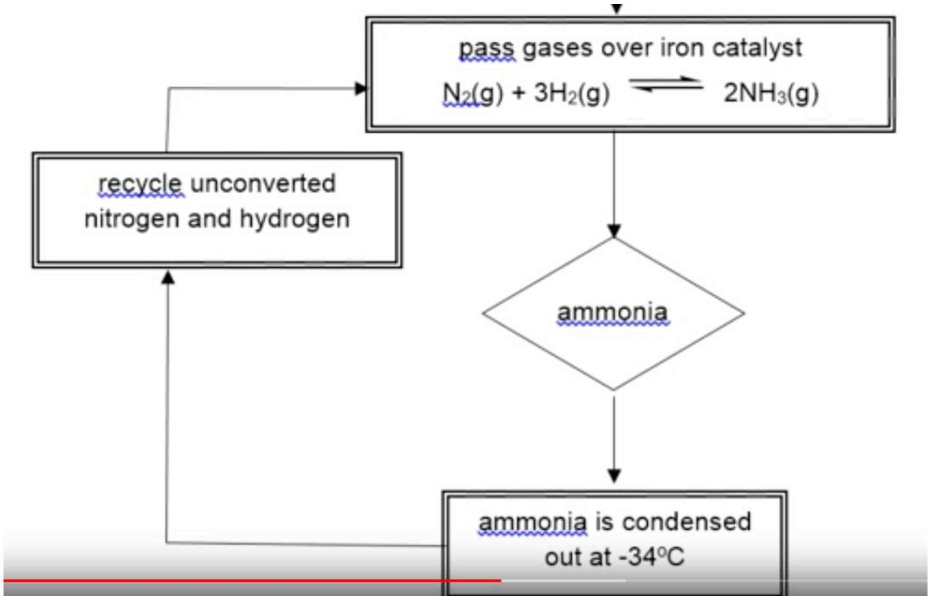
- Due to gaseous reactants we can increase the pressure to increase rate of reaction which increases rate of forward reaction
- Too high a pressure is costly

Why 450˚C high temperature?

- The lower the temperature, the higher the yield
- This is because the decomposition of ammonia through the backward reaction is reduced
- However, if the temperature is too low, it will slow down the forward reaction

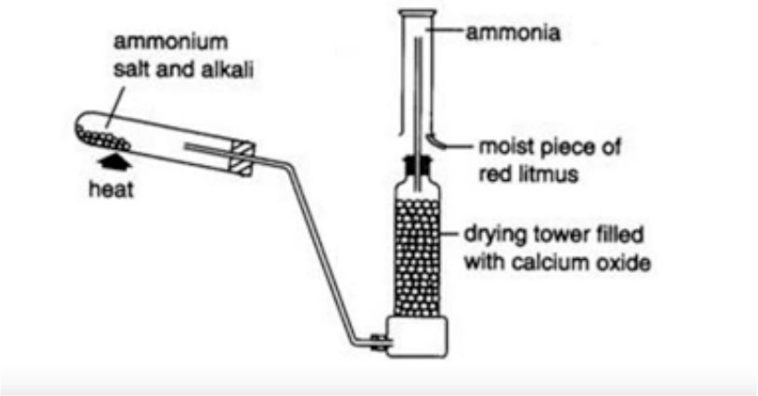
Why use a finely divided iron catalyst?

- Despite high pressure and temperature, the overall reaction is still slow
- Catalyst is used to speed up the reaction by lowering the activation energy through offering a alternative pathway with lower activation energy, and does not change the yield



OTHER REACTIONS FORMING AMMONIA

Ammonium compound with hydroxide reaction



The underlying reaction taking place in each case can be represented by the same ionic equation:

$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$

Ammonia gas is dried by passing it through calcium oxide:

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

Concentrated sulfuric acid cannot be used, because it reacts with ammonia gas:

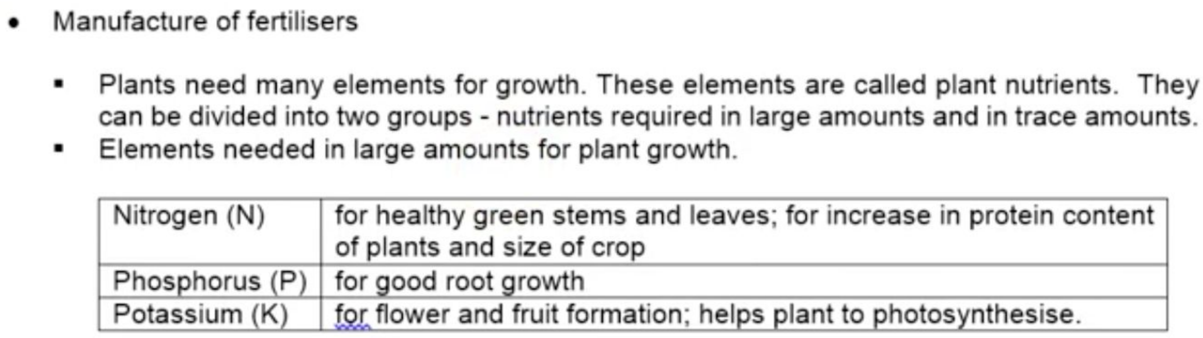
$2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$

Ammonia is collected by upward delivery because it is much lighter than air and soluble in water.

USES OF AMMONIA

Ammonia is used in the manufacture of fertilizers.

- Ammonium nitrate
- Ammonium sulfate



- Manufacture of fertilisers
  - Plants need many elements for growth. These elements are called plant nutrients. They can be divided into two groups - nutrients required in large amounts and in trace amounts.
  - Elements needed in large amounts for plant growth.

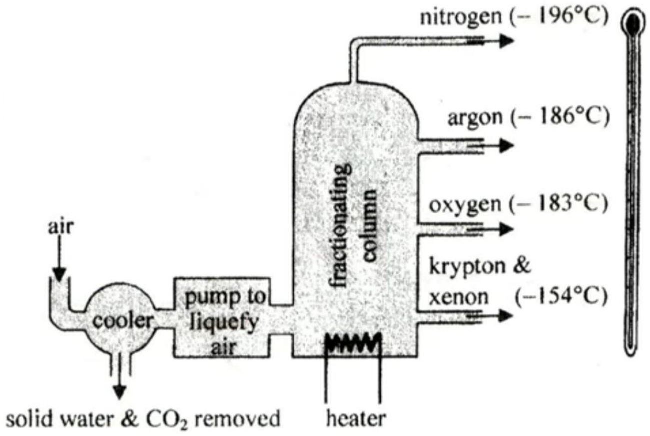
Nitrogen (N)	for healthy green stems and leaves; for increase in protein content of plants and size of crop
Phosphorus (P)	for good root growth
Potassium (K)	for flower and fruit formation; helps plant to photosynthesise.

AIR AND THE ENVIRONMENT

COMPOSITION OF AIR

Component	Percentage
Nitrogen	78%
Oxygen	21%
Carbon dioxide	0.03%
Noble gases	>1%

Air can be separated using fractional distillation.



POLLUTION

Carbon monoxide (CO)

Source: Incomplete combustion of carbon-containing fuels in vehicles

- Colorless, odorless, preventing easy detection
- Combines with haemoglobin to form stable carboxyhemoglobin, which prevents

the haemoglobin from transporting oxygen to cells, depriving them of oxygen
  - Essentially it just stops your ability to transport oxygen around your body so you die

Catalytic converters in vehicles remove carbon monoxide in vehicles by oxidising it into CO_2.

- $2CO + O_2 \rightarrow 2CO$,

Nitrogen oxides (NO, NO_2)

Source: Nitrogen and oxygen in air combine to form nitrogen monoxide at high temperatures in car engines / power stations / lightning strikes.

- Oxides of nitrogen are highly corrosive and toxic
- Oxides of NO_2 lead to the formation of acid rain and photochemical fog
- Irritates:
  - Lungs
  - Skin
  - Breathing passages
  - Eyes
- $NO_2 + H_2O \rightarrow HNO_3$
- $NO + H_2O \rightarrow HNO_2 \rightarrow$ oxidation $\rightarrow HNO_3$

Catalytic converters in vehicles remove oxides of nitrogen as N_2.

Sulfur dioxide

Source: Produced during combustion of fossil fuels such as coal and petroleum which contain sulfur as an impurity. It is also produced during volcanic eruptions and the extraction of iron in the blast furnace when the sulfur impurities are oxidised off.

- Acid rain
  - $SO_2 + H_2O \rightarrow H_2SO_3$
  - $2H_2SO_3 + O_2 \rightarrow 2H_2SO_4$
- Enter leaves and affect growth of plants
- Irritates eyes, lungs → breathing problems

Sulfur dioxide reacts with alkali to neutralise, removing sulfur dioxide.

