The ChemisTree of KNOWLEDGE

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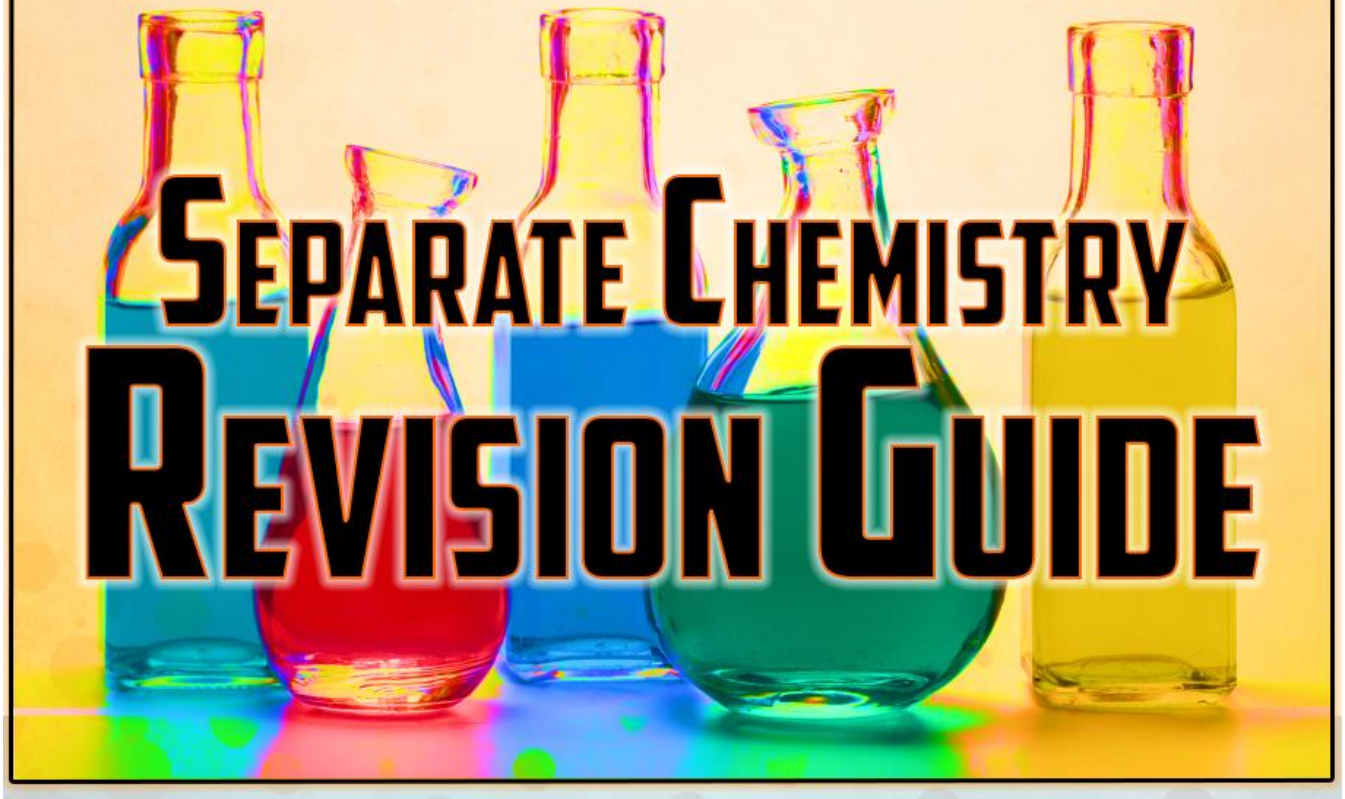
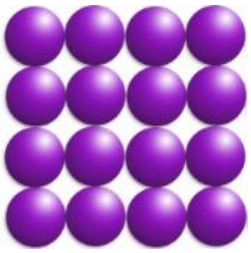
Mr Barnes Teaches Chemistry



Separate Chemistry Revision Guide

Section 1: The Atom:

