

Chemistry

Paper 1+2 Knowledge

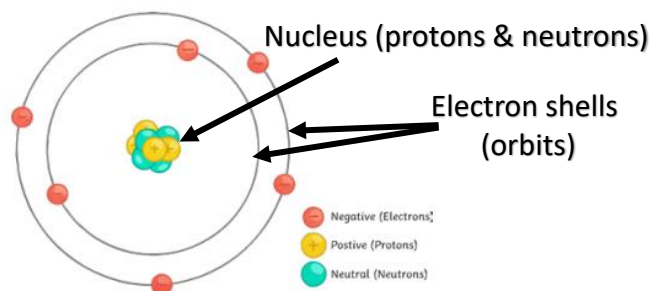
Organisers

AQA Triple Chemistry

C1 – Atomic Structure and The Periodic Table

Atoms

- Made up of **protons, electrons** and **neutrons**.



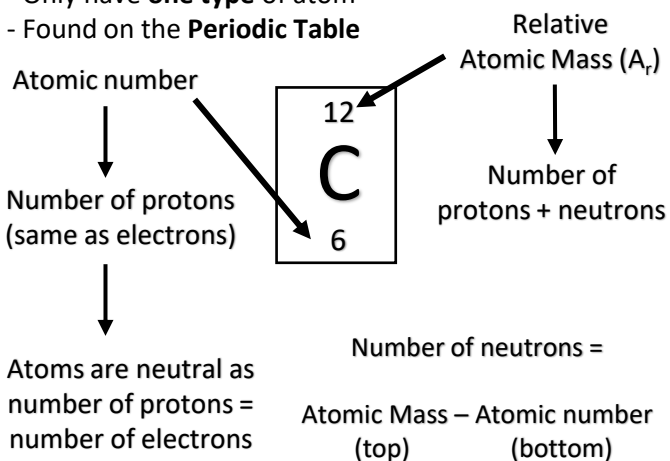
Subatomic particle	Relative Mass	Charge
Proton	1	Positive
Neutron	1	Neutral
Electron	Very small	Negative

Atoms have a radius of about 0.1nm (1×10^{-10} m)

Radius of nucleus = about 1×10^{-14} m

Elements

- Only have **one type** of atom
- Found on the **Periodic Table**

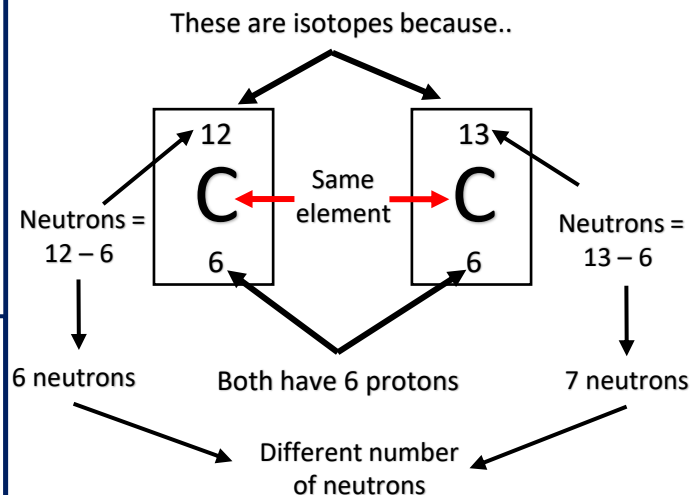


Compounds

- Two or more elements **chemically combined**.
- Formed by chemical reactions
- For example: CO_2 H_2O CH_4 HCl NaCl

Isotopes

Isotope = atoms of the **same element** which have the **same number of protons**, but a **different number of neutrons**.



Chemical Equations

- Shown by using a **word equation**.
e.g. magnesium + oxygen $\rightarrow$ magnesium oxide

Left of the arrow = **reactants**
Right of the arrow = **products**.

- Also can be shown by a **symbol equation**
e.g. $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

Mixtures and Separation

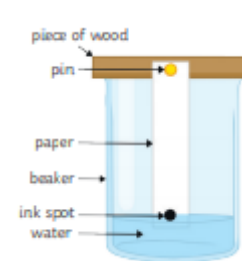
Mixtures – two or more elements or compounds **not** chemically joined.

This means the different components of the mixture can be separated by physical methods (below)

E.g. air is a mixture mainly made of nitrogen, oxygen and carbon dioxide.

Chromatography

to separate out mixtures (usually liquids) (e.g. colours in ink)



Filtration

To separate insoluble solids from liquids (e.g. sand and water)



Evaporation

To quickly separate soluble solids from a solution. (e.g. salt and water)



Crystallisation

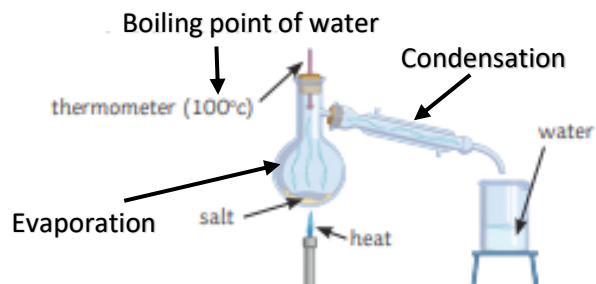
To slowly separate a soluble salt from a solution. (e.g. copper sulfate crystals)



C1 – Atomic Structure and The Periodic Table

Distillation

Simple distillation – separating a liquid from a solution.



- Liquid is heated to boiling point and evaporates
- Vapours travel up into the condenser
- Condenser has cold water around it.
- Vapours cool and condense (turn back into a liquid).

Electronic Structure

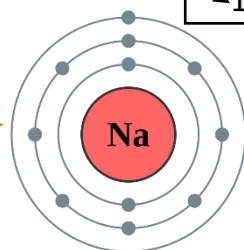
- Electrons are found on shells (orbits) orbiting the nucleus.
- There is a maximum number of electrons allowed on each shell:

First shell = 2 electrons
Second shell = 8 electrons
Third shell = 8 electrons.

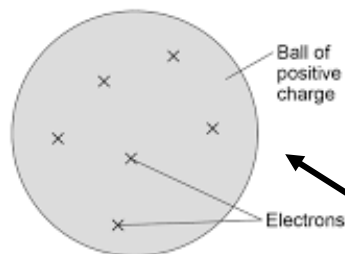
11 electrons

1st shell = 2
2nd shell = 8
3rd shell = 1

Total = 11 electrons



Plum pudding model



Differences to nuclear model

- Ball of positive charge (no protons)
- No nucleus
- No neutrons
- Evenly distributed mass

Rutherford tested the plum pudding model

History of the atom

Scientist	Time	Discovery
John Dalton	Start of the 19 th century	Atoms were first described as solid spheres.
JJ Thomson	1897	Plum pudding model – atom is a ball of + charge with electrons scattered
Ernest Rutherford	1909	Alpha scattering experiment - mass concentrated at the centre, only the nucleus is + charged. Most of the atoms is empty space.
Niels Bohr	Around 1911	Electrons are in shells orbiting the nucleus
James Chadwick	Around 1940	Discovered that there are neutrons in the nucleus.

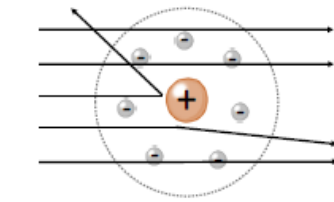
Rutherford's scattering experiment

alpha particles are positively charged

Fired at gold foil

some alpha particles are deflected/ repelled

most alpha particles passed straight through



What happened?

Conclusions made

Observation	Conclusion
Most of the particles passed straight through	Most of the atom is empty space
Some were deflected to the sides	The particles had passed close by a positive charge
A very small number were repelled straight back	The alpha particles had approached the nucleus straight on. the tiny number told him that the positive charge is in a very small dense core

- Elements **are less** reactive because:
- More electron shells
- Outer shell is further from nucleus and is **more shielded** by the other shells
- The electrostatic force of attraction between free electron and nucleus is **weaker**
- Harder to attract an electron into the outer shell.

Compared to group 1, the transition elements:

- Are harder and stronger
- Have higher melting points (except for mercury) and higher densities
- Much less reactive and don't react as vigorously with oxygen or water

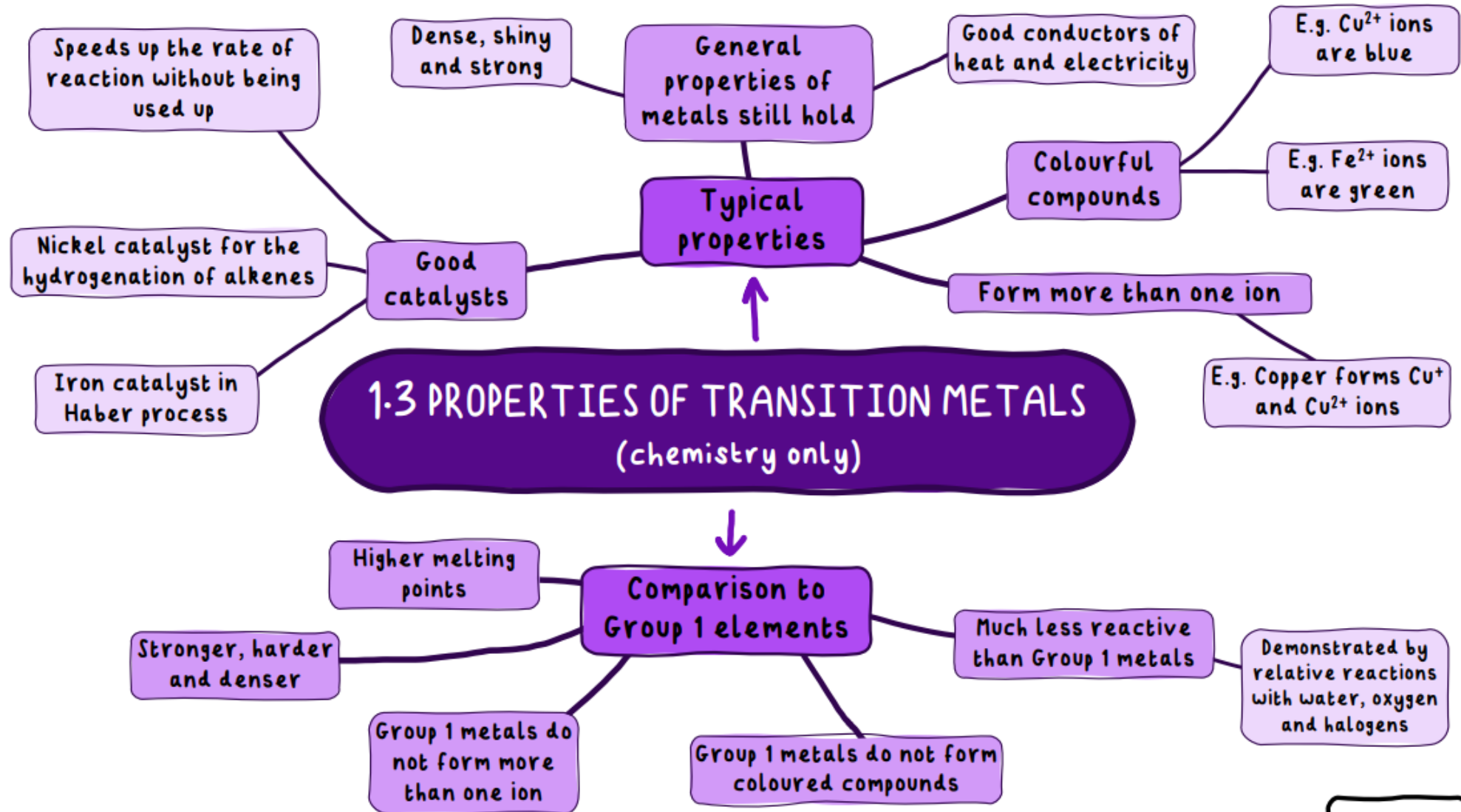
Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg

element:	chromium	manganese	iron	cobalt	nickel	copper
properties: to be compared to group 1)	lustrous, brittle, hard metal	hard and very brittle, difficult to fuse, but easy to oxidise	good conductor, rusts easily in air, strong, ductile malleable	brittle, hard, high melting point	hard, malleable, and ductile, fairly good conductor of heat and electricity	highly ductile and conductive. malleable and soft

Typical properties

- They have ions with many different charges
- Form coloured compounds
- Are useful as catalysts.

element:	chromium	manganese	iron	cobalt	nickel	copper
ion charges:	+2 +3 +4 +5 +6	+2 +3 +4 +5 +6 +7	+2 +3 +4 +5 +6	+2 +3 +4 +5	+2 +3 +4	+1 +2 +3
colours of compounds with transition metals having these charges:	+2 +3 +6	+2 +4 +6 +7	+2 +3	+2 +3	+2	+2
uses as a catalyst:		decomposition of hydrogen peroxide ($2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$)	for haber process ($\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$)		manufacture of margarine (adding H_2 to double bonds)	



C2 – Bonding, structure, and the properties of matter

Formation of Ions

- **Ions** = a charged particle made when atoms lose or gain electrons
- **Positive ion** = atom has lost electrons
- **Negative ion** = atom has gained electrons.

Metals form **positive ions**

Non-metals form negative ions

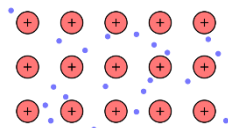
Group	Ions	Example
1	+1	$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$
2	+2	$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$
6	-2	$\text{O} + 2\text{e}^- \rightarrow \text{O}^{2-}$
7	-1	$\text{Br} + \text{e}^- \rightarrow \text{Br}^-$

Lost electrons

Gained electrons

Metallic Bonding

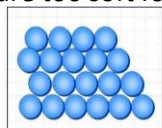
- Happens in **metals only**.
- Positive metal ions surrounded by **sea of delocalised electrons (can move)**.
- Ions tightly packed in rows.
- Strong **electrostatic forces of attraction** between positive ions and negative electrons.



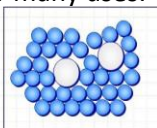
Alloys

- **Alloys** = mixture of two or more metal atoms
- Pure metals are too soft for many uses.

Pure Metal



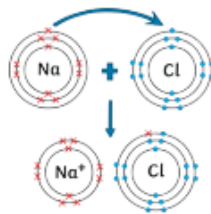
Alloy



- | | |
|-------------------|-------------------------|
| • Atoms same size | • Different sized atoms |
| • Layers slide | • Layers cannot slide |
| • Softer | • Stronger |

Ionic Bonding

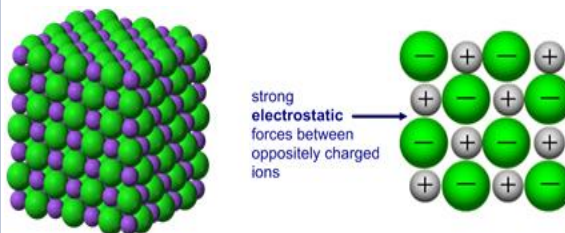
- Between a metal and non-metal.
- Metals give electrons to non-metals so both have a full outer shell.
- **Electrostatic force of attraction** between positive and negative ions.



E.g. Sodium loses one electron to become Na^+ . Chlorine gains one electron to become Cl^- . The two ions attract to form sodium chloride.

Ionic compounds

- Form **giant lattices**, as the attraction between ions acts in all directions



Properties of Ionic Compounds

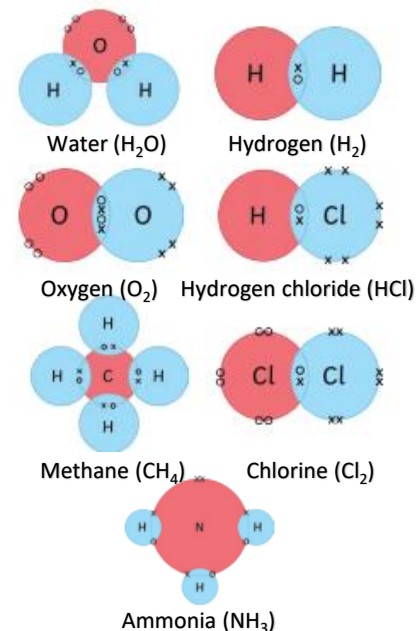
- **High melting point** – lots of energy needed to overcome electrostatic forces.
- **High boiling point**
- **Cannot conduct electricity as solid** – ions cannot move
- **Conducts electricity when molten or dissolved** – ions are free to move.

Covalent Bonding

- **Covalent bonding** = sharing a pair or pairs of electrons for a full outer shell.
- Between **non-metals only**.