- $CaO + SO_2 \rightarrow CaSO_3$

SO_2 is also removed from flue gases.

- Flue: a gas heater, a power station, or other fuel burning installation

Unburnt hydrocarbons

Source: Hydrocarbons in fuels that have not been burnt in vehicle engines.

- Carcinogenic
- React with sunlight and other pollutant to form photochemical smog

Methane

Source: Produced when plant and dead matter decay, as well as by cows and other farm animals.

- Contributes to global warming

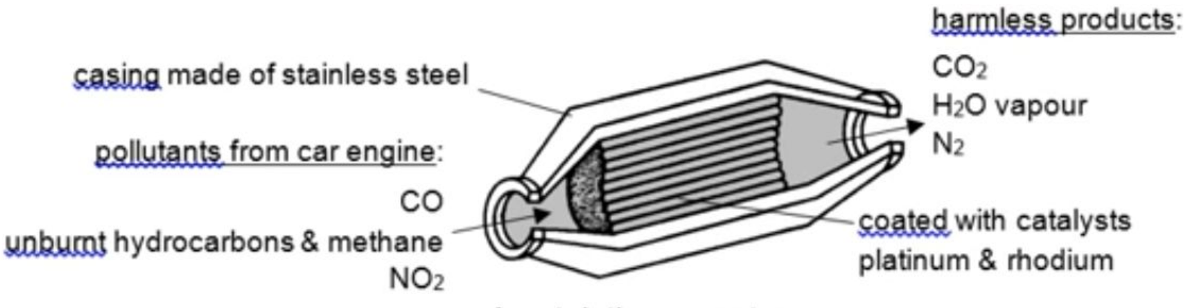
Ozone

Source:

- Irritates:
  - Eyes
  - Nose
  - Throat
- Asthma-inducing
- Damages plant crops
- Sunlights acts on air and other pollutants, producing photochemical smog

REMOVING HARMFUL GASES

Catalytic converters



A catalytic converter

- Can be fitted into motor vehicles to convert unburnt hydrocarbons and other harmful gases into harmless compounds
- Nitrogen oxides react with carbon monoxide in the presence of platinum and rhodium catalyst to form nitrogen and carbon dioxide
  - $2NO + 2CO \rightarrow N_2 + 2CO_2$
- Unburnt hydrocarbons are oxidised to form carbon dioxide and water
  - $2C_8H_{18} + 25O_2 \rightarrow 18CO_2 + 18H_2O$

Using cleaner fuels

- Use fuels with low sulfur content
- Changing to natural gas (methane), which burns to produce carbon dioxide and water

Reducing sulfur dioxide emissions in power stations

Flue gas desulfurization removes sulfur dioxide and nitrogen oxides from waste gases in power stations.

1. Calcium carbonate is added to hot flue gases. Heat causes calcium carbonate to decompose into calcium oxide and carbon dioxide.
   a. $CaCO_3 \rightarrow CaO + CO_2$
2. Calcium oxide reacts with sulfur dioxide to form calcium sulfite which is further

oxidised to form calcium sulfate (Neutral)

   a. $CaO + SO_2 \rightarrow CaSO_3$
   b. $2CaSO_3 + O_2 \rightarrow 2CaSO_4$

ACID RAIN

Acid rain comes from largely sulfur dioxide and nitrogen dioxide.

- Sulfur dioxide in the atmosphere dissolves in water to form sulfurous acid,
    - $SO_2 + H_2O \rightarrow H_2SO_3$
    - $2H_2SO_3 + O_2 \rightarrow 2H_2SO_4$
- Nitrogen dioxide in the atmosphere dissolves in water to form nitrous acid, which is further oxidised by oxygen to form nitric acid
    - $4NO_2 + 2H_2O + O_2 \rightarrow 4HNO_3$

Harmful effects of acid rain include:

- Makes soil more acidic, plants do not grow well
- Corrodes metal, cement and stone structures
- Removes important nutrients from soil

This can be prevented by:

- Catalytic converters convert oxides of nitrogen to harmless nitrogen
- Adding CaO or $Ca(OH)_2$ to lakes/soils to neutralise the acid

OZONE

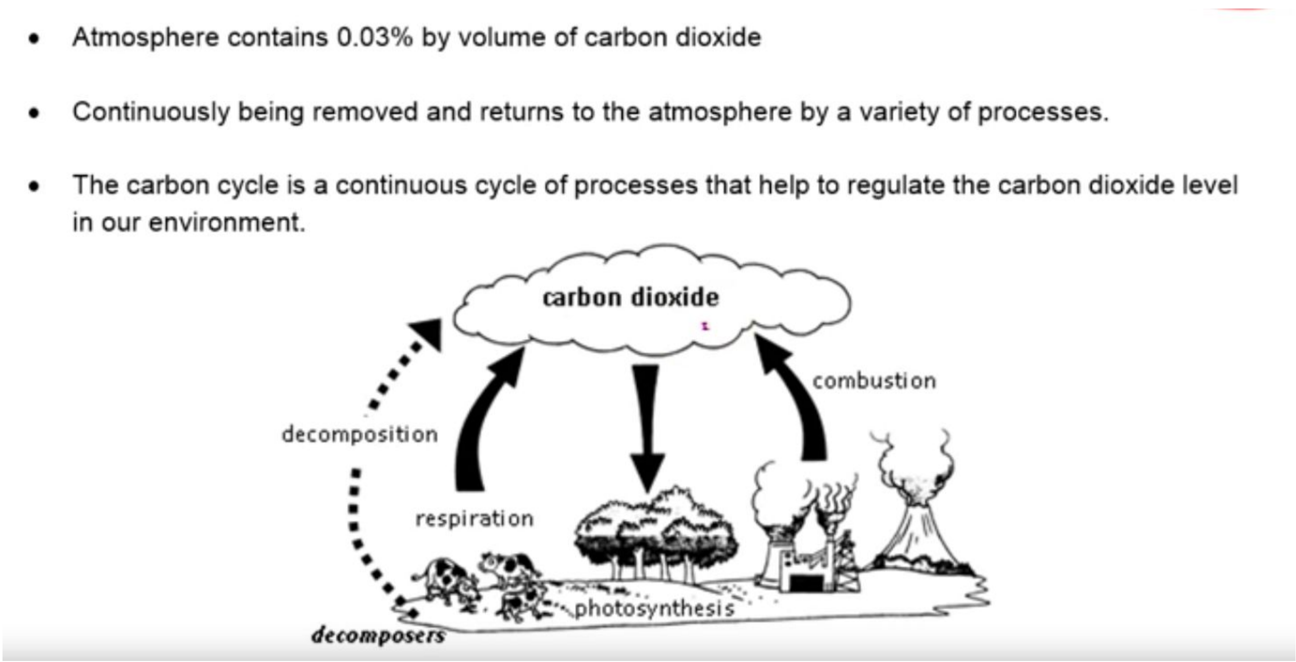
Ozone (O_3) is a pale blue and pungent gas formed in engines/motor vehicles when an electric spark passes through O_2 in the air.

- It is also produced when photochemical reactions happen between oxygen molecules and oxygen atoms in the lower atmosphere layer
    - $O_2 + O \rightleftharpoons O_3$
- Ozone surrounds the earth about 40km above the ground and absorbs UV rays from the sun

Chlorofluorocarbons (CFC) contain chlorine, fluorine and carbons.

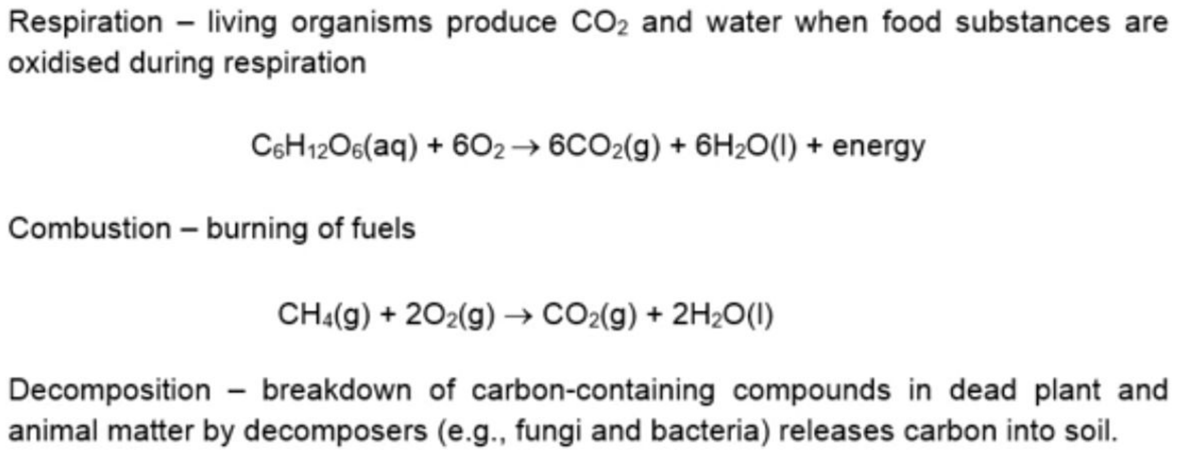
- They are unreactive and remain in the atmosphere for a long time
- Not soluble in water, hence not removed by rain
- They can form free chlorine atoms when exposed to UV light and then react with ozone removing it from the ozone layer
    - $O_3 \rightarrow O_2$
- CFCs were used in aerosol propellants and refrigerants
    - Banned since they are directly linked to the depletion layer

CARBON CYCLE



- Atmosphere contains 0.03% by volume of carbon dioxide
- Continuously being removed and returns to the atmosphere by a variety of processes.
- The carbon cycle is a continuous cycle of processes that help to regulate the carbon dioxide level in our environment.