A. Changing ideas of the Atom

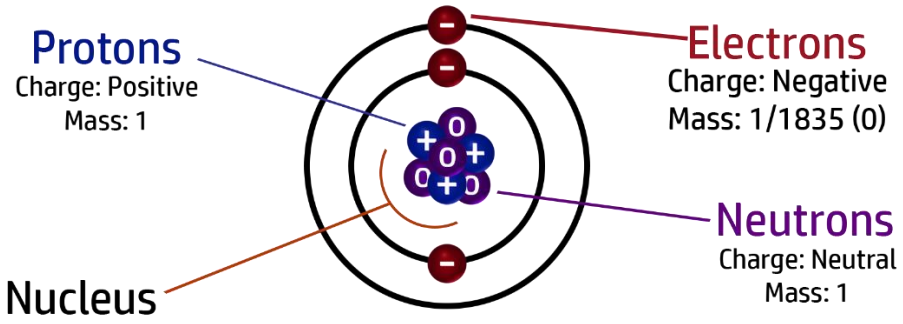


1. **Dalton** described atoms as **indestructible spheres** that could not be broken down

2. **JJ Thompson** measured the mass of rays in a cathode tube. They were lighter than atoms – proving that **electrons** existed.

3. **Rutherford** fired *20,000 alpha particles* into gold foil. Most went through, some refracted and *a few repelled*. This proved the **nucleus was positive and tiny**.

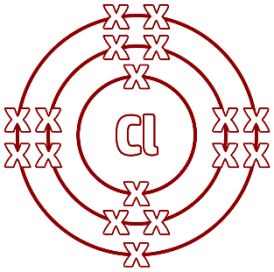
B. Structure of the Atom



C. Protons, Neutrons and Electrons:

Element:	Protons:	Neutrons:	Electrons:
7 Li Lithium 3	Bottom number = 3	Big-little number = 4 (7-3)	Bottom number = 3
27 Al Aluminium 13	Bottom number = 13	Big-little number = 14 (27-13)	Bottom number = 13

D. Drawing the Electronic Configuration



When drawing the electronic configuration, remember these key things:

1. The electrons fill up from the inside (smallest shell) first
2. The number of electrons is the atomic number (smallest number on the Periodic Table)
3. The first shell can only hold 2 electrons
4. All other shells can hold 8 electrons

- If you are asked to write the electronic configuration, it's as simple as stating how many electrons are in each shell.
- The written electronic configuration for aluminium is 2.8.3

F. The link between electronic configuration, group and period:

Chlorine (above) is in group 7 because it has 7 electrons in the outer shell.

Chlorine is in period 3 because it has 3 shells.

Section 2: The Periodic Table:

A. Dmitri Mendeleev's Periodic Table

- **Similarities**: Both our Periodic table and Mendeleev's were in groups based on **chemical properties**.
- Differences: Mendeleev arranged his in order of **atomic weight** and had **gaps** because there were undiscovered elements. Ours is in order of **atomic number** and **doesn't have gaps**.

B. Isotopes:

An isotope is an atom that has the same number of protons, but different number of neutrons.

The main example you need to know is chlorine:

Protons:	17	17
Neutrons:	18	20
Electrons:	17	17

In the above example, there are two chlorine atoms – one with a mass number of 37 and one with a mass number of 35. As the atomic number hasn't changed, the number of protons must be the same. Therefore, the only thing that can change is the neutrons.

As electrons can change (see ionic bonding), we do not include this in the definition of an isotope.

C. Calculating the Relative Atomic Mass of an isotope. (Higher)

In an exam, you may be given the relative abundance (amount)of an isotope and asked to calculate the Relative Atomic Mass of an isotope.

The Relative Atomic Mass is the **average mass of all the isotopes.**

To calculate this, use the following equation:

$$\text{Relative Atomic Mass of an Isotope} = \frac{\text{mass x \% of isotope 1}}{100} + \frac{\text{mass x \% of isotope 2}}{100} \quad \text{etc...}$$

For example, if Chlorine has two isotopes – 35-Cl, with an abundance of 75% and 37-Cl with an abundance of 25%, the calculation would look like this:

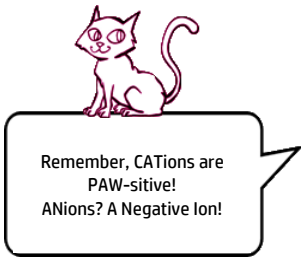
$$\text{Relative Atomic Mass of an Isotope} = \frac{35 \times 75}{100} + \frac{37 \times 25}{100}$$