Dot and cross diagrams

- Show the bonding in simple molecules.
- Uses the outer shell of the atoms
- Crosses and dots used to show electrons
- You should be able to draw the following:



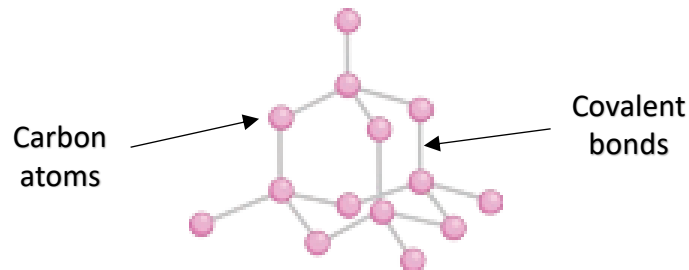
Simple Covalent Molecules

- Form when all atoms have full outer shells so bonding stops
- Examples are the molecules shown above.
- Have **low melting and boiling points**
- Due to **weak intermolecular forces**
- Do not conduct electricity

C2 – Bonding, structure, and the properties of matter

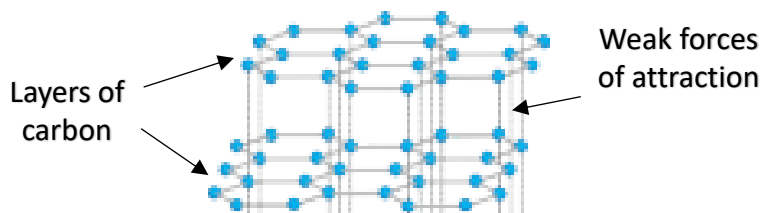
Giant Covalent Structure – Diamond

- Each carbon atom **covalently** bonded to **four** others.
- Forms a giant structure
- This makes diamond **strong** → a lot of **energy** needed to break lots of strong covalent bonds.
- **Does not conduct electricity** – has no free electrons.



Giant Covalent Structure – Graphite

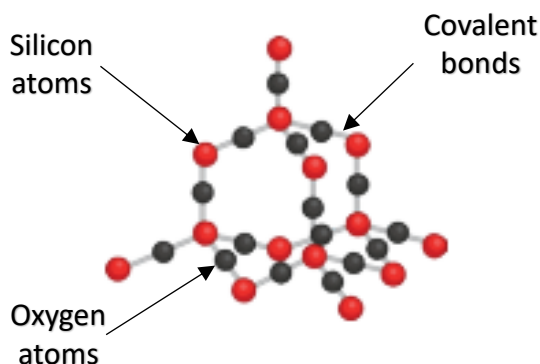
- Layers of **carbon** arranged in **hexagons**.
- Each carbon bonded to **three** other carbons.
- Leaves **one delocalised electron** → moves to carry electrical charge **throughout structure**.



- Layers held together by **weak forces**
- Layers can **slide** over each other easily
- Makes graphite **soft/slippery** → good lubricant.
- Has **high melting point** as has many strong covalent bonds.

Silicon Dioxide

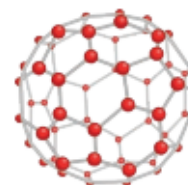
- Similar structure to diamond
- Giant covalent structure.
- Lots of **strong covalent bonds**.
- These require lots of **energy** to break.
- High melting and boiling points.



Fullerenes and Nanotubes

- Molecules of carbon shaped into hollow tubes or balls.
- Used to **deliver drugs into body**

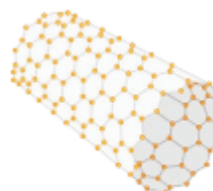
Buckminsterfullerene
Formula = C₆₀



- **Carbon nanotubes** = long narrow tubes
- Can conduct electricity

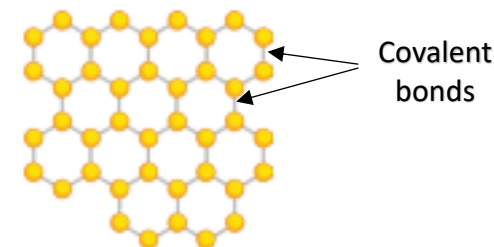
- Can strengthen materials without adding weight.

- Used in electronics and nanotechnology.



Graphene

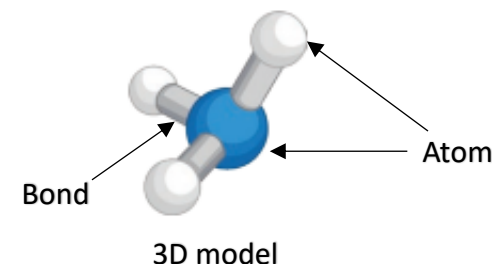
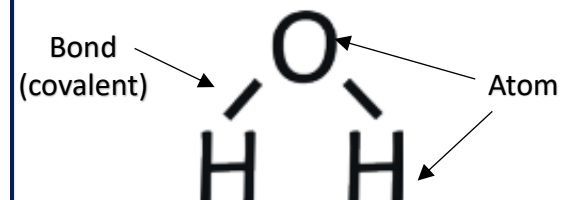
- Graphene = one layer of graphite.
- Very strong → lots of strong covalent bonds.



- Each carbon bonded to three others.
- One **free delocalised electron** → can move to **carry electrical current** throughout the structure.

Molecular models

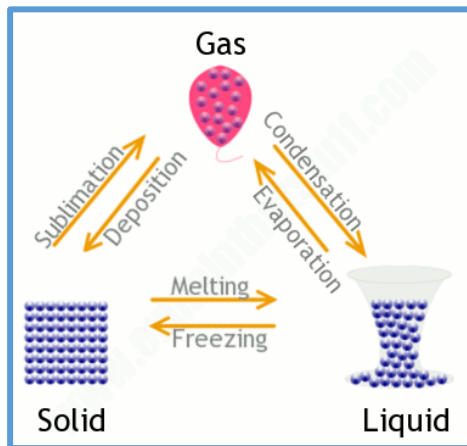
- There are different ways to show a molecule other than dot and cross diagrams.



C2 – Bonding, structure, and the properties of matter

States of Matter

- Three states of matter: **solid, liquid & gas.**
- To change state, **energy** must be **transferred.**



- When heated, particles **gain energy.**
- **Attractive forces** between particles begin breaking when melting or boiling points are reached
- **Amount of energy** needed to change state depends on how strong forces are.

Gas

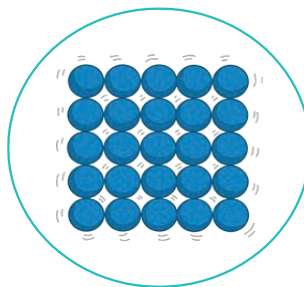
- Randomly arranged.
- Particles **move quickly** – all directions.
- Highest **amount of kinetic energy.**



- Gases **are able to flow** – fill containers
- **Can be compressed** as there is **space between particles**

Solid

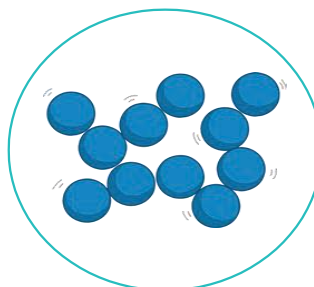
- **Regular** pattern (rows and columns)
- Particles **vibrate** in a **fixed position.**
- Particles have **low amount of kinetic energy.**



- Have a **fixed shape** – cannot flow because of strong forces of attraction between particles
- **Cannot be compressed** – particles close together.

Liquid

- Particles **randomly** arranged and touching.
- Particles can **move around.**
- **Greater amount of kinetic energy** than solid



- Liquids **able to flow** – take shape of containers.
- **Cannot be compressed** – particles are close together and cannot be pushed closer

State symbols

- States of matter shown in chemical equations:
- Solid (**s**)
- Liquid (**l**)
- Gas (**g**)
- Aqueous (**aq**)
- **Aqueous** solutions = substance dissolved in water.

Identifying Physical State of Substances

- If the temperature is **lower** than a substance's melting point – substance is **solid.**
- If the temperature is **between** the melting point and boiling point – substance is **liquid.**
- If the temperature is **higher** than the boiling point – substance is a **gas.**

Limitations of Particle Model (HT)

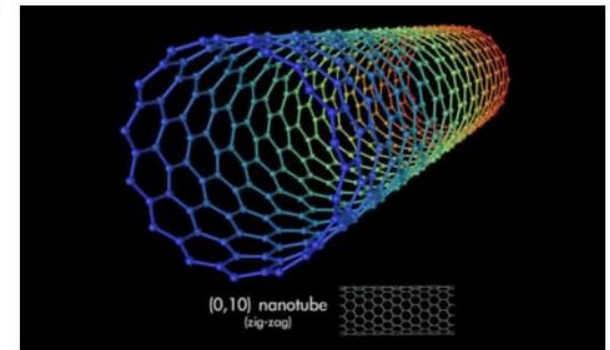
- No chemical bonds are shown.
- Particles shown as solid spheres – not the case, particles are mostly empty space like atoms.
- The diagrams don't show any of the forces between particles
- The diagrams are unable to show the movement of the particles.

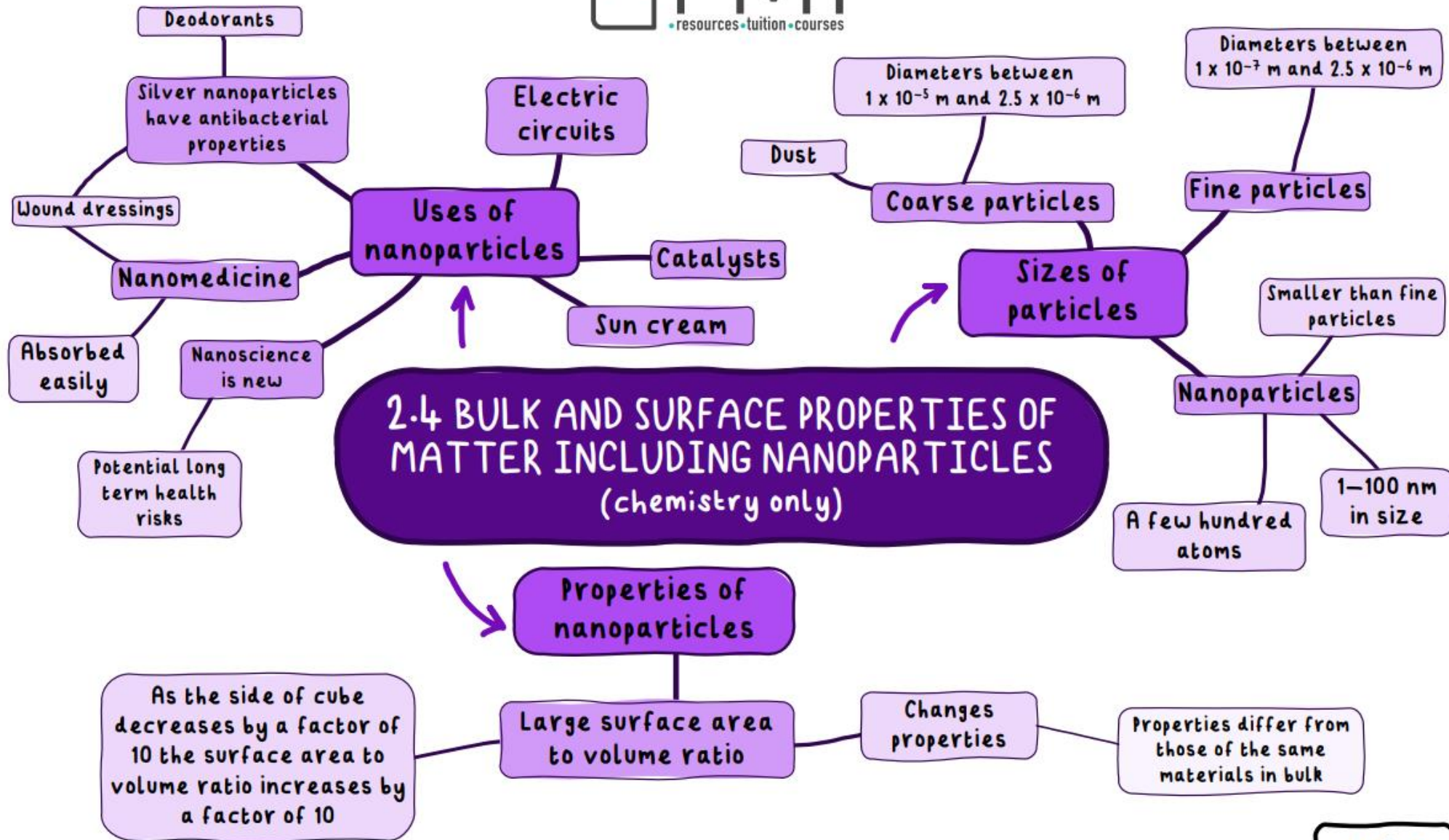
Nanoparticles

- Nanoparticles are 1-100 nanometers across.
- They contain a few hundred atoms.
- Nanoparticles, are smaller than fine particles (PM2.5), which have diameters between 100 and 2500 nm (1×10^{-7} m and 2.5×10^{-6} m).
 - Coarse particles (PM10) have diameters between 1×10^{-5} m and 2.5×10^{-6} m.
 - Coarse particles are often referred to as dust.
- As the side of cube decreases by a factor of 10 the surface area to volume ratio increases by a factor of 10
- Nanoparticles involve fullerenes.
- A nanoparticle has different properties to the 'bulk' chemical it's made from, because of their high surface area to volume ratio. It may also mean that smaller quantities are needed to be effective than for materials with normal particle sizes. e.g fullerenes have different properties to big lumps of carbon.

Their uses

- They have a high surface area to volume ratio, and therefore would make good catalysts.
- They can also be used to produce highly selective sensors.
- Nanotubes could make stronger, lighter building materials.
- New cosmetics, e.g sun tan cream and deodorant. They make no white marks.
- Lubricant coatings, as they reduce friction. These can be used for artificial joints and gears.
- Nanotubes conduct electricity, so can be used in small electrical circuits for computers.
- possible disadvantages: there are some concerns that nanoparticles may be toxic to people. They may be able to enter the brain from the bloodstream and cause harm.





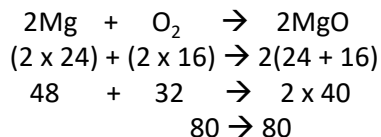
C3 – Quantitative Chemistry

Conservation of Mass

- Atoms cannot be created or destroyed during reactions.
- **Mass of reactants = mass of products.**

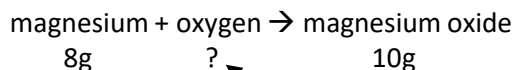
To show mass is conserved in a reaction:

M_r on the left-side must be same as the right side.



Reacting masses

Use conservation of mass to predict masses:



Both sides need to be equal:
 $10\text{g} - 8\text{g} = 4\text{g}$ of oxygen

Percentage Mass

- Percentage mass of an element in a compound

$$\frac{\text{Mass of the element in compound}}{\text{Total mass of compound}} \times 100$$

Example Question:

Find the percentage mass of oxygen in magnesium oxide (MgO).

A_r of magnesium = 24 A_r of oxygen = 16

M_r of MgO = $24 + 16 = 40$

$$\% \text{ mass} = \frac{A_r}{M_r} = \frac{16}{40} = 0.4 \times 100 = 40\%$$

X 100 to make a % 40% of the mass of MgO is oxygen

Mass Changes

- Mass is always conserved in a reaction.
- Sometimes it may seem like the mass has increased/decreased.
- If a **reactant** is a gas – mass may **increase**.

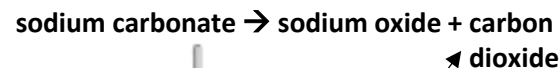


Oxygen is in the air before it combines with magnesium – you cannot find the mass of oxygen on the balance.

It will look like the mass has increased when it is re-weighed at the end.



- If a **product** is a gas and the gas is able to escape the system – mass will **decrease**.

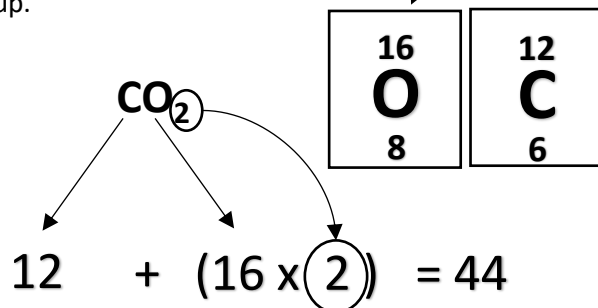


It will look like the mass has decreased as some of the atoms have been given off as gas and have escaped – so cannot be re-weighed.

Atomic mass (A_r) and Relative Formula Mass (M_r)

- Atomic mass (A_r) is the mass number – ie the mass of one atom
- Relative formula mass (M_r) = all the **relative atomic masses (A_r)** of the atoms in a compound or molecule added up.

Example



The Mole (HT only)

- Avogadro constant – 6.02×10^{23}
- One mole contains 6.02×10^{23} atoms or molecules
- The mass, in g, of one mole is the A_r (if an element) or M_r if a compound or molecular element

Iron has a A_r of 56, so 1 mole of iron is 56 g and contains 6.02×10^{23} atoms of iron

Ammonia (NH_3) has an M_r of 17, so 1 mole of ammonia has a mass of 17g. and contains 6.02×10^{23} molecules of ammonia

C3 – Quantitative Chemistry

Concentrations of Solutions

- Concentration = mass of dissolved substance in specific volume (eg dm³)
- More substance dissolved = more concentrated solution

$$\text{Concentration} = \frac{\text{mass}}{\text{volume}}$$

(g/dm³) (g) (dm³)

Can be rearranged to find mass dissolved:

$$\text{mass} = \text{concentration} \times \text{volume}$$

(g) (g/dm³) (dm³)

$$1000\text{cm}^3 = 1\text{dm}^3$$

$$\text{cm}^3 \rightarrow \text{dm}^3 = \text{divide by } 1000.$$

Calculating mass in a given volume

If you have a known volume of a solution of known concentration then you can calculate the mass of dissolved solid.