Introducing CO_2 into the environment:



Respiration – living organisms produce CO_2 and water when food substances are oxidised during respiration

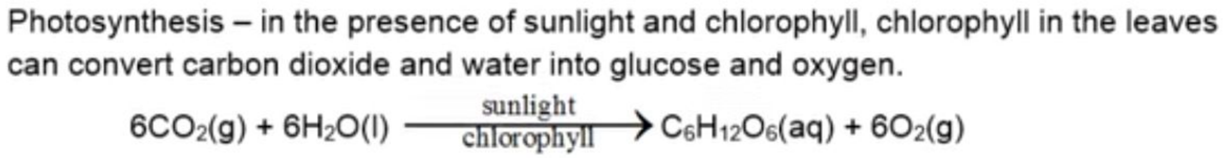
$$C_6H_{12}O_6(aq) + 6O_2 \rightarrow 6CO_2(g) + 6H_2O(l) + \text{energy}$$

Combustion – burning of fuels

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

Decomposition – breakdown of carbon-containing compounds in dead plant and animal matter by decomposers (e.g., fungi and bacteria) releases carbon into soil.

Removing CO_2 from the environment:



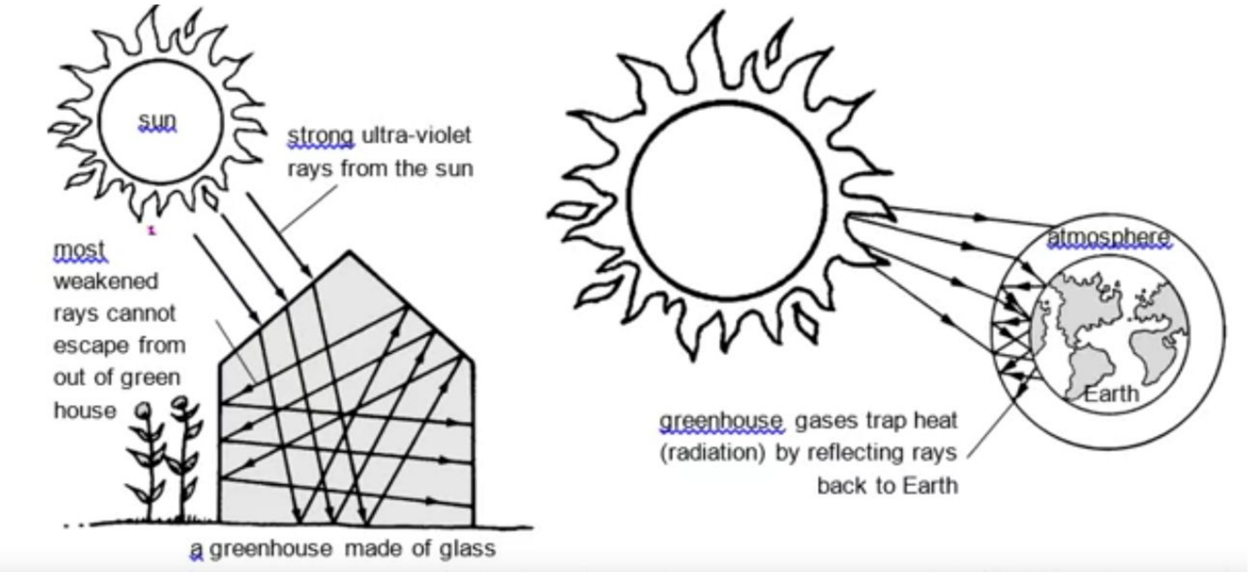
Photosynthesis – in the presence of sunlight and chlorophyll, chlorophyll in the leaves can convert carbon dioxide and water into glucose and oxygen.

$$6CO_2(g) + 6H_2O(l) \xrightarrow[\text{chlorophyll}]{\text{sunlight}} C_6H_{12}O_6(aq) + 6O_2(g)$$

GREENHOUSE EFFECT AND GLOBAL WARMING

Greenhouse gases include CO_2, water vapour, nitrous oxide, methane, as it traps heat energy from the sun as it is reflected back from the earth.

- Greenhouse effect is a natural process as it maintains a temperature suitable for sustaining life on earth.
- Due to the enhanced greenhouse effect, too much greenhouse gases are in the atmosphere
    - Results in excess heat getting trapped
    - Results in increase in global temperatures
    - Increase in sea level
        - Melting of polar ice caps
        - Floods and heavy rains → washes away fertile top soil leaving soil with very little nutrients


sun
strong ultra-violet rays from the sun
most weakened rays cannot escape from out of green house
a greenhouse made of glass
atmosphere
Earth
greenhouse gases trap heat (radiation) by reflecting rays back to Earth

REDOX

REDUCTION AND OXIDATION

Reduction and oxidation can be defined in terms of:

- Gain or loss in oxygen
- Gain or loss of hydrogen
- Gain or loss of electrons
- Increase or decrease in oxidation state

When both reduction and oxidation take place simultaneously (at the same time), they are known as **redox** reactions.

REDOX BY OXYGEN, HYDROGEN, ELECTRONS

Oxygen

A substance is **oxidised** when it **gains oxygen** in a chemical reaction.

A substance is **reduced** when it **loses oxygen** in a chemical reaction.

$Zn (s) + CuO (s) \rightarrow ZnO (s) + Cu (s)$

In the above reaction, zinc gains oxygen to form zinc oxide, hence zinc is oxidised; copper oxide loses oxygen to form copper, hence copper is reduced. As oxidation and reduction occurs simultaneously, it is a redox reaction.

Hydrogen

A substance is **reduced** when it **gains hydrogen** in a chemical reaction.

A substance is **oxidised** when it **loses hydrogen** in a chemical reaction.

$H_2S\ (g) + Cl_2\ (g) \rightarrow 2HCl\ (g) + S\ (s)$

In the above reaction, hydrogen sulfide loses hydrogen to form sulfur, hence it is oxidised; chlorine gas gains hydrogen to form hydrogen chloride, hence it is reduced. As oxidation and reduction occurs simultaneously, it is a redox reaction.

Electrons

A substance is **oxidised** when it **loses electrons** in a chemical reaction.

A substance is **reduced** when it **gains electrons** in a chemical reaction.

$2Na\ (s) + Cl_2\ (g) \rightarrow 2NaCl\ (s)$

In the above reaction, the sodium atom loses an electron, to form a sodium ion in sodium chloride (Na^+), hence it is oxidised; the chlorine atoms in chlorine gas gain an electron, to form a chlorine ion in sodium chloride (Cl^-). As oxidation and reduction occurs simultaneously, it is a redox reaction.

REDOX BY OXIDATION STATE

Oxidation state is a number. It is a charge an atom would have if it existed as an ion.

There are a few rules when finding and working with oxidation states:

1. **Elements have an oxidation state of zero**
    a. Cu, Fe, C, O_2, He
2. **The oxidation state of a simple ion is the charge on the ion**
    a. Oxidation state of Cl^- is -1
    b. Oxidation state of Cu^{2+} is +2
3. **The oxidation state of some elements in their compounds is fixed**
    a. Hydrogen in most of its compounds has oxidation state +1
        i. Exception: Hydrogen in peroxides have oxidation state -1 in compounds containing elements that are less electronegative than

hydrogen, as in CaH_2 (You will be told in the question)

   b. Oxygen in most of its compounds has oxidation state -2
      i. Exception: Oxygen in peroxides have oxidation state -1
   c. All Group I elements have oxidation state +1
   d. All group II elements have oxidation state +2
4. **The oxidation states of the elements in a compound add up to zero**
   a. Copper oxide (CuO): Cu (+2) + O (-2) = 0
5. **The sum of the oxidation states of the elements in an ion is equal to the charge on the ion**
   a. OH^-: O (-2) + H (+1) = -1

A substance is **oxidised** when it **increases in oxidation state** in a chemical reaction.

A substance is **reduced** when it **decreases in oxidation state** in a chemical reaction.