$$= 26.25 + 9.25 = 35.5$$

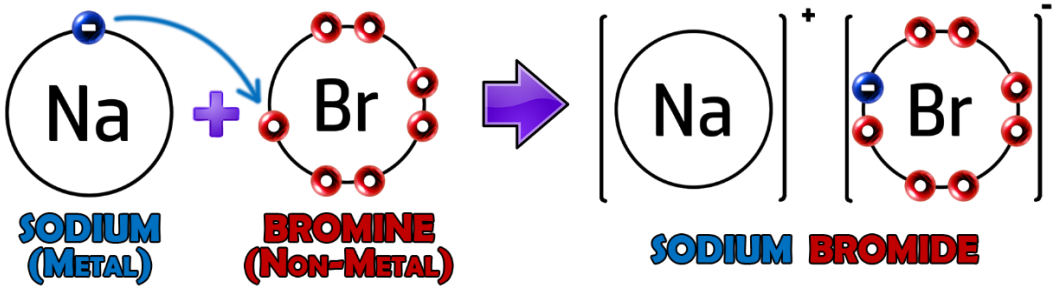
You can also work out which isotope is more abundant from the Relative Atomic Mass. If the Relative Atomic Mass of Boron is 10.8 and it has two isotopes, 10-B and 11-B, Boron-11 is more abundant as it the average mass is closest to 11.

Section 3: Ionic Bonding:

A. What is an ion:



- An **ionic substance** is a compound formed when a **metal loses electrons** and **non-metals gain electrons** to get a full outer shell. They do this by **transferring electrons** from the metal to the non-metal.



- All elements want to get a full outer shell to become stable. In the above diagram:
  - Sodium has **one electron** in its outer shell. It is easier to lose one electron (than gain 7 more), so it will transfer its electron to bromine.
  - Bromine has **seven electrons** in its outer shell. It is easier to **gain one electron** (than lose 7), so it will accept an electron from sodium.

B. Finding the Formula:

Rules:	Example: Calcium Chloride		
Step 1. Work out the **valency** – the number of electrons the atom loses/gains	Calcium = group 2 = wants to lose $2e^-$: Ca^2	Chlorine = group 7 = wants to gain $1e^-$: Cl^1	
Step 2. Swap the numbers and put them down below (ignoring 1's)	$Ca^2 + Cl^1 \rightarrow$	$Ca_1Cl_2 \rightarrow$	$CaCl_2$
Step 3: Simplify if possible. If both numbers are the same, they cancel out.	$Mg^2 + O^2 \rightarrow$	$Mg_2O_2 \rightarrow$	MgO
Step 4: If you multiply a **polyatomic** ion, like NO_3, put a **bracket** around it.	$MgNO_{32}$ ✘	$Mg(NO_3)_2$ ✓	

C. What are the properties of ionic compounds?

Why do ionic compounds have high melting points?

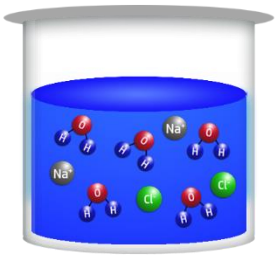
- There is a strong **electrostatic** force of attraction between the ions.
- **Lots of energy** is needed to **break this strong electrostatic attraction**.

Why can't Ionic Compounds conduct when SOLID?

- When solid, there is a strong force of attraction between the ions.
- The ions are **not free to move** and cannot carry a charge.

Why can ionic Compounds conduct when LIQUID/MOLTEN/AQUEOUS?

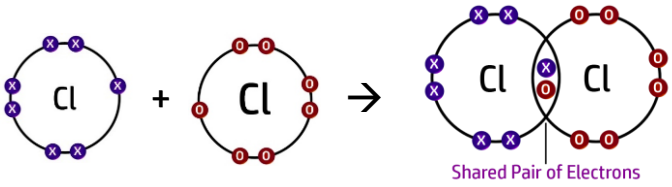
- When molten (melted) or dissolved, the **ions are free to move**.
- The ions can now carry a charge and conduct electricity.



Section 4: Simple Covalent Bonding:

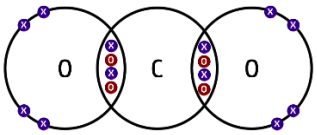
A. What is a covalent bond?

A covalent bond is a **shared pair of electrons** between non-metals

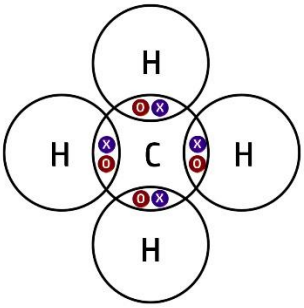


A molecule is a substance that involves **covalent bonding** and only contains a few atoms. They are tiny, usually around $1x10^{-10}$m small.