E.g Calculate the mass of dissolved solid in 25cm³ of a 96g/dm³ solution

96g/dm³ means 96g in every 1000cm³

Do the same to the other side (÷40)

↓
2.4g

25cm³

How do we get from 1000 to 25? (÷40)

Moles and Equations (HT only)

- You can use moles to help you write balanced symbol equations.

Example Question

18.4g of Sodium reacted with 6.4g of oxygen to give 24.8g sodium oxide. Use the masses to write the balanced equation.

Step	Example
Write the equation for the reaction (unbalanced)	Na + O ₂ → Na ₂ O
write down the mass or % given in the question	18.4 + 6.4 → 24.8
Write the mass of one mole of each element or compound	23 32 62 (e.g 18.4 ÷ 23)
Divide the mass given in question by the mass of one mole	0.8 0.2 0.4
Turn the answers into whole number simple ratio	8 2 4 (cancel down) 4 1 2
Put the numbers into the equation	4Na + O ₂ → 2Na ₂ O

Calculating reacting masses (HT)

Example Question

Calculate the mass of calcium needed to make 11.2g Calcium oxide

Step	Calculation
Write the balanced equation	2Ca + O ₂ → 2CaO
Write the masses of each substance	80 + 32 → 112
Write down the given mass in the question.	11.2
Work out the 'scale' factor (ie what did you have to do to the original number to get to the desired mass	÷ 10
Do the same to the other side	8g

Limiting Reactants (HT only)

- If one reactant runs out before the other, then the reaction will stop.
- The reactant that runs out first in a reaction is known as the limiting reactant.

C3 – Quantitative Chemistry

1. What does concentration mean?
2. How can you make a solution more concentrated?
3. State the equation to calculate concentration in g/dm^3 .
4. What is the unit for volume?
5. How many cm^3 are in a dm^3 ?

Calculating mass in a given volume

1. What does 36.5g/dm^3 mean?
2. Calculate the mass of dissolved solid in 25 cm^3 of a 36.5g/dm^3 solution

36.5

1000

25 cm^3

Do the same to the other side
($\div 40$)

g

How do we get from 1000 to 25?
($\div 40$)

Moles and Equations (HT only)

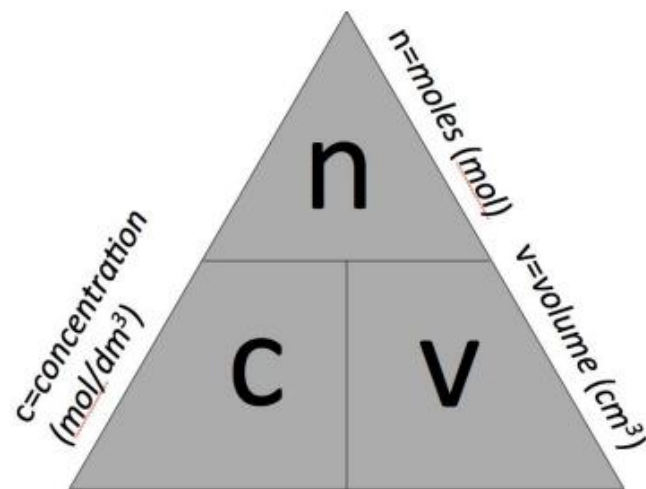
12g of magnesium (Mg) reacted with 8g of oxygen (O_2) to produce 20g magnesium oxide (MgO). Use the masses to write a balanced equation

Step	Example
Write the equation for the reaction (unbalanced)	
write down the mass or % <u>given in the question</u>	
Write the mass of one mole of each element or compound	
Divide the mass given in question by the mass of one mole	
Turn the answers into whole number simple ratio	
Put the numbers into the equation	

1. What is a limiting reactant?
2. Complete the calculation:
Calculate the mass of calcium needed to make 224g of calcium oxide

Step	Calculation
Write the balanced equation	$2\text{Ca} + \text{O}_2 \rightarrow 2\text{CaO}$
Write the masses of each substance	
Write down the given mass in the question.	
Work out the 'scale' factor (ie what did you have to do to the original number to get to the desired mass)	
Do the same to the other side	

- Concentration of a solution can be measured in mol. per given volume of solution e.g. mol. per dm^3 (mol./ dm^3)
- Mass of a solute and the volume of a solution are related to the conc. of the solution through the equation $\text{moles} = \text{concentration} \times \text{volume}$
 - Use mass to find mol: $\text{mol.} = \text{mass} / \text{molar mass}$, then use $\text{conc.} = \text{mol.} / \text{vol.}$ as seen below
- If the volumes of two solutions that react completely are known and the concentration of one solution is known, the concentration of the other solution can be calculated:
 - work out moles of solution where volume and concentration is known by $\text{moles} = \text{conc} \times \text{vol}$ (make sure volume is in dm^3 , to go from cm^3 to $\text{dm}^3 \div 1000$)
 - work out moles of other solution by mole ratio from equation
 - now work out unknown concentration by using $\text{conc} = \text{mol} / \text{vol}$



3.4 USING CONCENTRATIONS OF SOLUTIONS IN mol/dm^3 (chemistry only) (higher tier)

Can be measured in
moles per unit volume

Describes the number of
moles of a substance in a
certain volume of a solution

When the volumes of two reacting
solutions are known and the
concentration of one solution is
known, calculate the
concentration of the other solution

Units of concentration:
 mol/dm^3

Calculations

Need to know ratio of
moles reacting

$$\text{Concentration} = \frac{\text{Moles}}{\text{Volume}}$$

Where volume is measured in dm^3

AQA

Percentage yield

$$\text{Percentage yield} = \frac{\text{Amount of product produced}}{\text{Maximum amount of product possible}} \times 100$$

- It is not always possible to obtain the calculated amount of a product for 3 reasons:
 - Reaction may not go to completion because it is reversible
 - Some of the product may be lost when it is separated from the reaction mixture
 - Some of the reactants may react in ways different to the expected reaction
- Yield: amount of product obtained
- **To calculate the theoretical mass of a product from a given mass of reactant and the balancing equation for the reaction:**
 - Calculate mol. of reactant by using $\text{mol.} = \text{mass} / \text{molar mass}$
 - Use balancing numbers to find mol. of product (e.g. $2\text{HCl} + \text{Mg} \rightarrow \text{MgCl}_2$ if you have 2 mol. of HCl, you would divide by 2 to get 1 mol. of MgCl_2 .)
 - Calculate theoretical mass of a product by then using $\text{mass} = \text{mol.} \times \text{molar mass}$

Atom economy

- A measure of the amount of starting materials that end up as useful products
- Important for sustainable development and for economic reasons to use reactions with high atom economy

$$= (\text{Mr of desired product from reaction} / \text{sum of Mr of all reactants}) \times 100$$

- **Possible reasons why a particular reaction pathway is chosen/not chosen:**
atom economy, yield, rate, equilibrium position and usefulness of by-products

3.3 YIELD AND ATOM ECONOMY OF CHEMICAL REACTIONS (chemistry only)

Atom Economy

Raw materials need to be bought more regularly for low atom economy processes

Waste products can be expensive to remove

Higher atom economy is better for profits

The atom economy is likely to be lower, the more products there are

Higher atom economy implies a 'greener' process

Low atom economy produces waste products which need to be disposed of

Landfill

Calculations

$$\left(\frac{\text{Relative formula mass of desired product}}{\text{Sum of relative formula masses of all reactants}} \right) \times 100$$

Measure of amount of reactants which end up as useful products

Percentage yield

Calculations

$$\left(\frac{\text{Mass of actual yield}}{\text{Mass of theoretical yield}} \right) \times 100$$

Not always possible to obtain theoretical amount of product

Unexpected side reactions

Product lost when separated from reaction mixture

Reversible reaction

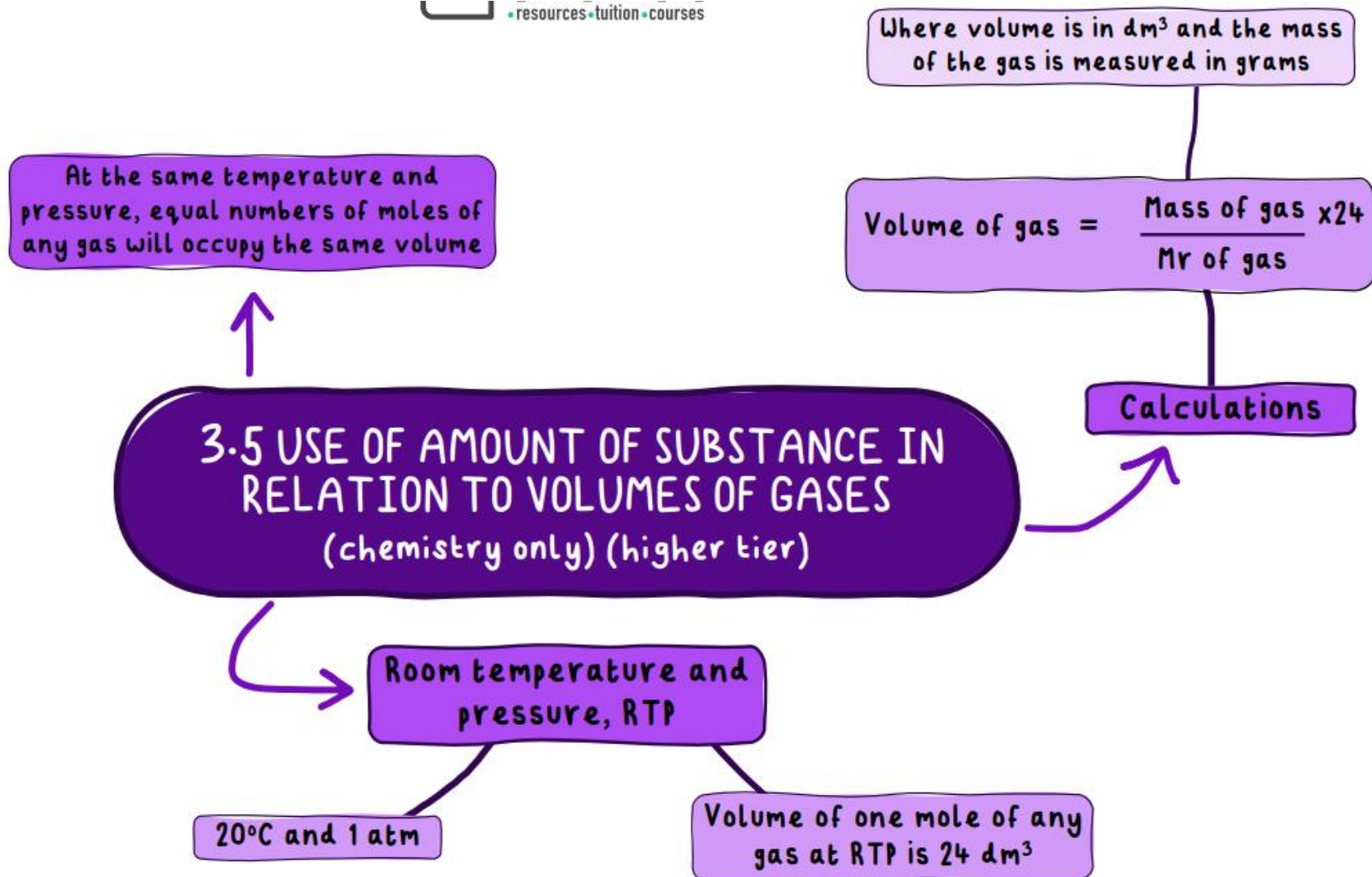
Incomplete reaction

E.g. Haber process

- Equal amounts (in mol.) of gases occupy the same volume under the same conditions of temperature and pressure (e.g. RTP)
- Volume of 1 mol. of any gas at RTP (room temperature and pressure: 20 degrees C and 1 atmosphere pressure) is 24 dm³
- This sets up the equation:

$$\text{Volume of gas (dm}^3\text{) at RTP} = \text{Moles} \times 24$$

- using this equation, if the reaction is at RTP, you can calculate moles of a gas produced and then x24 to get volume produced (e.g. if you produce 5 moles of hydrogen, you produce $24 \times 5 = 120 \text{ dm}^3$)



C4 – Chemical Changes

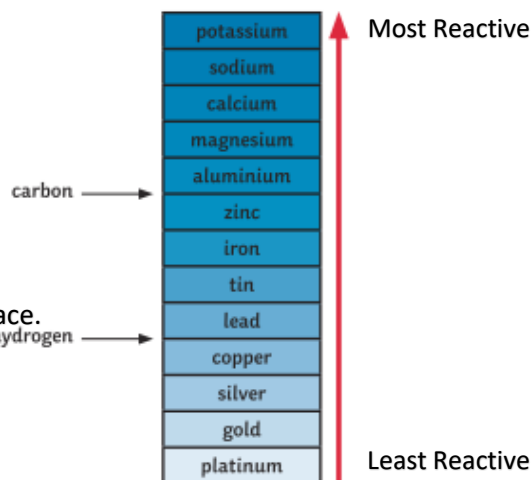
The Reactivity Series

- A more reactive metal will replace a less reactive metal in a compound (**displacement**)

- e.g. potassium + magnesium chloride → potassium chloride + magnesium

Potassium is more reactive than magnesium

Potassium **displaces** magnesium from the compound and takes its place.



Reactions of acids with metals

- Metal + acid → salt + hydrogen

E.g. iron + sulfuric acid → iron sulfate + hydrogen

iron sulfate

salt

To name salt :
1st name Metal
2nd name Acid used

Naming Salts

Acid used	Salt produced
Hydrochloric	Chloride
Sulfuric	Sulfate
Nitric	Nitrate

Extraction of Metals

- Extraction = remove metal from an ore or a compound.

Ore = a rock containing enough metal to make extracting metal worthwhile.

How to extract metals:

Less reactive than carbon – reduction with carbon

Reduction = loss of oxygen

E.g. iron oxide + carbon → iron + carbon dioxide

Oxygen has been removed to extract iron.

Carbon and the oxygen removed from the iron react to make carbon dioxide

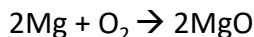
More reactive than carbon – electrolysis is used.

- Some metals are found in **native** form (not reacted, so in element form) – usually platinum and gold as **very unreactive**.

Reaction of metals with oxygen

- Metal + oxygen → metal oxide

e.g. magnesium + oxygen → magnesium oxide



Oxidation reaction as metal gained oxygen

- Oxidation = gaining oxygen

- Reduction = losing oxygen

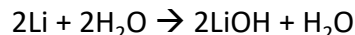
Reaction of metals with water

- Most metals don't react well with water

- Group 1 and group 2 react to form alkalis

- Metal + water → metal hydroxide + hydrogen

e.g. lithium + water → lithium hydroxide + hydrogen



Metal hydroxides are alkaline

Reactions of acids with alkalis

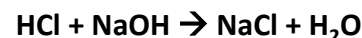
- Acid + alkali → salt + water

neutralisation

Hydrochloric acid + sodium hydroxide → sodium chloride + water

sodium chloride

salt



Reactions of acids with carbonates

- Acid + carbonate → salt + water + carbon dioxide

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

calcium chloride

salt



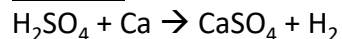
C4 – Chemical Changes

Redox Reactions (HT only)

- Redox = reduction and oxidation takes place at same time in a reaction.