$Zn (s) + 2H^+ (aq) \rightarrow Zn^{2+} (aq) + H_2 (g)$

In the above reaction, oxidation state for zinc increases from 0 in Zn to +2 in Zn^{2+}, hence zinc is oxidised; oxidation state of hydrogen decreases from +1 in H^+ to 0 in H_2, hence it is reduced. As oxidation and reduction occurs simultaneously, it is a redox reaction.

OXIDISING AND REDUCING AGENTS

An **oxidising agent** is a substance that **causes oxidation**. Oxidising agents undergo reduction in a redox reaction.

A **reducing agent** is a substance that **causes reduction**. Reducing agents undergo oxidation in a redox reaction.

Note: Causes reduction ≠ is reducing

Reducing agents can be tested with an oxidising agent. The common (testing) oxidising agents include:

Name of compound	**Formula**	**Applications**
Potassium manganate (VII), acidified with dilute H_2SO_4	$KMnO_4$	Used to test for reducing agent; oxidises alcohol to acids; used to test for SO_2 gas Colour change when a reducing present is from **purple → colourless**
Potassium dichromate (VI), acidified with dilute H_2SO_4	$K_2Cr_2O_7$	Used to test for reducing agent
Chlorine	Cl_2	Oxidises bromide to bromine and iodide to iodine

Oxidising agents can be tested with a reducing agent. The common (testing) reducing agents include:

Name of compound	**Formula**	**Applications**
Potassium iodide	KI	Used to test for oxidising agent
Carbon monoxide	CO	Used to reduce iron oxides to iron (and other metal oxides) in the blast furnace
Hydrogen	H_2	Reduces copper (II) oxide to copper
Sulfur dioxide	SO_2	Used as a bleach, and a preservative for food and drinks
Metals (high in the reactivity series)	Na, Mg…	Displacement of less reactive metals

METALS

PHYSICAL PROPERTIES OF METALS

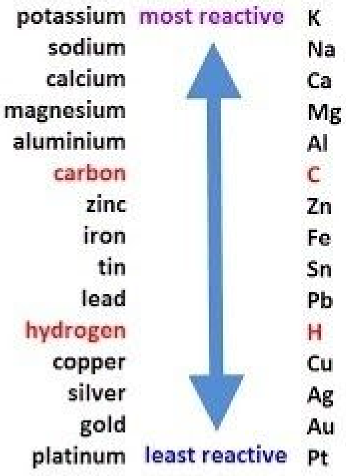
- Ductile
    - Ability to be stretched to form wires
- Malleable
    - Can be bent and beaten in to different shapes
- Good conductors of electricity
    - Presence of delocalised **electrons** can act as mobile charge carriers to conduct electricity
- Shiny
    - Some metals are dull on the surface, shiny underneath
- High melting point/boiling point
- High density
- Tensile strength

Metals can be used to form **alloys**

- Alloy: mixture of metallic elements. Some alloys contain non-metallic elements
- Most metals are soft and weak for most uses
    - Neatly arranged layers of atoms can slide over each other very easily
- In an alloy, atoms have different sizes
    - Disrupts orderly layer of atoms → makes it more difficult for the layers for slide over each other
    - Makes alloys harder and stronger than the pure metal

CHEMICAL PROPERTIES

The **reactivity** of metals can be determined by the reactivity series. This affects the type of reactions that occur.



Some metals react with **liquid water**, forming a metal hydroxide and hydrogen gas. Such metals include potassium, sodium, and calcium.

$2K\ (s) + 2H_2O\ (l) \rightarrow 2KOH\ (aq) + H_2\ (g)$

Potassium reacts very vigorously with water. A lot of heat is produced, causing the hydrogen produced to catch fire and explode.

Some metals react with **gaseous water (water vapour/steam)**, forming a metal oxide and hydrogen gas. Such metals include magnesium, zinc, and iron.

$Mg\ (s) + H_2O\ (g) \rightarrow MgO\ (s) + H_2\ (g)$

Magnesium reacts vigorously with water vapour, forming solid metal oxide and hydrogen gas.

Metals react with **acids**, forming a salt and hydrogen gas. Salt formed depends on the type of acid used.

$2K\ (s) + 2HCl\ (aq) \rightarrow 2KCl\ (s) + H_2\ (g)$

Potassium reacts explosively with hydrochloric acid, forming potassium chloride solid and hydrogen gas.

REACTIVITY SERIES

The reactivity series can be used to predict chemical reactions. This can be found on the previous page.

The reactivity series determines how readily a metal gives up electrons to form ions. The more readily it loses electrons, the more reactive the metal.

In some reactions, the reactivity of metals will affect the **displacement of metals** from solutions. This happens when one metal displaces (takes the place of) the ions of any less reactive metal in the reactivity series.

$Zn (s) + CuSO_4 (aq) \rightarrow Cu (s) + ZnSO_4 (aq)$

In this reaction, zinc displaces copper from its solution, as zinc is more reactive than copper.

$Mg (s) + CuO (s) \rightarrow MgO (s) + Cu (s)$

In this reaction, magnesium displaces copper from its oxide, as magnesium is more reactive than copper..

$Zn (s) + NaCl (aq) \rightarrow$ no reaction

Zinc cannot displace sodium from its solution, as zinc is less reactive than sodium. Hence, there is no reaction.

When a metal oxide reacts with carbon, it forms a pure metal and carbon dioxide. The lower the position of a metal in the reactivity series (less reactive), the easier it is for carbon to remove oxygen from the metal oxide.

$2CuO (s) + C (s) \rightarrow 2Cu (s) + CO_2 (g)$

In this reaction, carbon removes oxygen from copper oxide, forming copper solid and carbon dioxide.

Hydrogen can remove oxygen from metal oxides, producing the metal and water (steam). The lower the position of a metal in the reactivity series, the easier it is for hydrogen to remove oxygen from the metal oxide.

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

In this reaction, hydrogen removes oxygen from lead oxide, forming lead solid and steam.

Metal carbonates decompose when heated strongly, forming its metal oxide and carbon dioxide gas. The lower the position of a metal in the reactivity series, the easier it is for the metal carbonate to decompose when heated.

$CuCO_3 (s) \rightarrow CuO (s) + CO_2 (g)$

RUSTING

Rusting occurs when iron becomes iron oxide. The conditions for rusting to occur is the presence of oxygen and water. Rust formation occurs in 3 steps:

1. $2Fe (s) + O_2 (g) + 2H_2O (l) \rightarrow 2Fe(OH)_2 (s)$
2. $4Fe(OH)_2 (s) + O_2 (g) + 2H_2O (l) \rightarrow 4Fe(OH)_3 (s)$
3. Iron (III) hydroxide loses water to become hydrated iron (III) oxide.

There are ways of preventing rusting from occurring.

Surface protection is when the surface of an object is covered with a layer of substance.

- Layer stops air and water from reaching the iron/steel underneath
- Includes painting, oiling, greasing, plastic coating, and metal plating.
    - Paint, oil, grease, plastic are cheap
    - Metal plating is expensive
- When the layer is broken, air and water can reach the iron/steel, hence rusting occurs

Sacrificial protection is when the more reactive metals are placed next to the iron/steel, such that the metals will corrode in place of the iron/steel rusting.

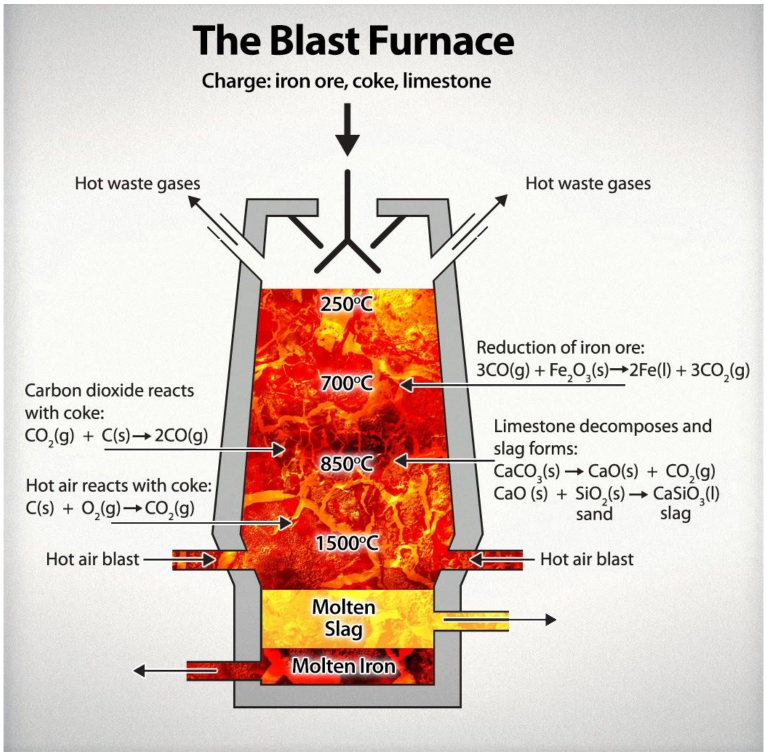
- When iron touches, or is covered by a more reactive metal (Mg, Zn, Al), no rusting occurs
- However, less reactive metals speed up the rusting of iron
- This is used in galvanized iron, protecting ships, and underground steel pipes
    - Galvanised iron is steel coated with zinc metal, It is most often used for roofs of houses in Southeast Asia
    - Steel hulls have blocks of zinc attached to them so they don’t rust as quickly as they should

Stainless steel is an alloy of ion, containing chromium or nickel.