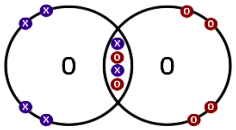
Carbon Dioxide, CO_2



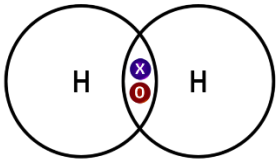
Methane, CH_4



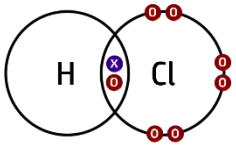
Oxygen, O_2



Hydrogen, H_2



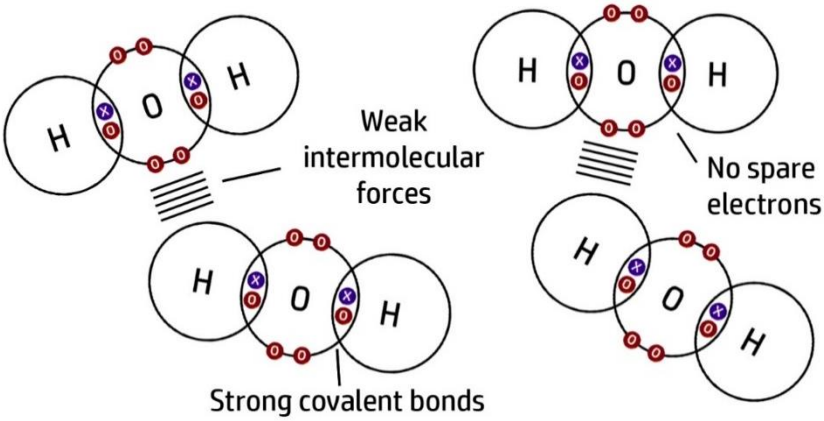
Hydrogen Chloride, HCl



B. The properties of simple (covalent) molecules:

There are two main properties you need to know about simple covalent compounds:

- **They cannot conduct electricity** – The electrons are not free to move, so cannot carry a charge.
- **They have low melting point and boiling points** – Not much energy is needed to break the weak intermolecular forces
- **They are soluble**.



C. Giant Covalent Structures:

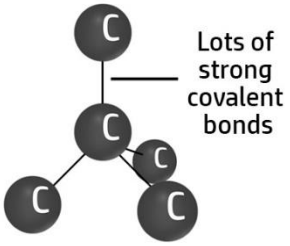
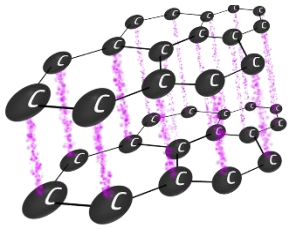
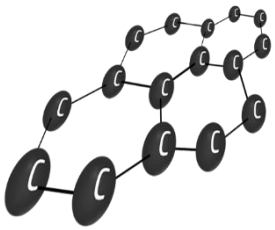
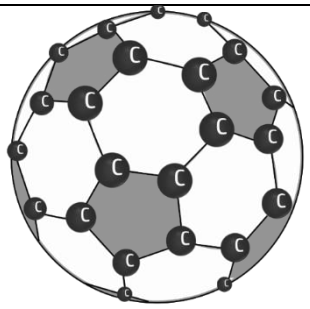
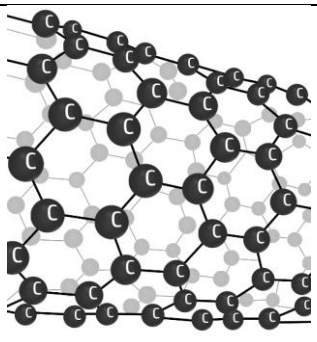
Giant covalent substances are bonded the same way as seen in simple covalent molecules. They still have **shared pairs of electrons** and involve only **non-metals.**

The only difference is the number of atoms and bonds involved – there are a lot more. This gives giant covalent structures the following properties:

Section 5: Giant Covalent Bonding:

E. Allotropes of Carbon:

The following structures are all classed as **allotropes** of carbon – this means they are all made up of carbon but have different structures. Some are giant covalent, and some are simple molecular.