- Metal + acid = redox reaction

Example



Ionic equation: $2\text{H}^+ + \text{Ca} \rightarrow \text{Ca}^{2+} + \text{H}_2$ Lost 2 electrons (oxidation)

Half equation 1: $\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$

Half equation 2: $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ Gained 2 electrons (reduction)

Strong/Weak Acids (HT only)

Strong acid = completely dissociates in a solution
e.g. $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$

Examples = nitric acid and sulfuric acid

Weak acid = partially dissociates in solution.

e.g. $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$
 $\rightleftharpoons$ = reversible reaction

Hasn't fully turned into ions – only partially

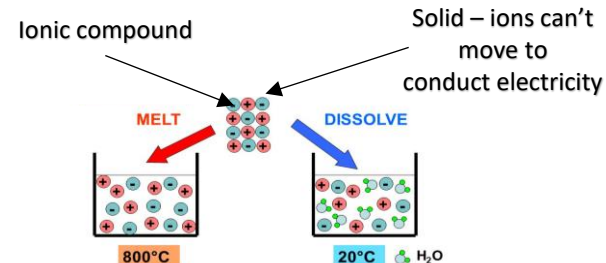
Concentration = how much is dissolved in every cm^3

Strong/weak = how well it ionises

As **pH** decreases by 1 unit, **hydrogen ion concentration** of solution increases by factor of 10

Electrolysis

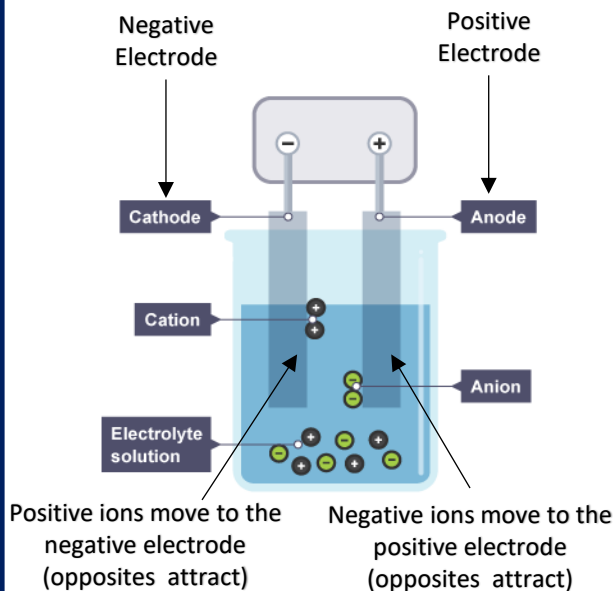
- **Splitting** up a **compound** using electricity.
- Used to extract metals from compounds, purify metals (eg copper)



- Must be **molten** or **aqueous** (dissolved in water) to allow **ions** to **move** to the electrodes

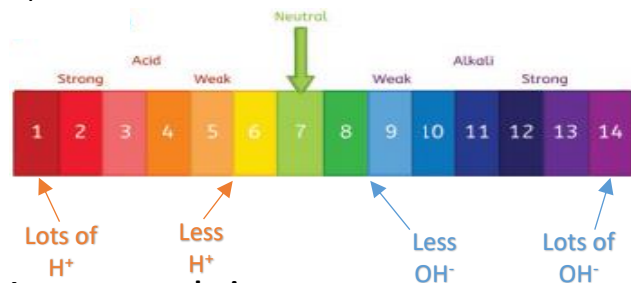
The Process of Electrolysis

Two **electrodes** – made of **inert** material (doesn't react)



pH Scale

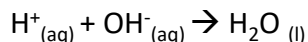
- Shows how acidic or alkaline solution is.
- pH 1-6 = acid
- pH 7 = neutral
- pH 8-14 = alkali



In aqueous solutions:

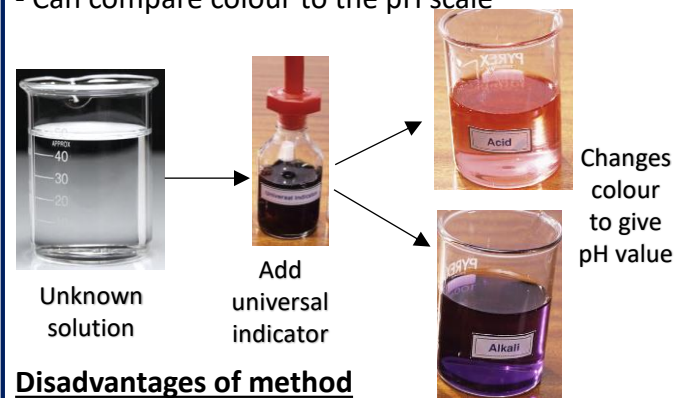
Acids – produce H^+ ions
Alkalis – produce OH^- ions

In neutralisation reactions:



Measuring pH of a solution

- Can use **universal indicator**
- Gives the solution a colour
- Can compare colour to the pH scale



Disadvantages of method

- Colour is **subjective** – different people may see different colours
- Doesn't give an exact pH number (could use **pH probe** to make more **accurate**).

C4 – Chemical Changes – Required Practical – Preparation of soluble salts

Aim

Prepare a pure, dry sample of a soluble salt from an insoluble **oxide or carbonate**.

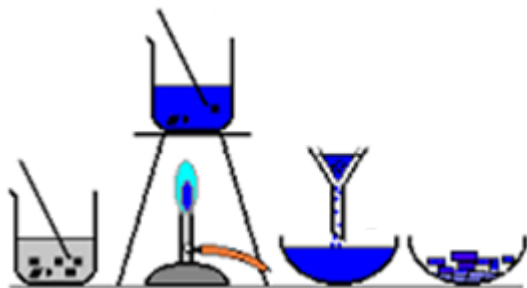
Equipment

- Beaker
- Measuring cylinder
- Bunsen burner and safety mat
- Filter funnel and filter paper
- Named acid (e.g. hydrochloric acid)
- Metal oxide or carbonate.
- Spatula
- Glass stirring rod

Change method
depending on reactants in
the question.

Method (example copper oxide and sulfuric acid to make copper sulfate)

1. Using measuring cylinder – 20cm³ **sulfuric acid** → beaker
2. Warm the acid gently (not boiling)
3. Using spatula add **copper oxide** to the acid and stir
4. Keep adding until no more oxide will dissolve (excess).
5. Using a filter funnel and filter paper – filter excess copper oxide.
6. Evaporate some of the filtrate using a water bath.
7. Pour remaining filtrate into an evaporating basin – leave overnight to evaporate water
8. Pat the crystals dry.



Common questions

Q1) Why do you heat the acid before adding the oxide?

A1) To speed up the reaction (particles have more energy to react).

Q2) Why is the oxide added in excess?

A2) To make sure that all the acid has been neutralised.

Q3) Why is the solution filtered?

A3) Remove any unreacted, excess solid.

Q4) Why is the solution left overnight in a warm, dry place?

A4) To evaporate excess water, to form crystals (crystallise).

Q5) Name 2 safety precautions you should take during this practical.

A5) Safety goggles and allow equipment to cool before putting away

C4 – Chemical Changes

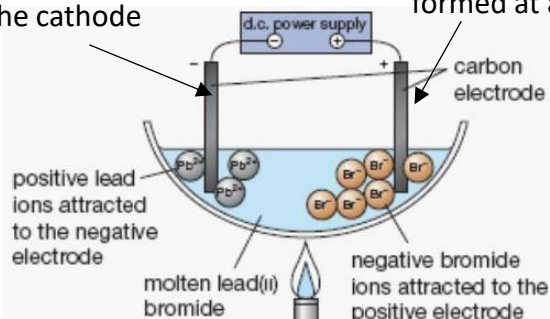
Electrolysis of Molten Ionic Compounds

Molten = melted so ions can move.

- Metal = produced at **anode**
- Non-metal = produced at **cathode**

Example: Lead Bromide - PbBr_2

Lead forms at the cathode
Bromine gas is formed at anode



Using Electrolysis to Extract Metals

- Used if metal is **too reactive** to be extracted by reduction with carbon.
- Requires **large amount of energy** to melt the compound and produce electrical current. (**expensive**)

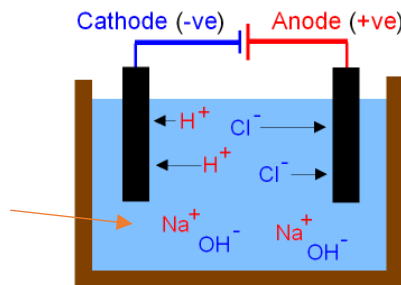
Example: Aluminium Oxide

- **Cryolite** is added – reduces the melting point (less energy needed – less expensive)
- **Carbon** used as positive electrode – needs to be replaced constantly as **oxygen** will react with it to produce CO_2 – it will degrade.

Electrolysis of Aqueous Solutions

- Compound is dissolved in water so ions can move.

When aqueous – H^+ and OH^- (from H_2O) are also present along with the two ions from the compound.



- Only **one** ion is discharged at each electrode.

Anode – Non-metal or oxygen

Cathode – Metal or hydrogen

Rules

+ ANODE Attracts – ions ('Anions')	- CATHODE Attracts + ions ('Cations')
If – ions are group 7 i.e. chloride Cl^- bromide Br^- iodide I^- Then the groups 7 element is produced as a gas If – ions are NOT Group 7 Eg sulphate SO_4^{2-} nitrate NO_3^- carbonate CO_3^{2-} OXYGEN is produced.	If + ions (metals) are MORE REACTIVE than hydrogen K, Na, Ca, Mg, Zn, Fe Then HYDROGEN is produced If + ions (metals) are LESS REACTIVE than hydrogen Cu, Ag, Au Then the METAL is produced

Examples

Solution	Product at cathode	Product at anode
Potassium chloride	Hydrogen – because K is more reactive than H	Chlorine – as it is a halogen
Copper sulfate	Copper – as copper is less reactive than H	Oxygen – as there is no halogen

Half-Equations at Electrodes (HT only)

During electrolysis:

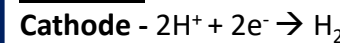
Cathode – positive ions **gain** electrons (**reduction**)

Anode – negative ions **lose** electrons (**oxidation**)

- Ions become **discharged** (lose their charge) at the electrodes to form the atoms again.

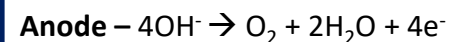
- Reactions at electrodes can be represented by half equations.

Examples



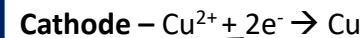
Gained 2 electrons (reduction)

molecules of hydrogen gas produced



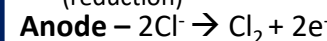
molecules of oxygen produced

Lost electrons (oxidation)



Gained electrons (reduction)

Copper atoms are formed at the cathode



chlorine molecules are formed

Lost electrons (oxidation)

Titration (chemistry only)

The volumes of acid and alkali solutions that react with each other can be measured by titration using a suitable indicator.

How to carry out a titration:

1. Wash burette using dilute hydrochloric acid and then water
2. Fill burette to 100cm³ with acid with the meniscus' base on the 100cm³ line
3. Use 25cm³ pipette to add 25cm³ of alkali into a conical flask, drawing alkali into the pipette using a pipette filler
4. Add a few drops of a suitable indicator to the conical flask (eg: phenolphthalein which is pink when alkaline and colourless when acidic)
5. Add acid from burette to alkali until end-point is reached (as shown by indicator)
6. The titre (volume of acid needed to exactly neutralise the acid) is the difference between the first (100cm³) and second readings on the burette
7. Repeat the experiment to gain more precise results

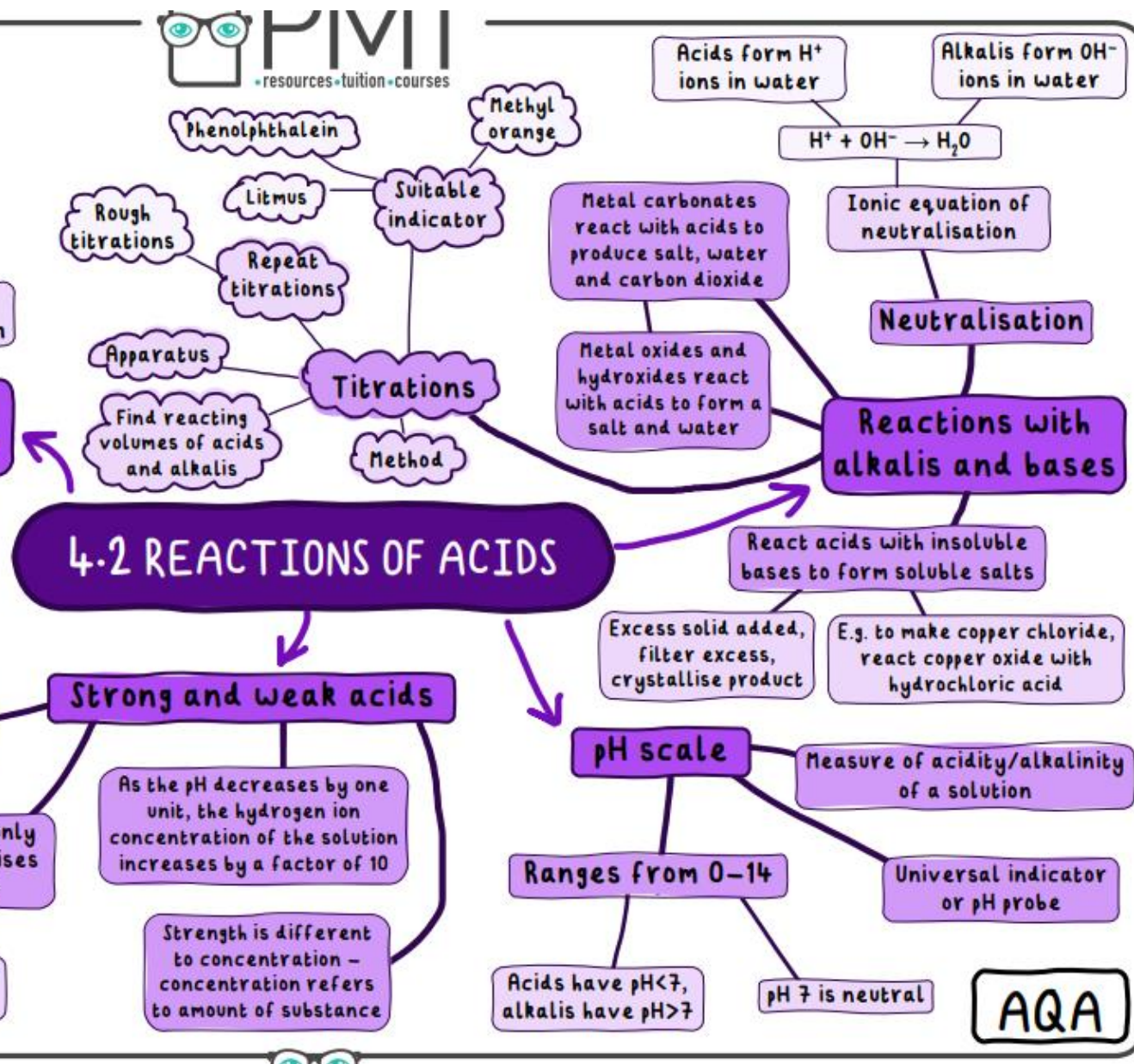
Titration calculations

- 1dm³ = 1000cm³
- One mole of a substance in grams the same as its relative atomic mass in grams.

Working out concentrations:

E.g 25 cm³ of dilute hydrochloric acid is neutralised by 20 cm³ of 0.5 mol/dm³ sodium hydroxide. What is the concentration of the hydrochloric acid?

1. Convert volumes into dm³.
 $25/1000=0.025\text{dm}^3$ $20/1000=0.02\text{dm}^3$
2. Work out the moles of NaOH
moles = volume x concentration
So, $0.02 \times 0.5 = 0.01$
3. Work out mole ratio from equation
 $\text{HCl} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{NaCl}$
1:1 ratio, so moles of NaOH = moles of HCl, so moles of HCl=0.01
4. Work out concentration
conc = moles / vol = $0.01 / 0.025 = 0.4 \text{ mol dm}^{-3}$



KEY

'Chemistry only' written in clouds.

'Higher only' written in yellow.

AQA

C5 – Energy Changes

Exothermic Reactions

- Energy transferred to the surroundings
- Temperature of the reaction mixture **increases**
- This energy is transferred **to** the surroundings

Examples include:

- Hand warmers
- Combustion reactions
- Respiration
- Neutralisation reactions
- Self-heating cans.



Exothermic

Endothermic Reactions

- Energy absorbed from the surroundings
- Temperature of reaction mixture often **decreases**
- Energy is transferred **from** the surroundings

Examples include:

- Ice packs (injuries)
- Reaction of citric acid and sodium hydrogen carbonate
- Thermal decomposition of calcium carbonate



Endothermic

Energy change of reactions (HT)

During a reaction:

- Energy is **absorbed** in order to **break** bonds in the reactants 
- Energy is **released** when bonds are **made** in the products. 

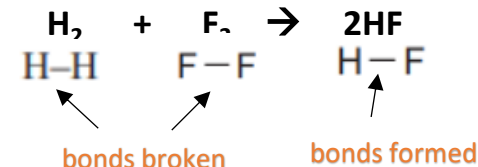
Bond energy = the amount of energy that is released when a bond is made or that is needed to break a bond

Calculating energy changes (HT)

Overall energy change = difference between energy needed to break bonds and the energy released when bonds formed.

To calculate energy change :

Energy change = bonds broken – bonds formed



Bond	Bond Energy / kJ mol^{-1}
F—F	158
H—H	436
H—F	568

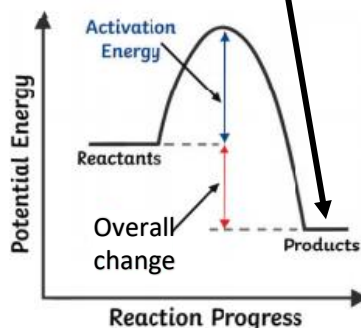
Bonds broken = $436 + 158$ 593	Bonds formed 2×568 1136
--	--

**Overall energy change = $593 - 1136$
= -543 kJ/mol Exothermic**

More energy is released in bond making than is required for bond breaking.