- Does not rust easily
- Expensive way of prevent rusting
- Used in knives and spoons, medical instruments used in hospital operations, kitchen sinks, some steel objects used in chemical factories/oil refineries

EXTRACTION OF IRON

Iron is extracted from its metal ore, which is haematite. Haematite is mainly iron (III) oxide. It is extracted in a blast furnace. Haematite, limestone (calcium carbonate), and coke (carbon) are fed into the blast furnace.



There are 5 main reactions taking place in the blast furnace.

1. Coke reacts with oxygen in the air, forming carbon dioxide and heat.

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

2. Carbon dioxide reacts with more coke, forming carbon monoxide.

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

3. Carbon monoxide reacts with haematite, forming molten iron and carbon dioxide.

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

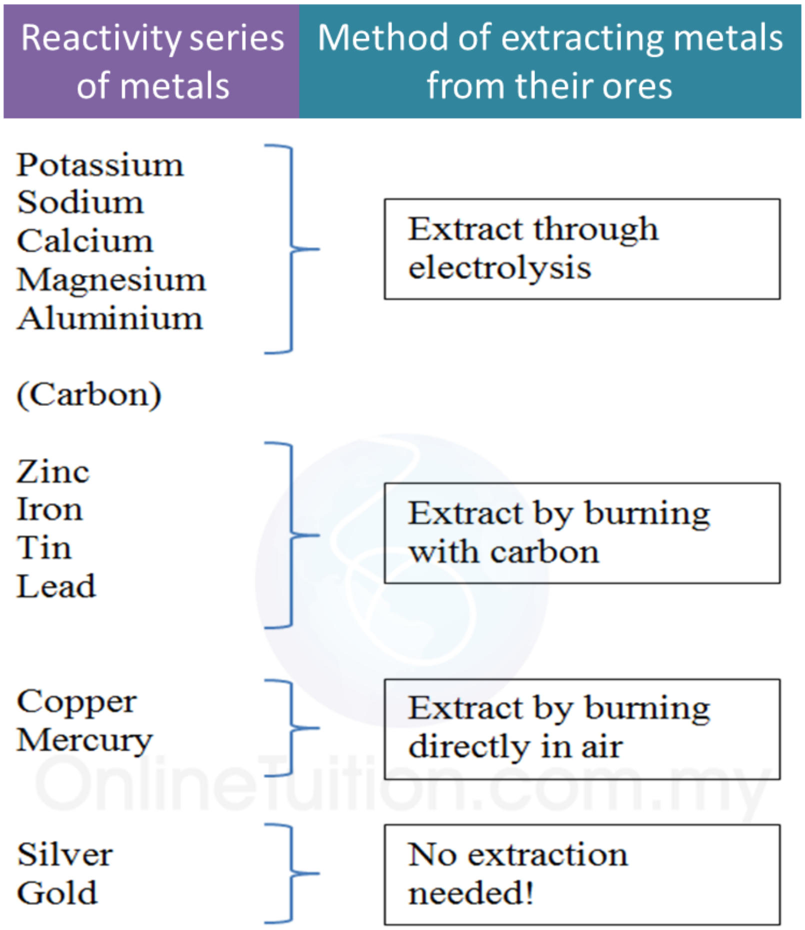
4. Limestone Is decomposed by heat to produce lime and carbon dioxide.

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

5. Impurities are removed from the iron ore. This includes silicon dioxide. Silicon dioxide reacts with calcium oxide, producing slag.

$SiO_2 (s) + CaO (s) \rightarrow CaSiO_3 (l)$

EXTRACTION OF METALS



The less reactive the metal, the easier it is to extract.

TYPES OF STEEL

There are different types of steel for different purposes.

- Mild steel/ Low-carbon steel
    - Strong and malleable
    - 0.25% carbon
    - Used for car bodies, ships, bridges, railway lines, steel rods
- Hard steel/ High-carbon steel

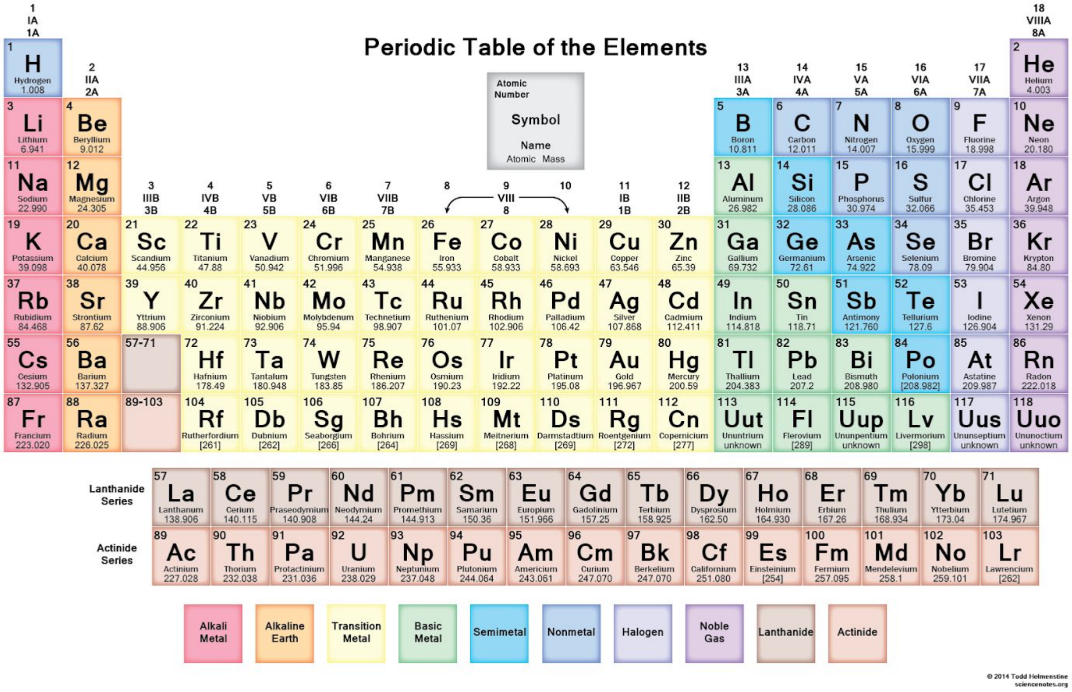
    - Harder than mild steel, less malleable
    - 1% carbon
    - Used to make tools
- Stainless steel
    - Hard, shiny, does not rust
    - Iron with large amounts of chromium and nickel added
    - Used to make cutlery, medical instruments, pipes in chemical industries

RECYCLING

Metal is a finite source. Hence, it is encouraged to recycle metals so that our metal reserves last longer. Recycling metals are also cheaper than extracting metals from its ore.

Benefits	**Problems**
• Conserves metals, especially valuable metals • Saves cost of extracting new metals • Beneficial to the environment. Recycling iron and steel prevents piles of rusty metal from being dumped on wasteland.	• Only recycled if it is economical to do so • Can cause much damage to the environment. Old car batteries are collected and melted down to recycle the lead. However fumes produced pollute the surroundings • Cost of separating metals from waste can be high • Transport costs for collecting scrap metal, difficulty of persuading people to deposit used materials in recycling containers.