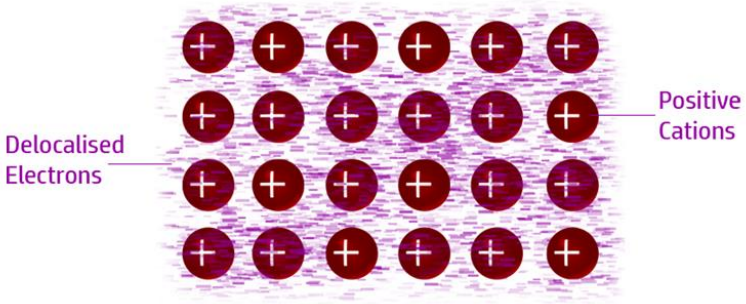
Structure:		Use/Properties
Diamond		**Bonding: Giant Covalent** **Structure:** Each carbon atom forms 4 strong covalent bonds **Use:** Cutting tools **Explanation:** Lots of energy is needed to break the strong covalent bonds – this makes diamond strong/hard
Graphite		**Bonding: Giant Covalent** **Structure:** Each carbon atom forms 3 strong covalent bonds, forming layers and delocalised electrons. **Uses:** Lubricant / Electrodes **Lubricant Explanation:** Layers can slide past each other reducing friction. **Electrodes Explanation:** Delocalised electron is free to move and can carry a charge.
Graphene		**Bonding: Giant Covalent** **Structure:** Each carbon atom forms 3 strong covalent bonds and is only one layer thick making it light and strong. **Use:** Electronics **Explanation:** Delocalised electron is free to move and can carry a charge.
Buckminsterfullerene		**Bonding: Simple Covalent** **Structure:** A ball made up of 60 carbon atoms – each carbon is bonded covalently to three other carbon atoms. **Use:** Drug Delivery **Explanation 1:** Low melting point – weak intermolecular forces easy to break. **Explanation 2:** Conducts - Delocalised electron is free to move and can carry a charge.
Nanotubes		**Bonding: Giant Covalent** **Structure:** A layer of graphene rolled into a cylinder. Long and thin. **Uses:** Electronics & Tennis Racquets **Tennis Racquet Explanation:** Lots of strong covalent bonds. Lots of energy to break these bonds. **Electronics Explanation:** Delocalised electron is free to move and can carry a charge.

Section 6: Metallic Bonding and Molecular Models:

A. What are the properties of Metals and Non-Metals?

Properties of Metals:	Properties of Non-Metals:
❋ Good conductors of heat	❋ Poor conductors of heat (insulator)
❋ Good conductors of electricity	❋ Poor conductors of electricity (insulator)
❋ Strong	❋ Brittle (Break easily)
❋ Malleable (Can be hammered into shape)	
❋ Ductile (Can be stretched into wires)	
❋ High Melting Points	❋ Low Melting Points
❋ High Boiling Points	❋ Low Boiling Points
❋ Sonorous (Makes a ringing sound)	❋ Dull sound when hit
❋ Mainly solids	❋ Solids, Liquids and Gases
❋ High Density	❋ Low Density
❋ Shiny when polished	❋ Dull

B. Explaining the properties of Metals

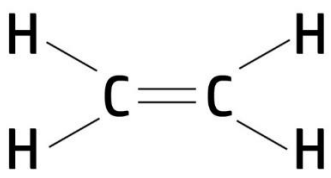


Property:	Explanation:
High Melting Points and Boiling Points	❋ **Lots of energy** needed to break the **strong electrostatic attraction** between *cation* and *delocalised electrons*.
Can conduct electricity	❋ **Delocalised electrons** can **flow**/move and carry a charge
Malleable	❋ The **layers can slide** past each other, but it **doesn't break the strong electrostatic attraction**.

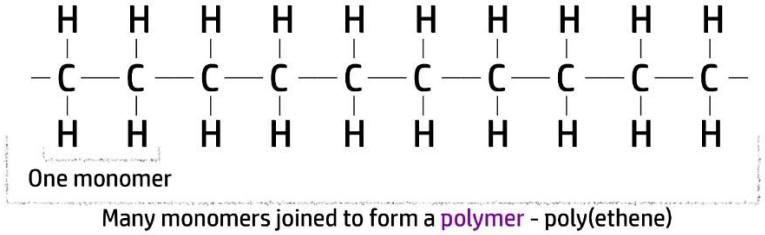
C. Polymers:

A polymer is a **long chain** of smaller molecules called **monomers**. These are examples of simple covalent compounds and have low melting points and cannot conduct electricity.

An example of this is **ethene** – which has the formula C_2H_4:



Monomer: Ethene



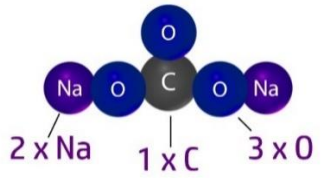
Polymer: Poly(ethene)

Section 7: Calculations involving Masses

A. Relative Formula Mass

The **relative formula mass** is simply the mass of all the atoms in a compound added together.