Reaction Profiles – Exothermic

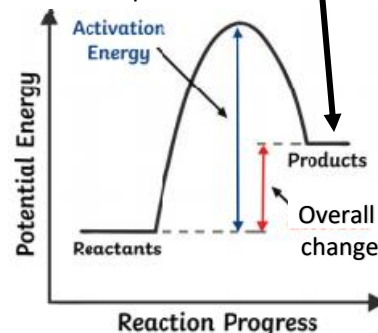
- Energy level diagrams show **difference in energy** between reactants and products.
- Exothermic = Energy of products is **lower than** reactants (energy is released)
- **Activation Energy** = minimum amount of energy needed to start the reaction.
- **Energy change** = the difference in energy between reactants and products.



You may need to draw and label this in the exam!

Reaction Profiles – Endothermic

- Energy level diagrams show **difference in energy** between reactants and products.
- Endothermic = Energy of products is **higher than** reactants (energy is absorbed)
- **Activation Energy** = minimum amount of energy needed to start the reaction
- **Energy change** = the difference in energy between reactants and products.



You may need to draw and label this in the exam!

C5 – Energy Changes

1. Which way is energy transferred in an exothermic reaction?
2. What happens to the temperature of the reaction mixture in an exothermic reaction?
3. State two examples of exothermic reactions.

1. Which way is energy transferred in an endothermic reaction?
2. What generally happens to the temperature of the reaction mixture of an endothermic reaction?
3. State two examples of endothermic reactions.

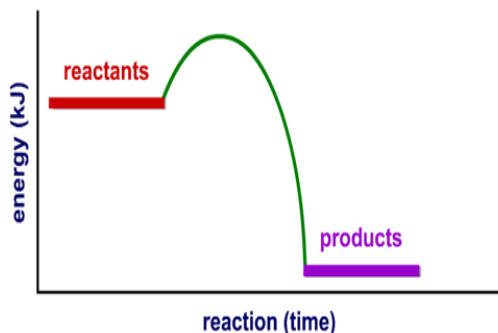
Higher Tier only

1. In terms of energy, what happens for bonds to be broken?
2. In terms of energy, what happens when bonds are formed?

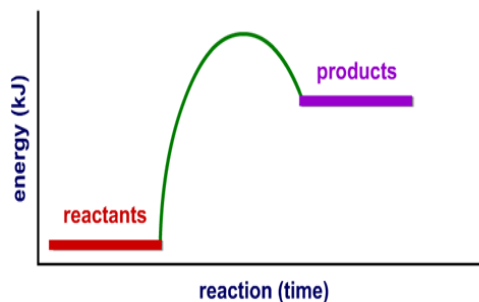
Higher Tier only

1. Define overall energy change.
2. How do you calculate energy change?
3. Why, in terms of bond breaking and making, is a reaction exothermic?
4. Why, in terms of bond making and breaking, is a reaction endothermic?

1. Define activation energy.
2. On the graph below, draw and label the :
 - overall energy change
 - activation energy



1. What does an energy level diagram show?
2. On the graph below, draw and label the :
 - overall energy change
 - activation energy



C5 – Energy Changes – Required Practical – Temperature Changes

Hypothesis

The energy change in the reaction between acid and alkali depends on the volume of alkali added.

Equipment

- Polystyrene cup and lid
- Thermometer
- 250cm³ beaker
- Measuring cylinder
- Liquid reactants



Method (example for hydrochloric acid and sodium hydroxide)

1. Using measuring cylinder to measure 30cm³ hydrochloric acid and put in polystyrene cup
2. Stand cup inside beaker to make stable.
3. Use a thermometer to measure the temperature of acid and record.
4. Using measuring cylinder – 5cm³ sodium hydroxide → polystyrene cup
5. Fit the lid and gently stir with thermometer through hole.
6. When reading stops on thermometer, record temperature in table.
7. Repeat, each time adding 5cm³ more sodium hydroxide up to a maximum of 40cm³.
8. Calculate the temperature change on each attempt.
9. Repeat the experiment 3 times and calculate a mean temperature change for each volume of sodium hydroxide.

Variables

Independent – Volume of sodium hydroxide

Dependent – Temperature change

Control – Volume of hydrochloric acid, concentration of acid, concentration of sodium hydroxide

Common questions

Q1) Why do you use a polystyrene cup and lid?

A1) Because polystyrene cups are insulators, which reduces heat loss in the experiment, making the results more accurate.

Q2) Why should you calculate the temperature change, instead of just using the final temperature?

A2) Because the initial (starting) temperature of the acid may have been different.

Q3) Why is it important to stir the mixture?

A3) To make sure all of the reactants have reacted and to get a uniform temperature.

Q4) Why is the experiment conducted 3 times?

A4) So that anomalies can be seen and removed and a mean calculated

Energy changes could also be investigated using:

1. Changing the **mass of metal** added to acid and measuring the **temperature increase**
2. Changing the **type of metal** added to acid and measuring the **temperature increase**
3. Dissolving different **masses of potassium nitrate** into water and observing the **temperature decrease**.

Cells and batteries

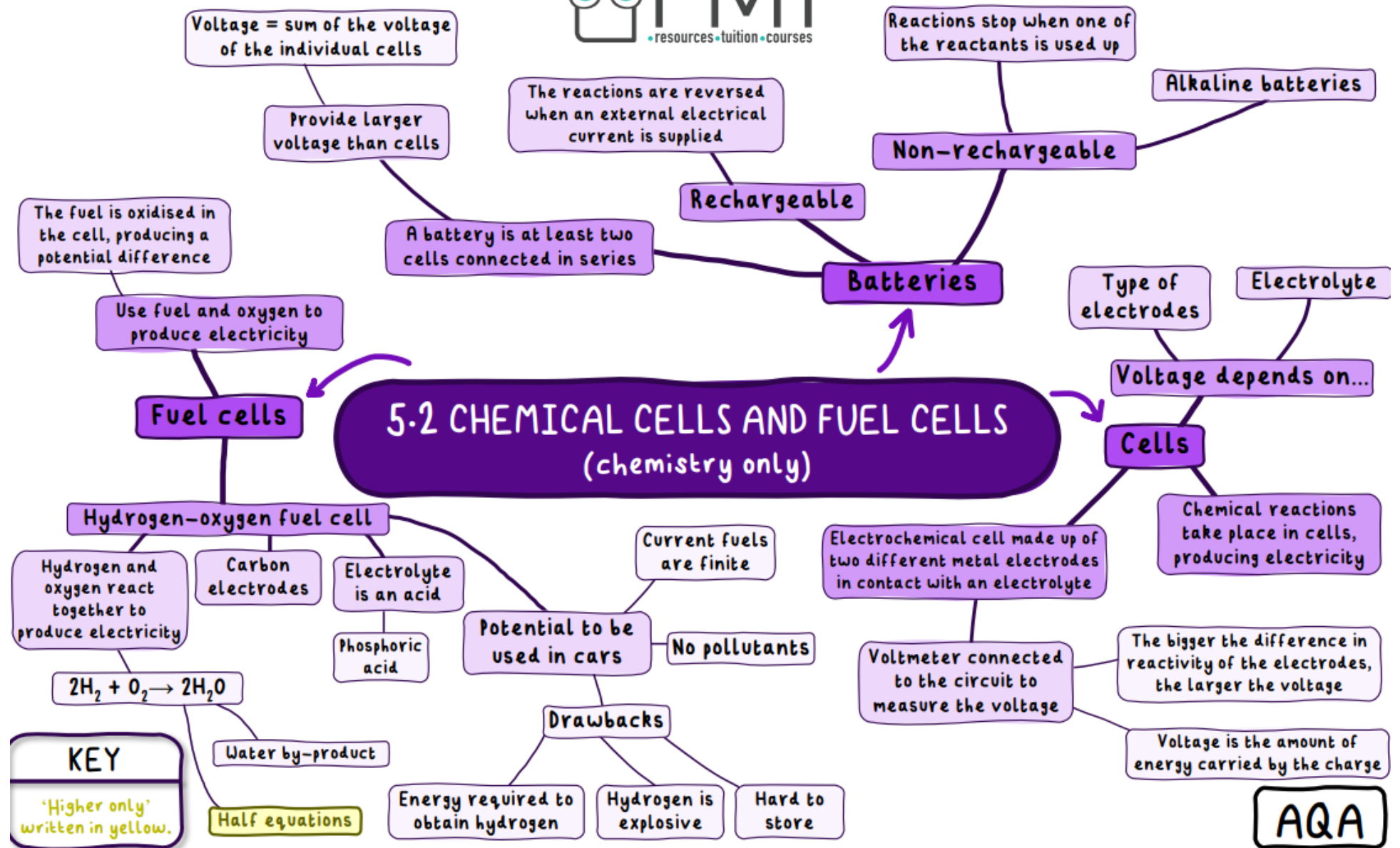
- Cells contain chemicals which react to produce electricity
- The voltage produced by a cell is depended upon a number of factors
 - E.g. type of electrode & electrolyte
- A simple cell can be made by connecting two different metals in contact with an electrolyte
- Batteries = two or more cells connected together in series to provide a greater voltage
- Non-rechargeable cells & batteries:
 - Chemical reactions stop when one of the reactants has been used up
 - Alkaline batteries are non-rechargeable
- Rechargeable cells & batteries:
 - Can be recharged because the chemical reactions are reversed when an external electrical current is supplied

Fuel cells

- Supplied by an external source of fuel (e.g hydrogen) and oxygen or air. the fuel is oxidised electrochemically within the fuel cell to produce a potential difference
- Overall reaction a hydrogen fuel cell involves the oxidation of hydrogen to produce water
- Hydrogen fuel cells offer a potential alternative to rechargeable cells & batteries:

hydrogen fuel cell	rechargeable cells and batteries
fuel cells can be used constantly provided fuel keeps being put in	can be recharged by reversing reaction, so fuel doesn't need to keep being supplied
hydrogen is a gas so needs to be stored at high pressure and so is harder to transport	hard to dispose of- non-biodegradable
only produces water when burnt	will eventually stop working

- equations for each half cell:
 - **At the anode (positive electrode):** $\text{H}_2(\text{g}) \rightarrow 2\text{e}^- + 2\text{H}^+(\text{aq})$
 - **At the cathode (negative electrode):** $4\text{H}^+(\text{aq}) + \text{O}_2(\text{g}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{g})$



KEY
 'Higher only' written in yellow.

Half equations

AQA

C6 – Rate and extent of chemical change

Rate of reaction.

Measuring the rate of anything always involves a **measurement of time**

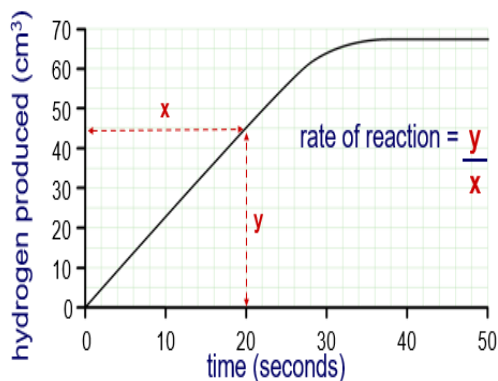
The rate of a chemical reaction can be found using:

$$\text{rate} = \frac{\text{quantity of reactant used}}{\text{time}}$$

$$\text{rate} = \frac{\text{quantity of product formed}}{\text{time}}$$

Quantities for reactants or products are measured in **mass in g** or by **volume in cm³**

Rate calculations can be done from tables of data or graphs:



Volume of hydrogen produced = 45cm³

Time taken = 20 seconds

Rate = $\frac{45}{20}$ cm³

20 s

rate = 2.25 cm³/s

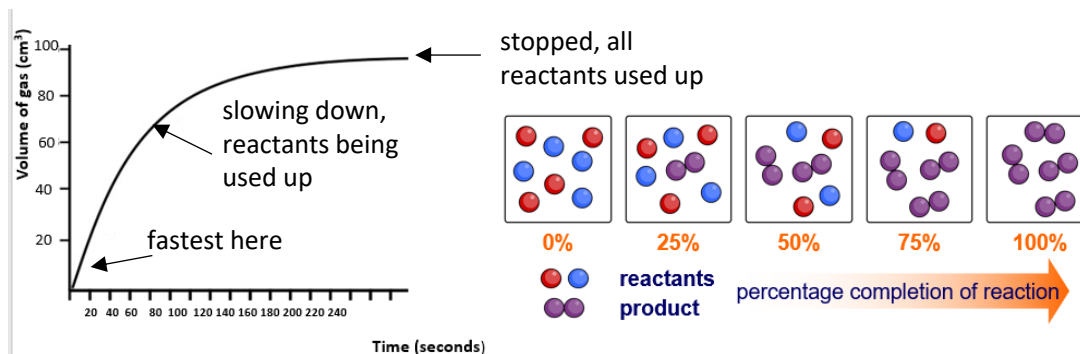
The progression of a chemical reaction

For a reaction to take place, reactant particles have to collide.

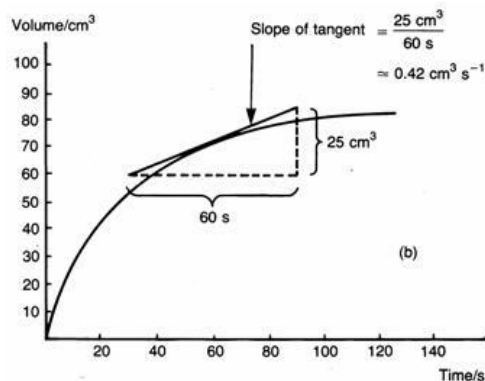
The rate of a reaction depends on the **frequency of collisions** and **the energy with which the particles collide**.

The minimum amount of energy needed to start a reaction is called the **activation energy**.

A reaction is always **fastest at the beginning** and slows down over time as the reactants get used up and the frequency of collisions decreases.



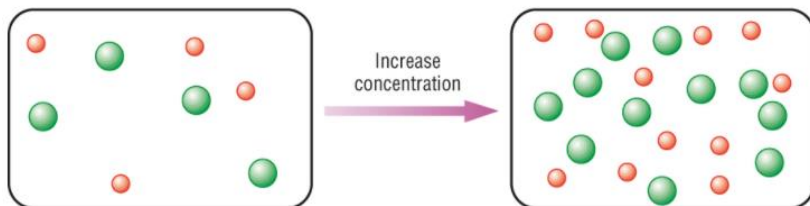
Using a tangent to calculate rate (HT)



- Draw a line along the point you're interested in. The line should touch the curve at the point given.
- Make a triangle. Try to make the angles either side of the line equal.
- Measure the change in volume and change in time
- Calculate the gradient
- Use units from the axes to determine the units for rate

C6 – Rate and extent of chemical change

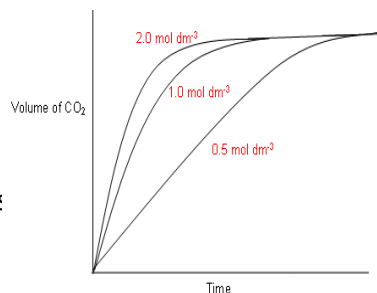
The effect of concentration



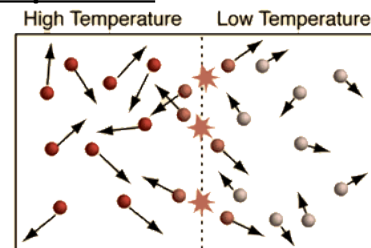
Concentration means number of particles per cm^3

Increasing the concentration of any of the reactants increases the rate of the reaction

This is because there are more particles per cm^3 so there are **more frequent collisions**, increasing the rate.

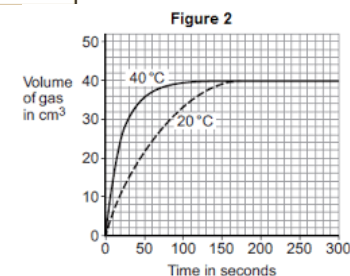


The effect of temperature



Increasing the temperature of the reactants increases the rate of the reaction.

This is because the particles have more kinetic energy and therefore move faster, so there are **more frequent collisions**, increasing the rate.

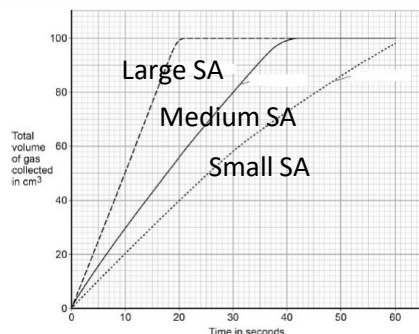


The effect of surface area

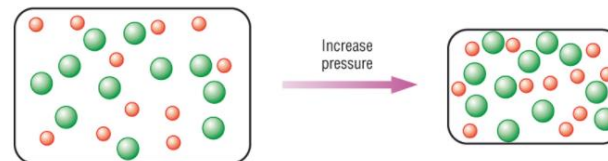


The smaller the pieces of a solid, the higher the surface area. Increasing the surface area of solid reactants increases the rate of reaction.

This is because there is a greater area available for collisions to occur so there are **more frequent collisions**, increasing the rate.