THE PERIODIC TABLE



Periodic Table of the Elements

Atomic Number
Symbol
Name
Atomic Mass

1 IA 1A	2 IIA 2A	3 IIIB 3B	4 IVB 4B	5 VB 5B	6 VIB 6B	7 VIIB 7B	8	9 VIII 8	10	11 IB 1B	12 IIB 2B	13 IIIA 3A	14 IVA 4A	15 VA 5A	16 VIA 6A	17 VIIA 7A	18 VIIIA 8A
1 H Hydrogen 1.008																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.066	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.88	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.933	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.39	31 Ga Gallium 69.732	32 Ge Germanium 72.61	33 As Arsenic 74.922	34 Se Selenium 78.09	35 Br Bromine 79.904	36 Kr Krypton 84.80
37 Rb Rubidium 84.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.94	43 Tc Technetium 98.907	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.411	49 In Indium 114.818	50 Sn Tin 118.71	51 Sb Antimony 121.760	52 Te Tellurium 127.6	53 I Iodine 126.904	54 Xe Xenon 131.29
55 Cs Cesium 132.905	56 Ba Barium 137.327	57-71	72 Hf Hafnium 178.49	73 Ta Tantalum 180.948	74 W Tungsten 183.85	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.967	80 Hg Mercury 200.59	81 Tl Thallium 204.383	82 Pb Lead 207.2	83 Bi Bismuth 208.980	84 Po Polonium [208.982]	85 At Astatine 209.987	86 Rn Radon 222.018
87 Fr Francium 223.020	88 Ra Radium 226.025	89-103	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [269]	109 Mt Meitnerium [268]	110 Ds Darmstadtium [269]	111 Rg Roentgenium [272]	112 Cn Copernicium [277]	113 Uut Ununtrium unknown	114 Fl Flerovium [289]	115 Uup Ununpentium unknown	116 Lv Livermorium [298]	117 Uus Ununseptium unknown	118 Uuo Ununoctium unknown

Lanthanide Series	57 La Lanthanum 138.906	58 Ce Cerium 140.115	59 Pr Praseodymium 140.908	60 Nd Neodymium 144.24	61 Pm Promethium 144.913	62 Sm Samarium 150.36	63 Eu Europium 151.966	64 Gd Gadolinium 157.25	65 Tb Terbium 158.925	66 Dy Dysprosium 162.50	67 Ho Holmium 164.930	68 Er Erbium 167.26	69 Tm Thulium 168.934	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.967
Actinide Series	89 Ac Actinium 227.028	90 Th Thorium 232.038	91 Pa Protactinium 231.036	92 U Uranium 238.029	93 Np Neptunium 237.048	94 Pu Plutonium 244.064	95 Am Americium 243.061	96 Cm Curium 247.070	97 Bk Berkelium 247.070	98 Cf Californium 251.080	99 Es Einsteinium [254]	100 Fm Fermium 257.095	101 Md Mendelevium 258.1	102 No Nobelium 259.101	103 Lr Lawrencium [262]

Alkali Metal
Alkaline Earth
Transition Metal
Basic Metal
Semimetal
Nonmetal
Halogen
Noble Gas
Lanthanide
Actinide

Red (Group I) - Alkali metals

Yellow - Transition metals

Indigo (Group VII) - Halogens

Purple (Group VIII/Group 0) - Noble gases

ALKALI METALS

Alkali metals have similar physical properties.

- Shiny, silvery solids
- Soft and easily cut with a knife
- Low densities/low melting points
  - Densities increase down the group

Alkali metals are reactive metals, hence they have similar reactions which change down the group.

- Metals tarnish (react) easily in air, they are kept in oil to prevent contact with the air
- They react with water to give alkaline solution
    - Turns red litmus paper blue
- Metals react vigorously with water, usually catch fire, may cause explosions
    - Products are hydrogen gas and metal hydroxide
- Metals become more reactive down the group
    - E.g. Lithium fizzes, Potassium catches fire and explodes
- React to form ionic compounds, ions have a charge of +1
    - Compounds have similar chemical formulae

HALOGENS

Halogens have similar physical properties.

- Elements consist of small molecules. They are found as diatomic molecules. (Each molecule has 2 atoms)
    - E.g. Cl_2 (g)
- Melting and boiling points increase down the group
- Halogens have low melting and boiling points
- Elements become darker down the group
    - E.g. Chlorine is greenish-yellow, bromine is reddish-brown

Halogens are reactive non-metals. They are similar in ways which change down the group.

- Halogens react with most metals to form compounds called halides
    - Halides have similar chemical formulae
- Halogens become less reactive down the group
- More reactive halogen will displace a less reactive halogen from an aqueous solution of its ions
    - $Cl_2 \text{ (g)} + 2KI \text{ (aq)} \rightarrow 2KCl \text{ (aq)} + I_2 \text{ (aq)}$

NOBLE GASES

Noble gases have similar physical properties.

- They are monoatomic
- Noble gases have stable electronic arrangements
    - Atoms have 8 electrons (other than helium) in their outermost shell. Helium has 2 electrons in its outer shell.
- They have **very low** melting and boiling points

Noble gases are generally unreactive.

TRANSITION METALS

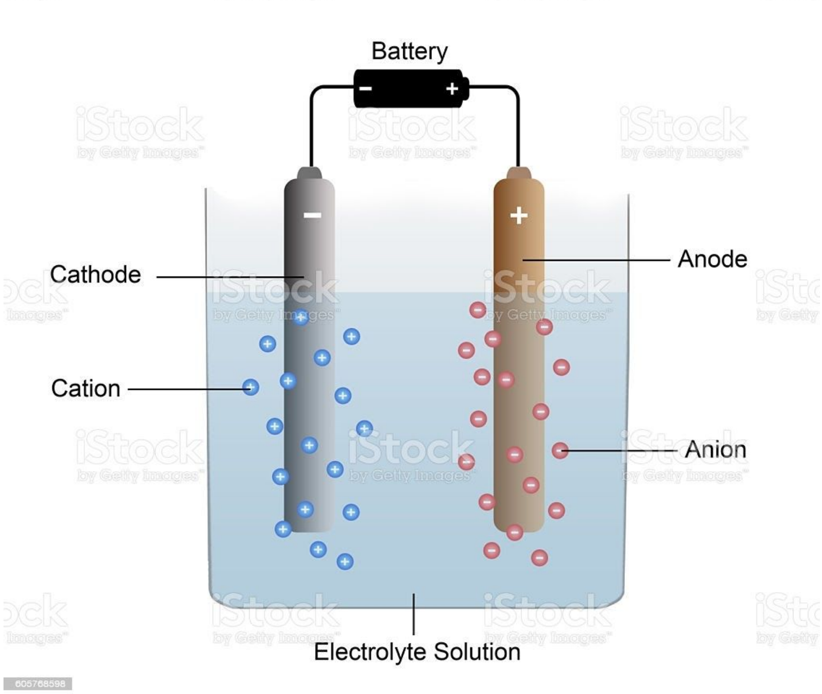
There are a few common physical properties of the transition metals.

- They are shiny and silvery in appearance (when clean)
- Good conductors of electricity
- Hard and strong
    - Stronger than those in group I and group II metals
- Have high densities and high melting and boiling points
- Compounds of most transition metals are coloured
    - Useful in identifying compounds

Transition metals (or transition elements) also have some similar chemical properties.

- They have variable valency and oxidation states in compounds
    - E.g. Iron forms compounds containing iron (II) ions (Fe^{2+}) or iron (III) ions (Fe^{3+})
    - Metals that are not transition metals usually form only one ion (E.g. Na^{+} ion for sodium)
- They are good catalysts to speed up chemical reactions

ELECTROLYSIS



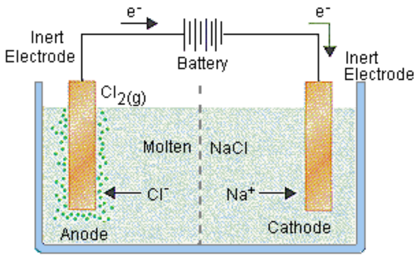
Cations are positive ions. **Anions** are negative ions.

During electrolysis, reactions occur only at the electrodes.

- Oxidation occurs at the anode (AN$_{\text{(ode)}}$ OX$_{\text{(idaton)}}$)
- Reduction occurs at the cathode (RED$_{\text{(uction)}}$ CAT$_{\text{(hode)}}$)

ELECTROLYSIS OF MOLTEN IONIC COMPOUNDS

Molten sodium chloride



1. Sodium chloride is an ionic compound, consisting of Na^+ ions and Cl^- ions
2. **Positively charged ions are attracted to the cathode, negatively charged ions are attracted to the anode**
3. At the cathode (negatively charged), sodium ions take electrons from the electrode to become sodium atoms
    3.1. $Na^+ (l) + e^- \rightarrow Na (l)$
    3.2. Reduction takes place
    3.3. Sodium ions are **discharged**
4. At the anode (positively charged), chlorine ions give up electrons to the electrode to become chlorine atoms
    4.1. $Cl^- (l) \rightarrow Cl (l) + e^+$
    4.2. Oxidation takes place
    4.3. Chlorine atoms combine to form chlorine molecules (chlorine gas)
    4.4. $2Cl (l) \rightarrow Cl_2 (g)$
    4.5. Chlorine gas is discharged around the anode during the electrolysis