Example: Sodium carbonate, Na_2CO_3



A_r: Na = 23; C = 12; O = 16

Here the little number applies to **only** the element before it – so I have 2xNa (sodium), 1xC (carbon) and 3xO (oxygen)

2 x Na = 2 x 23 = **46**

1 x C = 1 x 12 = **12**

3 x O = 3 x 16 = **48**

Relative Formula Mass = 46 + 12 + 48 = **106**

B. Empirical Formula:

The Empirical formula is the simplest ratio of atoms in a compound:

Here, all three atoms can be simplified by dividing by 6. → $C_6H_{12}O_6$ Molecular Formula — CH_2O Empirical Formula ← This is the simplest ratio – the empirical formula.

You can also work backwards. For example, if asked to *work out the molecular formula of CH_2O when the M_r is 120*

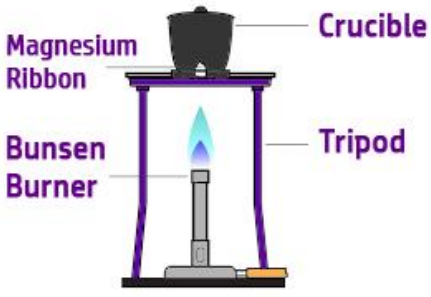
- CH_2O has a formula mass of 30. 30 fits into 120 four times, so $CH_2O \times 4 = C_4H_8O_4$.

C. Calculating the Empirical Formula from reacting masses

Example: 3.2 g of sulfur reacts with 3.2g of oxygen to produce 6.4 g of sulfur oxide. What is the formula of the oxide?

	S	O
Step 1: Divide the mass (g) by the atomic mass	3.2g / 32 = **0.1**	3.2g / 16 = **0.2**
Step 2: Divide both of these numbers by the **smallest** of the two numbers.	0.1 / 0.1 = **1**	0.2/0.1 = **2**
Step 3: Put these numbers into the formula	**SO_2**	

D. Investigating the Empirical Formula



If asked to describe a practical to work out the empirical formula of an oxide, do the following

1. Weigh a crucible and lid.
2. Weigh a strip of Magnesium – record the weight.
3. Clean the Magnesium strip to remove the oxide layer
4. Heat the magnesium strongly in a crucible. Keep lifting the lid of the crucible to replenish the oxygen.
5. Once the magnesium has fully reacted and is a white powder, reweigh the crucible, lid and magnesium oxide.
6. Subtract the mass of the crucible to leave you with the mass of magnesium oxide.
7. Subtract the mass of magnesium used from the mass of magnesium oxide to give you the mass of oxygen.
8. Carry out the calculation as shown above.

E. Calculating the Concentration in gdm^{-3}

To calculate the concentration of a solution in grams per decimetre cubed (gdm^{-3}), you need two things: - **the mass in grams, and the volume in dm^3 (cm^3 / 1000)**

Example: 20g of sodium hydroxide dissolves in 750cm^3 of water. Calculate the concentration in gdm^{-3}

- Volume = 750 / 1000 = 0.75dm^3
- Concentration = mass / volume = 20g / 0.75dm^3 = **= 26.7gdm^{-3}**

F. Conservation of Mass

The conservation of mass states that the total mass of reactants is **always the same** as the total mass of products. If the mass goes down, it is because a gas has been produced. Solution? **Add a lid**! You can use this to work out the maximum mass produced in a reaction:

Example: Work out how many grams of sodium chloride are produced when 5.3g of sodium carbonate reacts with dilute hydrochloric acid.

$$\mathbf{Na_2CO_3} + 2HCl \longrightarrow \mathbf{2NaCl} + H_2O + CO_2$$

	$\mathbf{Na_2CO_3}$	**2NaCl**
Step 1: Calculate the formula mass, M_r, for each substance involved:	46 + 12 + 48 **= 106**	2x(23+35.5) **= 117**
Step 2: Calculate the moles – mass / M_r:	5.3 / 106 = **0.05**	
Step 3: Work out the mass – moles x M_r:		0.05 x 117 **= 5.85g**

G. Moles and Avogadro's number (H)

- To calculate the moles, divide the mass in grams by the formula/atomic mass.
- To calculate the number of particles, multiply the moles by Avogadro's constant – 6.02x10^{23}