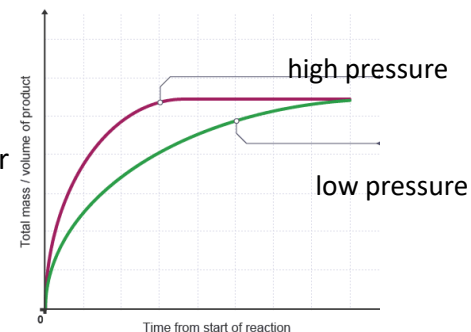


The effect of pressure



Increasing the pressure of gaseous reactions increases the rate of the reaction.

This is because the same number of particles are now in a smaller volume, so there are **more frequent collisions**, increasing the rate.

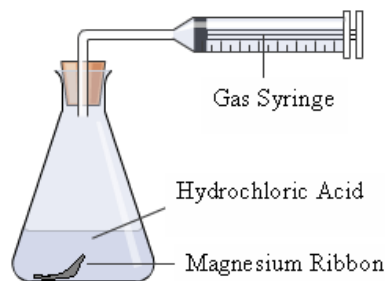
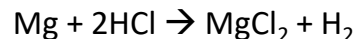


In all cases, the overall amount of product is the SAME, the end point of the reaction is just reached faster

C6 – Required practical – the effect of concentration on rate of reaction

Experiment 1

Using volume of gas collected over time as a measure of the rate



Independent variable: concentration of HCl

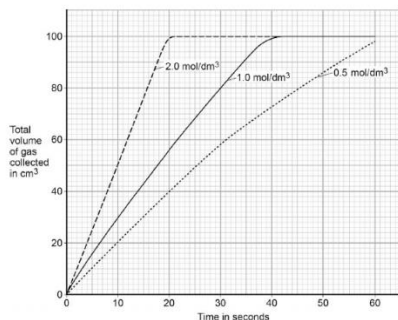
Dependent variable: Volume of gas produced / min

Control variables: volume of HCl, mass of Mg, temperature of acid

Method

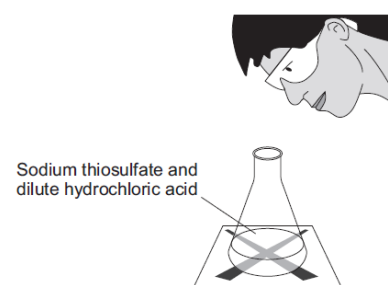
1. Measure 20cm³ 0.5M HCl into a conical flask.
2. Insert 2 x 2cm pieces of Mg and attach a gas syringe
3. Start a stopwatch and measure the volume of gas collected every 20 seconds until the reaction is over.
4. Repeat using different concentrations of HCl.

An increase in the concentration leads to an increase in the rate of the reaction, but the same volume of product overall



Experiment 2

Investigating the effect of changing the concentration of HCl on the rate of reaction



Independent variable: concentration of HCl

Dependent variable: Time taken for the cross to disappear

Control variables: volume of HCl, volume of sodium thiosulphate, temperature of both solutions, concentration of sodium thiosulphate

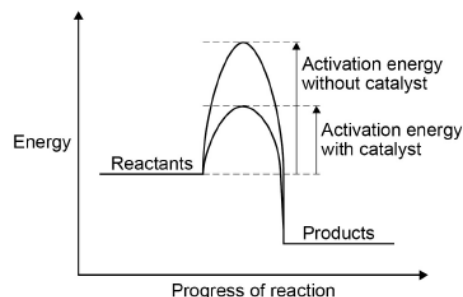
Method

1. Use a measuring cylinder to put 10 cm³ sodium thiosulfate solution into the conical flask.
2. Put the conical flask on the black cross.
3. Put 10 cm³ of 0.5M hydrochloric acid into the 10 cm³ measuring cylinder.
4. Put this acid into the flask. At the same time swirl the flask gently and start the stopwatch.
5. Look down through the top of the flask. Stop the stopwatch when you can no longer see the cross. Record the time.
6. Repeat steps 1-5 using different concentrations of HCl – 1M, 1.5M, 2M and 2.5M

C6 – Rate and extent of chemical change

Catalysts

- Catalysts are substances that speed up chemical reactions without themselves being used up.
- They provide a different pathway for the reaction with a lower activation energy.
- Different reactions require different catalysts.



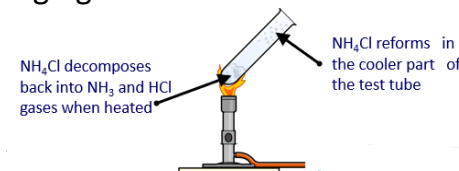
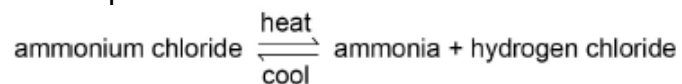
Reversible reactions

These are reactions in which the products can react to produce the original reactants

They are represented by the symbol $\rightleftharpoons$

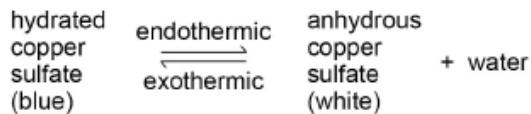
The direction of the reaction can be changed by changing the conditions

For example:



If a reaction is exothermic in one direction, it is endothermic in the opposite direction.

The same amount of energy is transferred in each case.



When a reversible reaction takes place in sealed apparatus, then a point occurs when the forward and backward reactions occur at the same rate. This is **equilibrium**

The effect of changing conditions on equilibrium (HT)

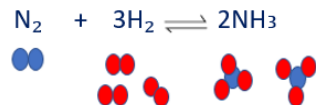
If a system is at equilibrium and a change is made to the conditions, then the system responds to counteract the change.

E.g. – if the temperature is increased, then the system will respond by increasing the rate of the endothermic reaction, to bring the temperature back down

If the concentration of the reactants is increased, then equilibrium will shift right and more products will be made.

In gaseous reactions, a change in pressure will result in equilibrium shifting to the side that restores the pressure.

E.g. :



In this reaction, there are 4 moles of gas on the reactants side and only 2 on the product side

If the pressure is increased, equilibrium will shift right as there are fewer moles on the products side, and this will decrease the pressure.

The effect of changing conditions on equilibrium

The relative amounts of all the reacting substances at equilibrium depend on the conditions of the reaction.

If a system is at equilibrium and a change is made to any of the conditions, then the system responds to counteract the change (Le Chatelier's principle)

The effect of changing concentration

- If the concentration of one of the reactants or products is changed, the system is no longer at equilibrium and the concentrations of all the substances will change until equilibrium is reached again.
- If **concentration of reactants is increased**: position of equilibrium shifts towards products so **more product is produced** until equilibrium is reached again
- If **concentration of products is increased**: position of equilibrium shifts towards reactants so **more reactant is produced** until equilibrium is reached again

The effect of temperature changes on equilibrium

- If **temperature is increased**: equilibrium moves in the **direction of the endothermic reaction** (e.g. if forwards reaction is endothermic and temperature is increased, equilibrium shifts right to produce more product)
- If **temperature is decreased**: equilibrium moves in the **direction of the exothermic reaction**
- For the forwards being exo/endothermic and yield meaning the amount of product from the forwards reaction:

	Exothermic	Endothermic
An increase in temperature...	Decreases yield of reaction	Increases yield of reaction
A decrease in temperature...	Increases yield of reaction	Decreases yield of reaction

The effect of pressure changes on equilibrium

- In gaseous reactions, an increase in pressure will favour the reaction that produces the least number of molecules as shown by the symbol equation for that reaction.
- If **pressure is increased**: equilibrium shifts to **side of equation with fewer moles of gas** (e.g. $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$, left side has 4 moles of gas (1+3) and right has 2 moles of gas. If you increase the pressure equilibrium moves right as there are fewer moles of gas on the right hand side, making more product)
- If **pressure is decreased**: equilibrium will shift to **side of equation with more moles of gas** (e.g. for previous example equilibrium would move left, making more reactant)

If a reaction produces a...	...larger volume of gas (more moles)	...smaller volume of gas (fewer moles)
An increase in pressure...	Decreases yield of reaction	Increases yield of reaction
A decrease in pressure...	Increases yield of reaction	Decreases yield of reaction

E.g. hydrated copper sulfate $\rightleftharpoons$ anhydrous copper sulfate + water

Forwards reaction is endothermic, backwards reaction is exothermic

Denoted by:
 $\rightleftharpoons$

The products can react to produce the original reactants

The reaction will be endothermic in one direction and exothermic in the other

Energy changes

Equal amounts of energy transferred in each direction

If you change the conditions of a reversible reaction, the system will try to counteract the change

Direction changed by altering the conditions

Reversible reactions

Reversible reactions will reach equilibrium

Nothing can get in

Closed system

None of the reactants or products can escape

Equilibrium

The forwards and backwards reaction are occurring at the same rate

Depends on temperature, pressure and concentration

Explain in terms of Le Chatelier's Principle

6.2 REVERSIBLE REACTIONS AND DYNAMIC EQUILIBRIUM

Le Chatelier's Principle

Change in temperature

Decreasing the temperature will shift equilibrium in the exothermic direction

Change in pressure

Increasing the pressure will move equilibrium in the direction where there are fewer molecules of gas

Change in concentration

Increasing the concentration of the reactants will increase the rate of the forwards reaction

Equilibrium will shift to the right

KEY

'Higher only' written in yellow

AQA

C7 – Organic Chemistry

Crude oil

Crude oil = a mixture of **hydrocarbons**.

- It is a **non-renewable resource (fossil fuel)**
- Made from remains of dead sea creatures **compressed** over millions of years

Hydrocarbons - molecules containing **hydrogen** and **carbon only**.

Two types of hydrocarbons are **alkanes** and **alkenes**.
The hydrocarbons in crude oil are mostly alkanes.

Alkanes

- Alkanes = **saturated** hydrocarbons.
- Held together by **single covalent bonds**.
- General formula = C_nH_{2n+2}
- Have different boiling points – longer the chain, higher the boiling point

You need to remember the names, and formulas of the first 4 alkanes.

Name of Alkane	Structural Formula	Molecular Formula
methane	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$	CH_4
ethane	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array}$	C_2H_6
propane	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array}$	C_3H_8
butane	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$	C_4H_{10}

Fractional Distillation

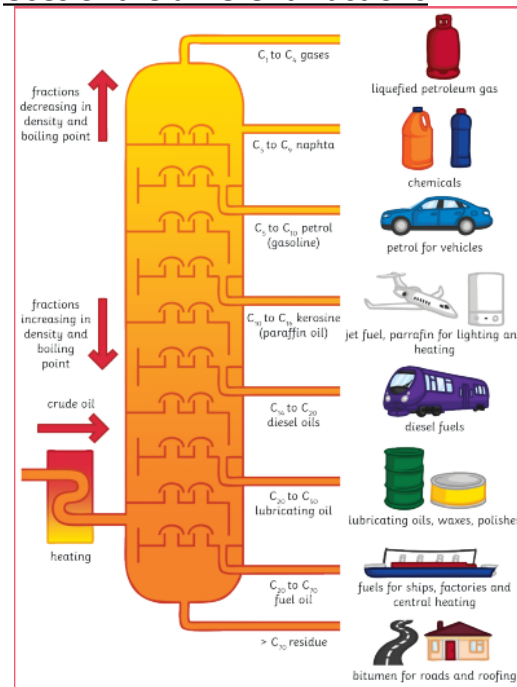
- Used to **separate** the mixtures of hydrocarbons in **crude oil**.

Steps in Fractional Distillation

1. Crude oil enters **fractioning column** and is heated to boiling point so the hydrocarbons evaporate.
2. It is **cooler** at the **top** of the fractionating column and **hotter** at the **bottom**.
3. Vapours rise up the column and, as they rise, they cool
4. The different hydrocarbons condense at different **boiling points**
5. The different 'fractions' have different **properties**

Short-Chain Molecules	Increasing Chain Length	Long-Chain Molecules
		
thin		thick
		
		

Uses of the different fractions



Supply and demand

Product	Supply in tonnes	Demand in tonnes
petrol	100	300
diesel	200	100
heating oil	250	50

After fractional distillation, we find:

- we have more of the long chain hydrocarbons than we need
- There are not enough short chain hydrocarbons.
- Short chain are more useful as they are more flammable so can be used as fuels.

C7 – Organic Chemistry

Cracking

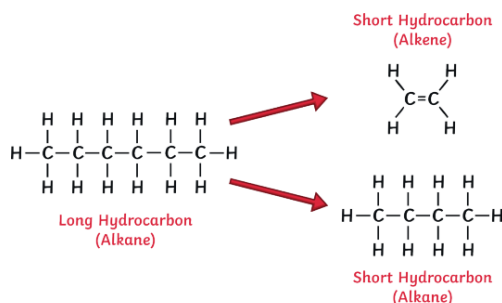
- This is done to solve the problem of having too many long chain hydrocarbons and not enough short ones
- Long hydrocarbons are **broken down** into smaller, more useful hydrocarbons.
- Short chain hydrocarbons are more useful as they are more flammable

Two types of cracking: catalytic and steam cracking.

Catalytic cracking – needs a **high temperature** and a **catalyst**.

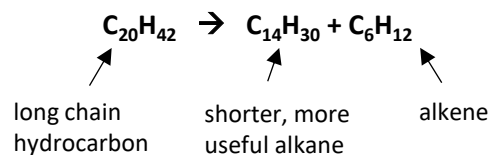
Steam cracking – **high temperature** and **steam**

- Cracking produces a **short-chain alkane** and an **alkene**.



Cracking equations

Same number of carbon and hydrogen atoms on both sides of the equation:



Alkenes

- Alkenes are **unsaturated** hydrocarbons.
- Contain carbon-carbon **double bonds**.

Test for Alkenes

Use bromine water to test for alkenes.

If an alkene is present, the bromine water turns from orange/brown to colourless.

Alkanes do not react with bromine water.

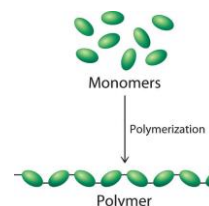


Uses for alkenes:

- Can be used as fuels
- Can be used as a starting material for other chemicals
- Can be used to make polymers (e.g. plastic)

Polymers

- Polymers are large molecules made of many repeating units (monomers)
- Alkenes (small molecules) are joined together to make polymers



Poly(ethene) – plastic bags/drinks bottles

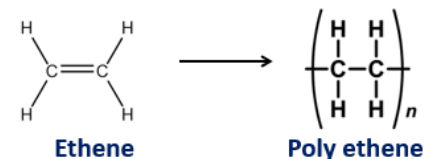
Poly(propene) – strong tough plastics

Drawing and naming polymers

1. Redraw the **monomer given**, but without the double bond. Make sure to copy all other elements exactly.
2. Put brackets around the monomer and extend joining bonds out through the brackets on both sides
3. Add an 'n' at the bottom right of the bracket
4. To name the polymer, you put **poly** in front of the monomer name

E.g.:

Draw and name the polymer made from the monomer ethene:



Combustion of Hydrocarbons

Combustion means burning.

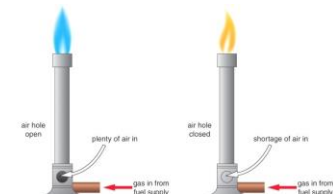
Complete combustion - when there is a good supply of **oxygen** for a fuel to burn.

Fuel + oxygen → carbon dioxide + water

Incomplete combustion - **not enough oxygen**

Products are **carbon monoxide** and water.