$2NaCl (l) \rightarrow 2Na (l) + Cl_2 (g)$

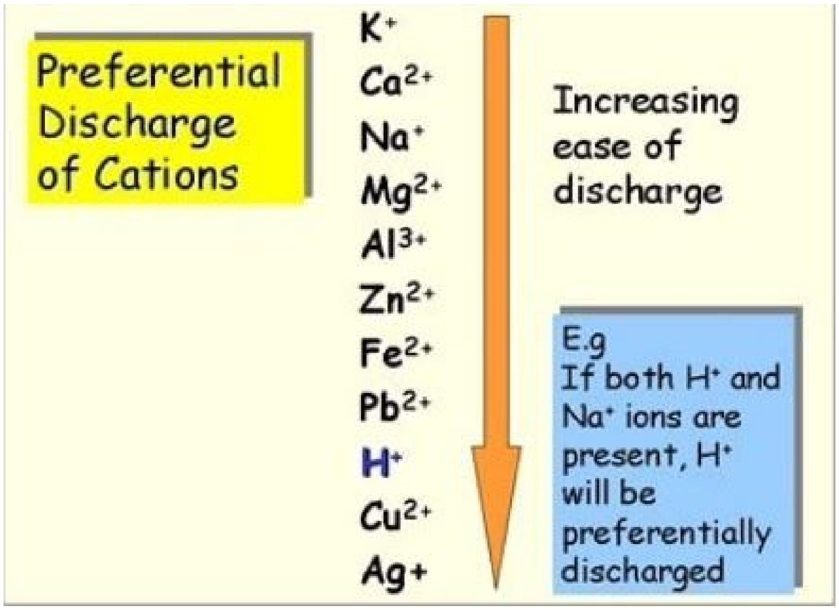
The carbon electrodes are called **inert electrodes** since they do not react in the electrolysis.

Note: Molten means liquid state. Aqueous means there is water added. Molten ≠ Aqueous

SELECTIVE DISCHARGE OF IONS

At the **cathode**, cation discharged is related to the position of metals in the reactivity series.

- Ions of metals with lower reactivity are preferentially discharged over H^+ ions
- If cations are of ions of highly reactive metals, they are not discharged in the presence of water. Instead, H^+ ions from the water are discharged to produce hydrogen gas.
    - $2H^+ (aq) + 2e^- \rightarrow H_2 (g)$

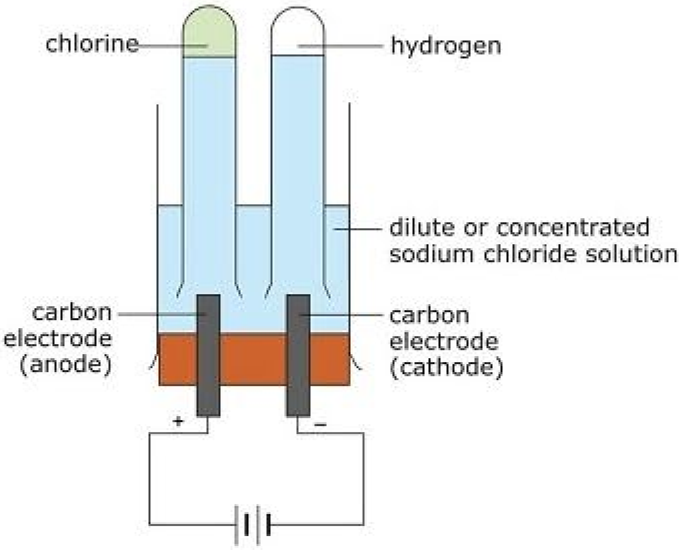


At the **anode**, halide ions (chloride, bromide, iodide) are only discharged in electrolysis of concentrated solutions.

- In diluted solutions, **hydroxide ions** are discharged instead
    - In electrolysis of a concentrated solution of sodium chloride, chloride ions are discharged, giving chlorine gas
    - $4OH^- (aq) \rightarrow O_2 (g) + H_2O (l) + 4e^-$
- Hydroxide ions (OH^-) are discharged preferentially over sulfate ions (SO_4^{2-}) and nitrate ions (NO_3^-)

ELECTROLYSIS OF AQUEOUS SOLUTIONS OF IONIC COMPOUNDS

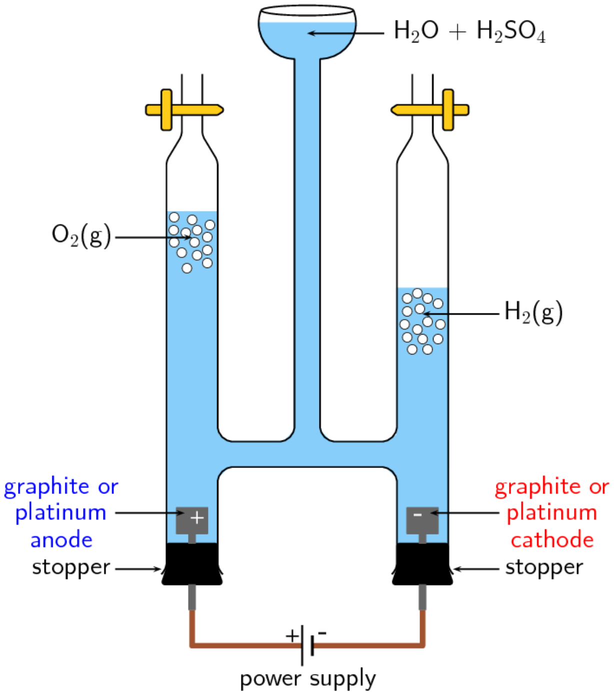
Concentrated sodium chloride solution



1. Sodium chloride is an ionic compound, consisting of Na^+ ions and Cl^- ions
2. At the cathode (negatively charged), both hydrogen and sodium are attracted
    2.1. H^+ ions are discharged instead of the Na^+ ions, giving hydrogen gas
    2.2. $2H^+ (aq) + 2e^- \rightarrow H_2 (g)$
    2.3. Reduction takes place
3. At the anode (positively charged), both hydroxide and chlorine ions are attracted
    3.1. Chlorine ions give up electrons to the electrode to become chlorine atoms
    3.2. $Cl^- (l) \rightarrow Cl (l) + e^+$
    3.3. Oxidation takes place
    3.4. Chlorine atoms combine to form chlorine molecules (chlorine gas)
    3.5. $2Cl (l) \rightarrow Cl_2 (g)$
    3.6. Chlorine gas is discharged around the anode during the electrolysis

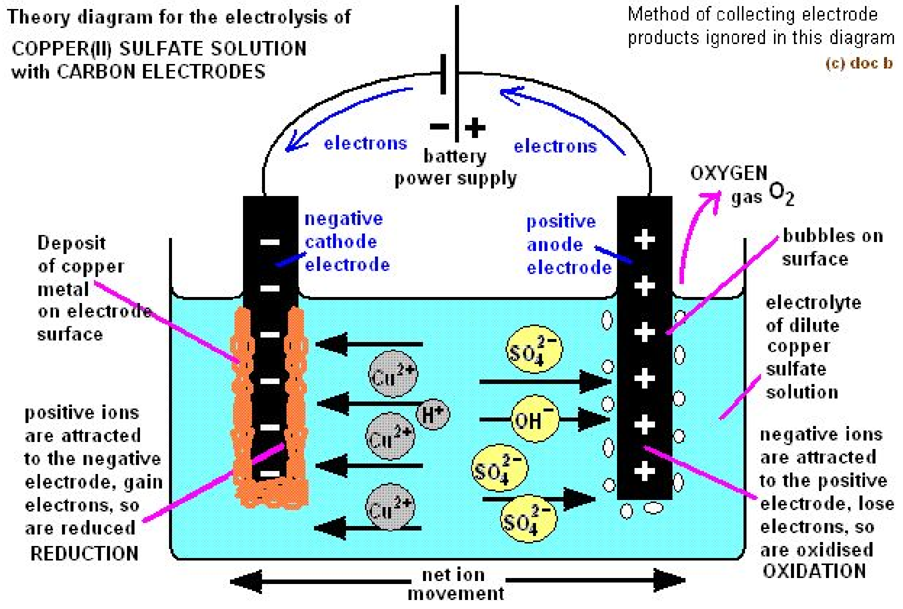
As OH^- ions and Na^+ ions remain in the solution, sodium hydroxide solution is formed.