H. Limiting Reactant (H) – The reactant that runs out first

Example: 1.50g of ammonium chloride reacts with 4.0g of calcium hydroxide. To form ammonia. Work out the limiting reactant. $2NH_4Cl + Ca(OH)_2 \rightarrow 2NH_3 + CaCl_2 + 2H_2O$

	$\mathbf{2NH_4Cl}$ +	$\mathbf{Ca(OH)_2} \rightarrow$
Step 1: Calculate the formula mass (M_r) for each substance:	2(14+4+35.5) **= 107**	40+2(16+1) **= 74**
Step 2: Calculate the moles – mass / M_r	1.5 / 107 = **0.0140**	4 / 74 = **0.0540**
Step 3: The smallest number is the limiting reactant.	Limiting Reactant	

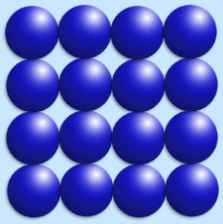
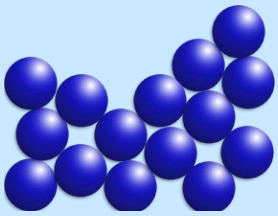
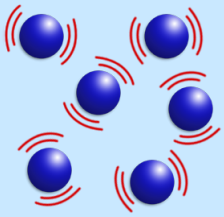
I. Stoichiometry (H) – Balancing Masses using Mole calculations

Example 1: 18.25g of hydrochloric acid, HCl, reacts with 10g of Magnesium oxide to form 23.75g of magnesium chloride, $MgCl_2$ and 4.5g of water. Deduce the balanced equation for the reaction.

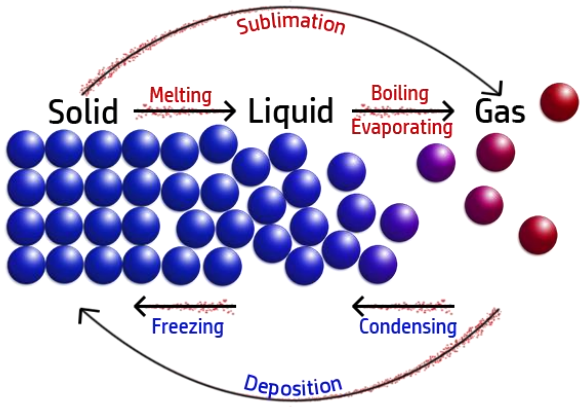
Step 1: Write the unbalanced equation:	HCl +	MgO →	$MgCl_2$ +	H_2O
Step 2: Calculate the Formula Mass (M_r) for each substance	1 + 35.5 **=36.5**	24 + 16 **= 40**	24 + 71 **=95**	2 + 16 **=18**
Step 3: Calculate the moles (Mass / M_r)	18.25 / 36.5 **= 0.5**	10 / 40 **= 0.25**	23.75 / 95 **= 0.25**	4.5/18 = **0.25**
Step 4: Divide both sides by the smaller number	0.50 / 0.25 **= 2**	0.25 / 0.25 **= 1**	0.25 / 0.25 **= 1**	0.25 / 0.25 **= 1**
Step 5: Put these values into the balanced equation.	**2HCl +**	**MgO →**	**$MgCl_2$ +**	**H_2O**

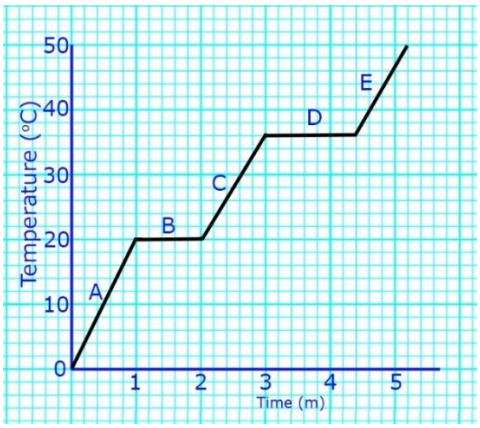
Section 8: States of Matter and Mixtures:

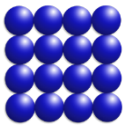
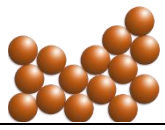
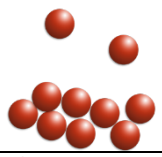
A. The Particle Model:

Solid:	Liquid:	Gas:
		
Movement: **Vibrate** about a fixed position.	**Movement:** Free to **move** / **flow** about each other.	**Movement:** Moving **fast** in **all directions**.
Arrangement: **Regular** pattern (in rows), and **touching**.	**Arrangement:** **Random** arrangement, but still **touching**.	**Arrangement:** **Random** arrangement and **not touching**.