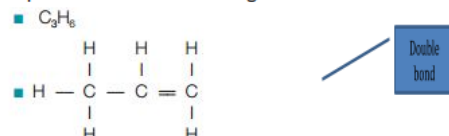
Carbon monoxide = poisonous gas



Alkenes

Structure and formulae of alkenes

- Alkenes have the general formula C_nH_{2n}
- they have at least one carbon-carbon double bond, making them unsaturated because they contain two fewer hydrogen atoms than the alkane with the same number of carbon atoms.
- The first 4 alkenes are ethene, propene, butene and pentene
- Unsaturated carbons can be represented in the following forms:



Reactions of alkenes

- Alkenes are hydrocarbons with the functional group $C=C$ (functional groups are usually responsible for the reactions of organic compounds)
- They react with oxygen in combustion reactions in the same way as other hydrocarbons, but they tend to burn in air with smoky flames because of *incomplete combustion* (meaning carbon or carbon monoxide is formed (CO))
- They also react with hydrogen, water and the halogens, by the addition of atoms across the carbon-carbon double bond so that the double bond becomes a single carbon-carbon bond :

ethene	C_2H_4	<pre> H H \ / C = C / \ H H </pre>	plus bromine (Br_2):
		<pre> H H H H Br - C - C - Br H H </pre>	
propene	C_3H_6	<pre> H H H H - C - C = C H H </pre>	plus water:
		<pre> H H H H H - C - C - C - OH H H H </pre>	propanol
butene	C_4H_8	<pre> H H H H H - C - C = C - C - H H H </pre>	plus hydrogen (H_2):
		<pre> H H H H H H H - C - C - C - C - H H H H H </pre>	butane

- In each reaction with hydrogen, water and halogens:
 - every reaction works the same for all alkenes, the table just shows examples.
 - the $C=C$ bond is broken to form a $C-C$ bond
 - the compound added splits into two groups and the two groups are added to the 2 different carbons in the $C=C$ bond (each group can be added to either carbon)
 - H_2 splits into 2 H 's, H_2O splits into a H and an OH , Br_2 splits into 2 Br 's (same for Cl_2 or I_2)

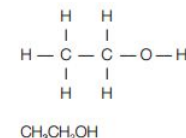
Alcohols

Alcohols

- Alcohols contain the functional group $-OH$
- The first 4 members of the series are methanol, ethanol, propanol and butanol:

methanol	CH_3OH	<pre> H H - C - O - H H </pre>
ethanol	C_2H_5OH	<pre> H H H - C - C - OH H H </pre>
propanol	C_3H_7OH	<pre> H H H H - C - C - C - OH H H H </pre>
butanol	C_4H_9OH	<pre> H H H H H - C - C - C - C - OH H H H H </pre>

- alcohols can be represented by:



Reactions of methanol, ethanol, propanol and butanol:

- They burn in air, which produces carbon dioxide and water
- They dissolve in water to form a neutral solution (has a pH of 7)
- They react with sodium to produce hydrogen and a salt (e.g. C_2H_5ONa - the H has been given off and Na has been added)
- They react with oxidising agents to form carboxylic acids

Uses of methanol, ethanol, propanol and butanol:

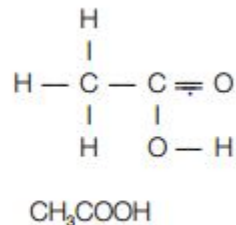
- methanol: chemical feedstock, in anti-freeze, to make biodiesel
- Ethanol: the main alcohol in alcoholic drinks, used as a solvent and fuel
- All 4: can be used as fuels

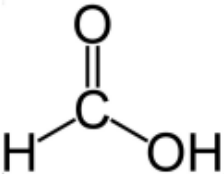
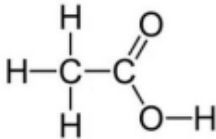
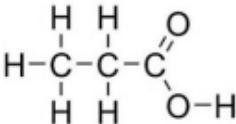
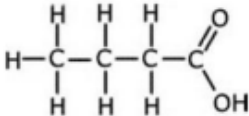
Producing ethanol:

- Ethanol can be produced by fermentation of sugar with yeast, using renewable sources.
- Conditions: about $35^\circ C$, anaerobic (without oxygen) and yeast enzyme catalyst
- Sugar $\rightarrow$ ethanol + carbon dioxide

Carboxylic acids

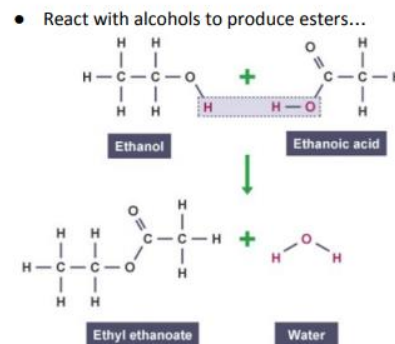
- Ethanoic acid is a member of the carboxylic acids, they have the functional group -COOH .
- First four members are: methanoic acid, ethanoic acid, propanoic acid and butanoic acid:



methanoic acid	CHOOH	
ethanoic acid	CH_3COOH	
propanoic acid	$\text{C}_2\text{H}_5\text{COOH}$	
butanoic acid	$\text{C}_3\text{H}_7\text{COOH}$	

Reactions of methanoic acid, ethanoic acid, propanoic acid and butanoic acid:

- They dissolve in water to produce acidic solutions (pH less than 7)
- They react with metal carbonates to produce carbon dioxide (turns limewater cloudy), a salt and water
- React with alcohols in the presence of an acid catalyst to produce esters

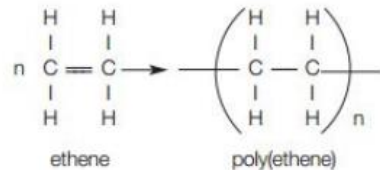


You don't need to know this equation and the only ester you need to know the name of is ethyl ethanoate

- They do not ionise completely in solutions, so do not release many H^+ ions, making carboxylic acids weak acids.
- This means they have a higher pH (less acidic) than solutions of strong acids of the same concentration.

Addition polymerisation

- Alkenes can be used to make polymers such as poly(ethene) and poly(propene) by addition polymerisation. In this reaction, many small molecules (monomers) join together to create very large molecules (polymers). For example:

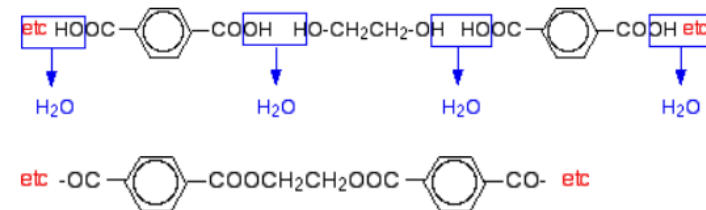


- The repeat unit has the same atoms as the monomer because no other molecule is formed in the reaction
- when you draw out a polymer make sure than you remember to draw the bonds coming off the C's out the brackets and the little 'n' (means there are large numbers of these molecules joined together)

Condensation polymerisation

- Involves monomers with two functional groups
 - When they react, they join together, usually losing small molecules such as water, and so the reactions are called condensation reactions
 - Simplest polymers are produced from two different monomers with two of the same functional groups on each monomer

E.g. this polyester has 1 monomer with 2 carboxylic acid functional groups and 1 monomer with 2 alcohol functional groups:



Amino acids

- They have two different functional groups in a molecule (an amine group and a carboxylic acid group)
- They react by condensation polymerisation to produce polypeptides (works the same as the polyester above just has different functional groups)
- Different amino acids can be combined in the same chain to produce proteins

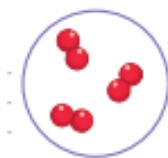
DNA (deoxyribonucleic acid) and other naturally occurring polymers

- DNA is a large molecule essential for life- it encodes genetic instructions for the development and functioning of living organisms and viruses
- Most molecules are two polymer chains, made from four different monomers called nucleotides, in the form of a double helix
- Other naturally occurring polymers important for life...
 - Proteins (monomer= amino acid), starch (monomer= glucose) and cellulose (monomer= glucose)

C8 – Chemical Analysis

Pure substances

Pure = single element or compound – not mixed with any other substance.



Testing to see if a substance is pure:

- Pure substances have specific melting and boiling points
- Compare your data to a library of known values.

E.g. Water has a boiling point of 100°C, if it is above or below this, it is not pure.

Formulations

Formulation = a mixture that is designed as a useful product.

- Components mixed carefully to get the required **properties**.

Examples of formulations:

- Fuels
- Cleaning agents
- Paints
- Medicines
- Alloys
- Fertilisers
- Food



Chromatography

- Technique used to separate mixtures of **soluble substances**.
- How soluble a substance is determines how far it travels across paper.

More soluble = travels further (higher up paper)

Mobile phase

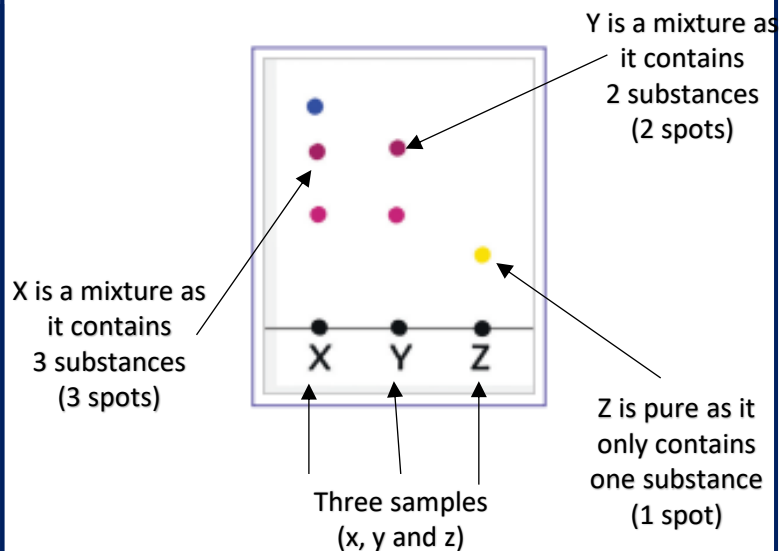
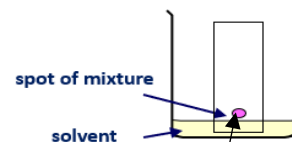
- **Solvent** is the mobile phase
- The substances dissolve in the solvent
- The solvent then moves through the stationary phase.

Stationary phase

- Does not move. The paper is the stationary phase.

Important – start line on paper must be drawn in **pencil** as pencil is **insoluble** and **will not run**

The spot and start line must be **above the solvent line** so the colours won't just wash into the solvent in the beaker.

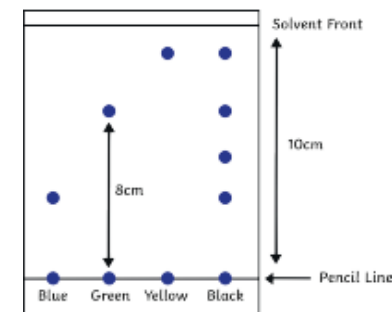


R_f Values

This is the ratio of the distance moved by a substance to the distance moved by the compound

$$R_f = \frac{\text{distance travelled by substance}}{\text{distance travelled by solvent}}$$

- Should always be between 0 and 1.
- Each substance has a unique R_f value.
- Can compare R_f values to a library of known substances
- Can identify unknown substances.



R_f value of green:

$$8\text{cm} / 10\text{cm} = 0.8$$

C8 – Chemical Analysis

Required Practical – Paper Chromatography

Aim: Investigate how paper chromatography can be used to separate and distinguish between coloured substances.

Method

- 1) Using a ruler, measure 1cm from bottom of chromatography paper and draw a line across the paper with a **pencil**.
- 2) Using a pipette, drop small spots of each ink onto pencil line (leave a gap so do not merge).
- 3) Pour solvent into a beaker, do not fill solvent above the pencil line on the paper.
- 4) Place chromatograph paper into beaker and allow solvent to move up the paper.
- 5) Remove paper just before solvent reaches top of the paper and leave to dry.
- 6) Calculate R_f values of all the spots using the equation below:

$$R_f = \frac{\text{distance travelled by substance}}{\text{distance travelled by solvent}}$$

Common questions

Q1) Why is a pencil used instead of a pen?

A1) Ink in the pen would move up the paper with the substances.

Q2) Why do you not fill the solvent above the line?

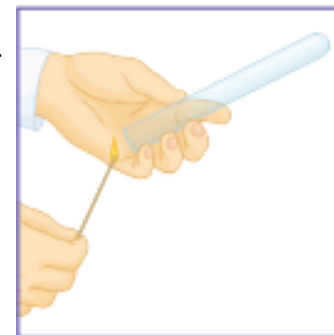
A2) Substances would wash off into the solvent instead of rising up the paper

Q3) Why might water not work as a solvent?

A3) Some substances are **insoluble** in water.

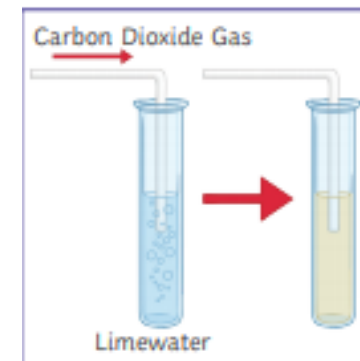
Identification of the Common Gases

Test for hydrogen – Place a **burning** splint at the opening of a test tube. If hydrogen gas is present, it will burn with a **squeaky-pop** sound.



Test for Oxygen – Place a **glowing** splint inside a test tube. The splint will **relight** in the presence of oxygen.

Test for Carbon Dioxide –Bubble the gas through the lime water – if the gas is carbon dioxide, the limewater turns **cloudy**.



Test for Chlorine – **Damp litmus paper** is held over the of gas. If the tube contains chlorine, the litmus paper becomes **bleached** and turns **white**.

Flame tests

- Flame tests can be used to identify metal ions.

Lithium	Crimson
Sodium	Yellow
Potassium	Lilac
Calcium	Orange- Red
Copper	Green

- However, if a sample containing a mixture of ions is used some flame colours can be masked (you won't be able to see them)



Metal hydroxides

- Aluminium, calcium and magnesium ions form a white precipitate with NaOH.
- Only aluminium's precipitate dissolves when excess NaOH is added.
- Copper(II) produces a blue precipitate
- Iron(II) produces a green precipitate
- Iron(III) produces a brown precipitate
- equations: e.g. $\text{Cu}^{2+} + 2\text{OH}^- \rightarrow \text{Cu}(\text{OH})_2$
 - you need as many OH^- ions as the charge on the metal ion
 - the Na from the NaOH and whatever the metal ion was bonded to will react to form a compound together: e.g. $\text{CuCl}_2 + 2\text{NaOH} \rightarrow \text{Cu}(\text{OH})_2 + 2\text{NaCl}$

Carbonates

- Carbonates react with dilute acids to create carbon dioxide.
- This gas can be bubbled through limewater, if the limewater goes cloudy, the gas is CO_2 .

Halides

- First add dilute nitric acid, followed by silver nitrate solution
- Chloride gives a white precipitate
- Bromide gives a cream precipitate
- Iodide gives a yellow precipitate
- (catswithbrainscanideallyyodel)

Sulfates

- First add dilute hydrochloric acid, followed by barium chloride solution
- A white precipitate will form when sulfate ions are in this solution

Instrumental methods

- Elements and compounds can be detected and identified using instrumental methods
 - These are: accurate, sensitive and rapid, making them advantageous compared to chemical tests

Flame emission spectroscopy

- Example of an instrumental method used to analyse metal ions in solutions
- Sample is put into a flame and the light given out is passed through a spectroscope
- Output is a line spectrum that can be analysed to identify the metal ions in the solution and measure their concentrations

C9 – Earth & Atmosphere

Early Atmosphere vs modern atmosphere:

Gas	Levels in earth's early atmosphere	Percentage in air today
Nitrogen	None	78
Oxygen	None	21
Others – CO ₂ and argon	Very High	1
Water vapour	Very high	Varies – but usually only around 1%
Ammonia	High	None



4.6
billion
years



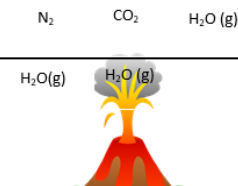
We think that the atmosphere on Earth was once like that of Mars or Venus is today

When Earth was formed it was so hot it was molten on the surface, and the atmosphere was full of toxic gases like methane and ammonia.

We cannot be sure about exactly what the Earth's early atmosphere as we have no evidence from so long ago

How did the atmosphere change?

1. Volcanoes released nitrogen, carbon dioxide and water vapour



2. The earth cooled and solidified



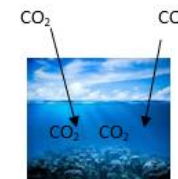
3. Water vapour in the atmosphere condensed and fell as rain



4. Oceans, lakes and rivers formed



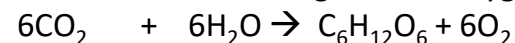
5. Carbon dioxide from the air dissolved in the oceans



6. Some of this reacted to form sedimentary rocks like limestone

7. Algae and then plants evolved, removing carbon dioxide from the air and produced oxygen by photosynthesis

Carbon dioxide + water → glucose + oxygen



8. Many early plants and marine organisms were buried and decayed underground, locking up carbon in fossil fuels like coal (plants) and oil (animals)



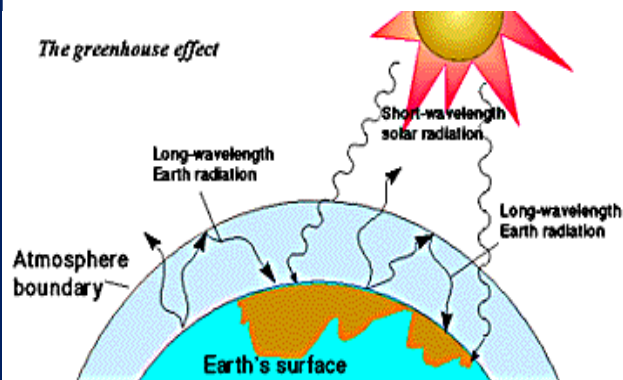
C9 – Earth & Atmosphere

The greenhouse effect

The greenhouse layer is a layer of gases in the atmosphere made of:

- carbon dioxide
- methane
- water vapour

The greenhouse effect



1. Short wavelength infrared radiation from the sun reaches Earth
2. Some energy is absorbed by the Earth
3. Longer wavelength IR is reflected by the Earth
4. Longer wavelength IR cannot get through the greenhouse layer as easily so some is trapped, warming the Earth

The thicker the layer of gases, the more heat is trapped

Global warming

The greenhouse layer is getting thicker, because:

- CO₂ released from fossil fuels to generate electricity
- CO₂ released from fossil fuels in vehicles
- Methane released from cattle
- Methane released from rotting landfill sites

Many scientists believe that human activities are causing the warming of the Earth.