Dilute sulfuric acid



1. Dilute sulfuric acid contains H^+ ions (from both sulfuric acid and water), SO_4^{2-} ions (from sulfuric acid), and OH^- ions (from water)
2. At the cathode (negatively charged), both hydrogen and sodium are attracted
    - 2.1. H^+ ions are discharged giving hydrogen gas
    - 2.2. $2H^+ (aq) + 2e^- \rightarrow H_2 (g)$
    - 2.3. Reduction takes place
3. At the anode (positively charged), both hydroxide and sulfate ions are attracted
    - 3.1. Hydroxide are discharged to become oxygen gas
    - 3.2. $4OH^- (aq) \rightarrow O_2 (g) + H_2O (l) + 4e^-$
    - 3.3. Oxidation takes place
    - 3.4. Oxygen gas discharged around the anode during the electrolysis

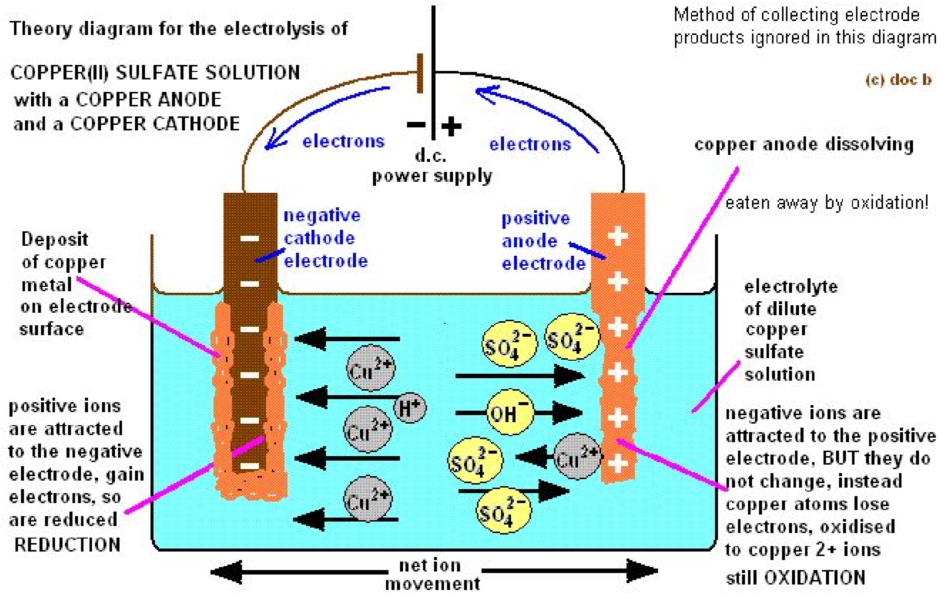
Copper (II) sulfate solution using carbon electrodes



1. Copper (II) sulfate solution contains Cu^{2+} ions and SO_4^{2-} ions (from copper (II) sulfate), H^+ ions and OH^- ions (from water)
2. At the cathode (negatively charged), both hydrogen and copper ions are attracted
    - 2.1. Cu^{2+} ions are discharged
    - 2.2. Cu^{2+} (aq) + $2e^-$ → Cu (s)
    - 2.3. Reduction takes place
    - 2.4. Copper metal is deposited on the cathode
3. At the anode (positively charged), both hydroxide and sulfate ions are attracted
    - 3.1. Hydroxide are discharged to become oxygen gas
    - 3.2. $4OH^-$ (aq) → O_2 (g) + H_2O (l) + $4e^-$
    - 3.3. Oxidation takes place
    - 3.4. Oxygen gas discharged around the anode during the electrolysis

The blue colour of the solution slowly fades as the concentration of copper (II) ions decreases. As hydrogen and sulfate ions are not discharged, the solution becomes sulfuric acid.

Copper (II) sulfate solution using copper electrodes



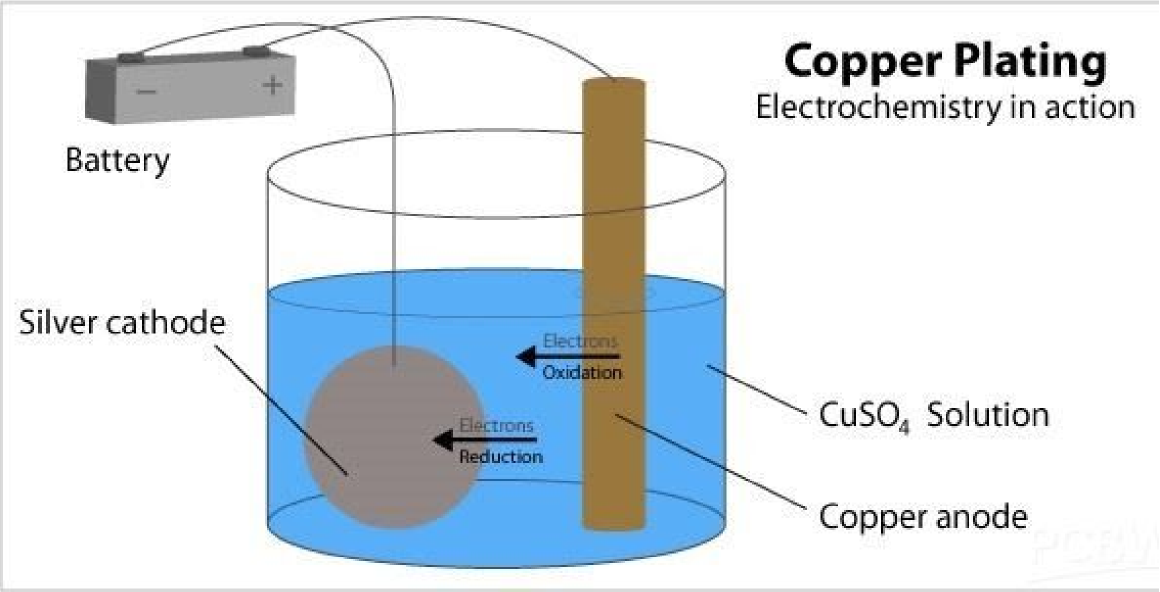
1. Copper (II) sulfate solution contains Cu^{2+} ions and SO_4^{2-} ions (from copper (II) sulfate), H^+ ions and OH^- ions (from water)
2. At the cathode (negatively charged), both hydrogen and copper ions are attracted
    - 2.1. Cu^{2+} ions are discharged
    - 2.2. $Cu^{2+} (aq) + 2e^- \rightarrow Cu (s)$
    - 2.3. Reduction takes place
    - 2.4. Copper metal is deposited on the cathode
3. At the anode (positively charged), neither the sulfate ions nor the hydroxide ions are discharged.
    - 3.1. Copper in the anode goes into the solution to from copper (II) ions
    - 3.2. $Cu (s) \rightarrow Cu^{2+} (aq) + 2e^-$

The copper cathode gains copper and becomes larger. The copper anode dissolves and becomes smaller. For each copper (II) ion discharged at the cathode, one copper (II) ion goes into the solution at the anode.

The concentration of copper (II) sulfate in the solution and the colour of the solution does not change.

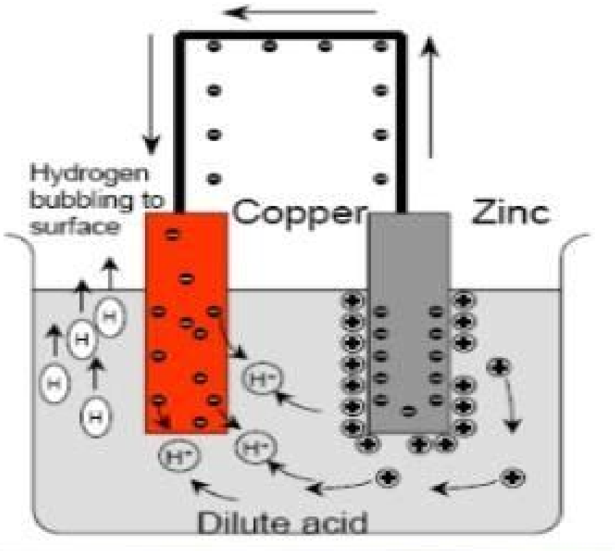
APPLICATIONS OF ELECTROLYSIS

- Electrolysis is used to purify metals. This can be seen during the electrolysis of copper (II) sulfate solution using copper electrodes.
    - The copper produced in this process is 99% pure
- **Electroplating**: Coating of an object with a thin layer of a metal by electrolysis
    - Electrolysis can electroplate an object



- Steel objects are also electroplated, where the metals used to electroplate steel objects are corrosion resistant and shiny
- A similar setup to an electrolytic cell can also be used instead of a sacrificial metal to protect steel against rusting

ELECTRIC CELLS



The metal which is higher up in the reactivity series gives up electrons more readily than the negative electrode.

1. At the zinc electrode, zinc atoms give up electrons to form zinc ions
    1.1. $Zn (s) \rightarrow Zn^{2+} (aq) + 2e^-$
    1.2. Zinc ions go into the solution
2. At the copper electrode, the solution contains positive sodium and hydrogen ions
    2.1. Hydrogen ions accept the electrons to form hydrogen gas
    2.2. $2H^+ (aq) + 2e^- \rightarrow H_2 (g)$
3. The metal lower down in the reactivity series is the positive electrode
    3.1. Copper metal
    3.2. When voltmeter is attached,voltage of about 1V is obtained

<u>The further apart the metals are in the reactivity series, the larger the voltage.</u>