Potential consequences:

- Melting ice caps
- Loss of habitats for animals and plants
- Damage to coral reefs caused by warmer oceans
- Changes to animal migration patterns
- Extreme weather patterns – more hurricanes, heat waves, droughts, snow and ice
- Difficulty growing crops so reduced food supply

Carbon footprint



The total amount of CO₂, CH₄ and water vapour released by of a product or service. E.g for a concert:

- electricity in performance
- Fossil fuels used by people travelling there
- Plastics used and disposed of in refreshments etc

Carbon footprints can be reduced by recycling, reducing energy use or eating vegetarian diets but this is hard to get people to do.

Pollutants :

Pollutant	Source	Effects
Carbon dioxide	Combustion	Global warming
Carbon monoxide	Incomplete combustion of fuels	Toxic gas, can be fatal
Sulfur dioxide	Traces of sulfur in coal react with oxygen when burned	Acid rain
Nitrogen oxides	Hot engines provide the energy for N ₂ to react with O ₂	Acid rain
particulates	Incomplete combustion	Global dimming, breathing problems

C10 – Using Resources

Earth's Resources

We use Earth's resources to provide **warmth, shelter, food and transport.**

E.g.:

- metals from the Earth's crust to build buildings and cars
- Timber and oil to burn for warmth
- Crop plants for food
- Products from crude oil to serve as fuels in cars, trains and planes

Finite resources – ones that will run out as they are being used much faster than they can be replaced, e.g. oil

Renewable resources – resources that will not run out, e.g. wood, wind etc.

Chemistry plays an important part in finding improvements or alternatives to current resources.

Natural	Improved or replaced by....
Wood for furniture	Plastic/polymers
Food crops	Fertilisers/artificially grown foods such as Quorn
Oil for fuel	Ethanol/hydrogen fuel cells
Rubber for tyres	Polymers

Life-Cycle Assessments (LCA)

- These assess the environmental impact of a product in these stages.

Stage 1 – extracting raw materials needed to make products.

- Energy cost and effect on habitats of extraction
- Are the raw materials finite/renewable?

Stage 2 – Manufacturing and packaging product

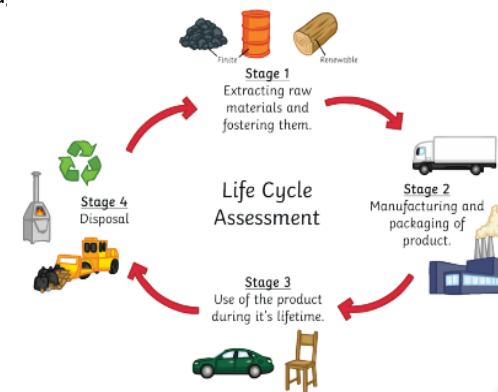
- How much energy and resources are needed?
- What waste products/pollution are released?
- Transportation of goods from factories to user need considering.

Stage 3 – Use of product during its lifetime

- E.g. a car has a significant impact as needs filled up with petrol which is a finite resource.

Stage 4 – Disposal at end of product's life.

- 1) Landfill – high environmental impact
- 2) Incineration – burning of product
- 3) Recycling – e.g. batteries contain metals that are harmful to environment – recycling means no new compounds need to be taken out of the ground.



Example LCA for plastic vs paper bags:

Stage of Life Cycle Assessment	Plastic Bag	Paper Bag
Stage 1 – raw material	Uses finite resource. Process of fractional distillation, cracking and polymerisation all require energy.	Made from trees/recycled paper. Making paper from trees required more energy than recycled paper. Less energy than plastic bags.
Stage 2 – Manufacture	Cheap to make	More expensive to make
Stage 3 – Use	Low environmental impact as can be re-used many times. Much stronger product.	Only be reused a limited number of times – short lifetime.
Stage 4 - disposal	Do not biodegrade easily in landfill.	Paper bags degrade easily in landfill sites.

- Different people have different opinions and so depends on who completes the LCA. Bias may be added.
- Some companies may only discuss some of environmental impacts of their product.
- Accurate numerical values should be used where possible – for example to show how much energy has been used.

C10 – Using Resources

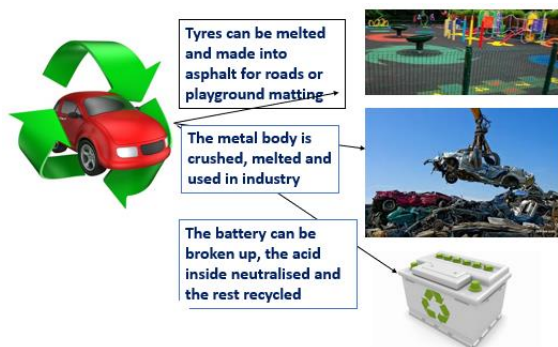
Reducing the use of resources

Metals, glass, ceramics, building materials and most plastics are produced from limited resources. The energy for the processes involved in making/extracting raw materials also comes from limited resources – e.g. oil. We can reduce the use of limited resources by reducing use, reusing materials and recycling materials at the end of their life.

Reduce, reuse, recycle.

E.g.

- Glass bottles can be reused.
- Metals can be melted down and recast and so recycled.
- Scrap steel can be added to extracted iron to reduce the amount of iron that has to be extracted in the blast furnace.



Evaluating methods to reduce, reuse, recycle

Advantages	Disadvantages
Fewer resources such as mines and quarries are needed to extract finite materials	Requires collection and transport of items – involving staff, vehicles and use of fuel
Crude oil does not need to be extracted – avoids high energy costs for fractional distillation etc.	Materials, such as metals, very often have to be separated from other materials first
Less greenhouse gases produced.	Some metals need melting before being reused – energy costs.
Less items in landfill	

Biological extraction techniques (HT only)

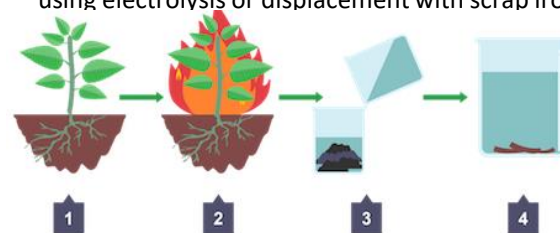
- Earth's supply of metal ores is limited.
- There are fewer sites that give lots of copper (high grade ore sites)
- New ways of extracting from low grade ore sites are:
 - Phytomining
 - Bioleaching

Disadvantage = **slow processes**

Advantage = reduce need for the traditional mining methods of digging, moving and disposing of large amounts of rock.

Phytomining (HT only)

- 1) Plants are grown on a low-grade ore
- 2) The plants absorb metal ions through their roots
- 3) The plants are harvested and burnt
- 4) Ash left behind contains metal compounds
- 5) Ash is dissolved in acid and copper is extracted using electrolysis or displacement with scrap iron.



Bioleaching (HT only)

- Uses **bacteria** to produce a solution called **leachate** – contains copper ions.
- The copper can be extracted by using iron to **displace** the copper from the leachate.
- Does not need high temperatures
- Produces **toxic substances** which can damage the environment.
- Iron is cheaper than copper – use of scrap iron is a cost-effective way to produce copper from leachate.
- Can also undergo **electrolysis** to produce copper.

C10 – Using Resources

Water

Potable Water

- Water is **essential** for life.
- **Potable water** is water that is safe to drink.
- Potable water is not pure as it contains some dissolved substances.

In the UK – rain water provides water with low levels of dissolved substances that collects in the ground and in lakes and rivers. This is fresh water.

Most potable water is produced by:

- 1) Choosing an appropriate source of fresh water
- 2) Passing the water through filter beds
- 3) Sterilising to kill bacteria

Sterilising agents used for potable water include:

- Chlorine
- Ozone
- Ultraviolet light



Desalination of Sea Water

- **Potable** water can be made from sea water through desalination.
- Required a lot of **energy** to **remove salt** in sea water.

Can be done by:

Distillation

- Sea water heated until it boils
- Steam is **condensed** to make potable water
- Requires a lot of **energy**

Reverse Osmosis

- Water put under **high pressure** and passed through **membrane** with tiny holes in.
- Holes allow water through but not salt/ions
- Very **expensive**
- Produces **large volumes** of waste water.

Waste Water Treatment

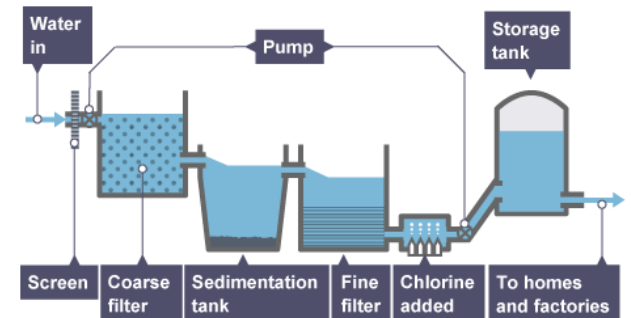
- Waste water needs to be treated before being released back into environment

Pollutants can be present in waste water including:

- Human waste contains harmful **bacteria** and nitrogen – can harm aquatic ecosystems.
- Industrial waste can contain **toxic** substances
- Agricultural waste water can contain **fertilisers** or **pesticides** – disrupt ecosystems.

Sewage treatment involves:

- 1) Screening and grit removal to remove large particles
- 2) Sedimentation – allows tiny particles to settle – produces sewage sludge and effluent (liquid that remains on the top)
- 3) Sewage sludge is digested anaerobically by specific bacteria
- 4) Effluent is treated with aerobic bacteria to reduce volume of solid waste.



C10 – Using Resources – Required Practical – Analysis and purification of water

Analysing the pH of Water Samples

- Test pH of each water sample using pH probe or universal indicator.
- Compare to pH chart if using universal indicator

Analysis the Mass of Dissolved Solids

1) Measure out 50 cm³ of water sample using measuring cylinder.

2) Take the mass of evaporating basin using top pan balance.

3) Heat the sample in the evaporating basin gently until all liquid evaporates.

4) Let the evaporating basin cool

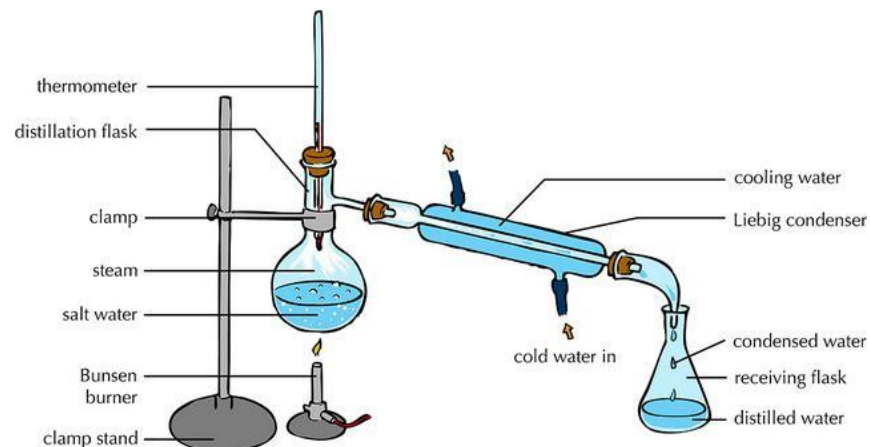
5) Re-take the mass of the evaporating basin.

6) Calculate the mass of the solid left behind by doing: final mass – initial mass.

7) Repeat with different water samples (e.g. rainwater, salt water, spring water)



Distillation of water Sample



- 1) Set up apparatus as shown in picture with the sample of water in the round bottom flask.
- 2) Heat water sample until it boils gently.
- 3) Water vapour enters the tube at the side (condenser)
- 4) There is cold water surrounding the tube
- 5) The water vapour cools and condenses and collects in the flask.
- 6) The water collected should be **pure**.

Corrosion and its prevention

- Corrosion = destruction of materials by chemical reactions with substances in the environment
 - E.g. rusting
 - Both air and water are necessary for iron to rust
- Corrosion can be prevented by applying a coating that acts as a barrier, such as greasing, painting or electroplating
 - Aluminium has an oxide coating that protects the metal from further corrosion
 - Some coatings are reactive and contain a more reactive metal to provide sacrificial protection
 - E.g. zinc is used to galvanise iron
 - sacrificial protection works by the more reactive metal donating electrons to any ions of the other metal that may have formed so they don't corrode

Alloys as useful materials

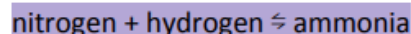
- Most metals in everyday uses are alloys. Pure copper, gold, iron and aluminium are all too soft for everyday uses and so are mixed with small amounts of similar metals to make them harder for everyday use.
 - Gold in jewellery is usually an alloy with silver, copper and zinc and the gold purity is measured in carats
 - Bronze is an alloy of copper and tin - used in electrical connectors
 - Brass is an alloy of copper and zinc - used for tools
- Steels are alloys since they used mixtures of carbon and iron
 - Some steels contain other metals. Alloys can be designed to specific uses.
 - Low-carbon steels are easily shaped - used for sheeting (malleable)
 - High carbon steels are hard - used for cutting tools
 - Stainless steels (containing chromium and nickel) are resistant to corrosion - used for cutlery
- Aluminium alloys are low density - used for aircraft

Ceramics, polymers and composites

- Soda-lime glass: made by heating a mixture of sand, sodium carbonate and limestone (most commonly used glass)
- Borosilicate glass: made from sand and boron trioxide, melts at higher temperatures than soda-lime glass
- Clay ceramics, including pottery and bricks, are made by shaping wet clay and then heating in a furnace
- Properties of polymers depend on what monomers they are made from and the conditions under which they are made.
 - Low density polyethene and high density polyethene, are both made under different reaction conditions using different catalysts.
 - Low density polyethene has weaker forces of attraction as the chains are further apart, meaning it has a low melting point and is soft
 - HD polyethene has higher forces of attraction, as the chains are closer together, giving it a higher melting point
- Thermosoftening polymers are made of individual, tangled polymer chains which are easily separated and are melted by heat
 - 1) There are weak intermolecular forces between the chains
 - 2) The chains are easy to separate
 - 3) At lower temperatures
 - 4) Less heat energy is needed to break the chains
- Thermosetting polymers consist of polymer chains, which cross links, so that they do not melt when heated.
- Most composites are made of two materials, a matrix or binder surrounding and binding together fibres or fragments of the other material, which is called the reinforcement
 - examples of composites: carbon fibre and fibreglass (both are very hard but very brittle)

The Haber process

- Used to manufacture ammonia, which is used to produce nitrogen-based fertilisers
- The raw materials are nitrogen and hydrogen.
- Nitrogen is obtained from the air and hydrogen may be obtained from natural gas or other sources.
- The purified gases are passed over a catalyst of iron at a high temperature (about 450 °C) and a high pressure (about 200 atmospheres).
- Some of the hydrogen and nitrogen reacts to form ammonia.
- The reaction is reversible so ammonia breaks down again into nitrogen and hydrogen.



- On cooling, the ammonia liquefies and is removed.
- The remaining nitrogen and hydrogen are recycled.
- Reaction conditions and compromise:
 - o the Haber Process is in dynamic equilibrium- the forward and backward reactions keep going once equilibrium is reached
 - o the chemical equation is $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$, so there are less moles of gas on the product side (2 compared to 4), this means you would increase pressure to move equilibrium to the right so more ammonia is produced
 - o the forwards reaction is exothermic, so a low temperature would favour the forwards reaction and mean more ammonia was produced
 - o HOWEVER the actual conditions used are not low temperature and very high pressure because a low temperature leads to a reaction rate that is too slow and a very high pressure requires too much energy

Production and uses of NPK fertilisers

- Compounds of nitrogen, phosphorus and potassium are used as fertilisers to improve agricultural productivity
- NPK fertilisers contain compounds of all three elements
- Industrial production of NPK fertilisers can be achieved using a variety of raw materials in several integrated processes
- NPK fertilisers are formulations of various salts containing appropriate percentages of the elements
- Ammonia can be used to manufacture ammonium salts and nitric acid
- Potassium chloride, potassium sulfate and phosphate rock are obtained by mining, but phosphate rock cannot be used directly as a fertiliser
 - o Phosphate rock is treated with nitric acid or sulfuric acid to produce soluble salts that can be used as NPK fertilisers

Use of Phosphate Rock

- The Earth's crust is rich in minerals that serve as essential raw materials for creating fertilizers.
- Phosphate rocks are a source of potassium chloride and potassium sulfate, which provide potassium.
- Since the rock is insoluble in water, it is generally reacted with acid to create useful water-soluble compounds.
 - o Reacting the rock with nitric acid yields phosphoric acid and calcium nitrate. The phosphoric acid is neutralised with ammonia, producing ammonium phosphate.
 - o Reacting the rock with sulfuric acid produces a mixture of calcium phosphate and calcium sulfate, known as single superphosphate.
 - o Reacting the rock with phosphoric acid results in calcium phosphate, referred to as triple superphosphate.