B. State Changes and State Change Graphs:





Part of Graph:	Particle Model:	Explanation:
A		❄ Particles are vibrating about a fixed position. ❄ As the temperature increases, the particles will vibrate more.
B		❄ The temperature remains constant. ❄ This is where **MELTING** is occurring. ❄ The energy is now being used to **break some of the weak** forces between the molecules (**intermolecular forces)**
C		❄ At this point, the substance is completely melted, in a random pattern. ❄ As the temperature increases, more particles are starting to vibrate more.
D		❄ The temperature remains constant. ❄ This is where **EVAPORATING** is occurring. ❄ The energy is now being used to **completely break** the weak forces between the molecules (**intermolecular forces)**
E		❄ The particles are now not touching and moving fast in all directions. ❄ As the temperature increases, the particles move faster and faster.

Section 9: Separation Techniques:

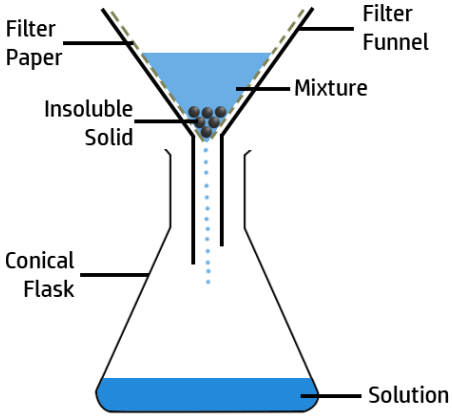
A. Pure or Mixture?

A pure substance will have only **one compound**/element present and will have only **one melting point.**

A mixture will have **more than one compound** (not bonded) and will have **multiple melting points**.

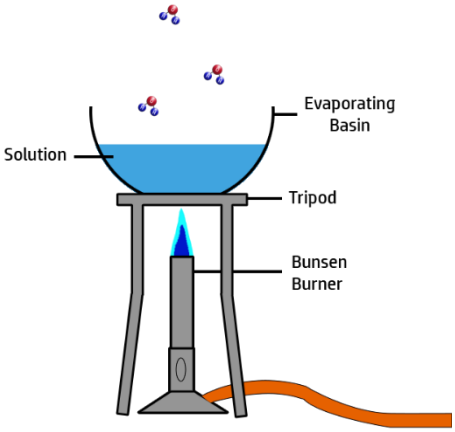
B. Filtration and crystallisation:

Filtration:



- Filtration is a technique used to separate a liquid from an **insoluble solid** – a solid that does not dissolve.
- Use filter paper and the solid will be left in the paper.

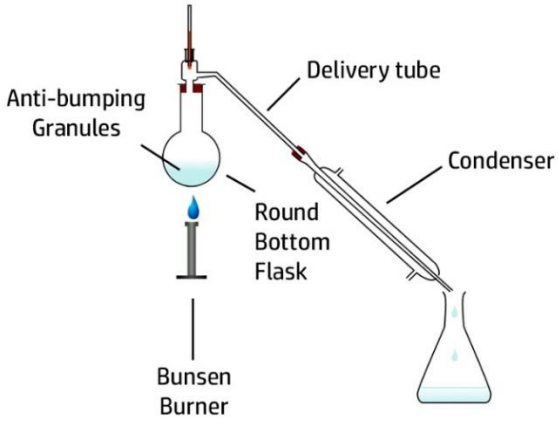
Crystallisation:



- Crystallisation is a technique used to separate a liquid from a **soluble solid** – a solid that does dissolve.
- **Heat** the solution to evaporate ½ of the water. Leave to **cool** and leave to **dry** to get crystals.

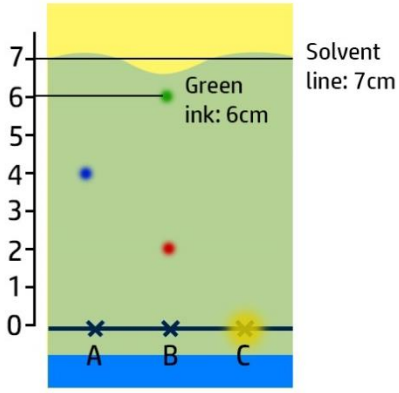
C. Distillation and Chromatography:

Simple Distillation:



- Simple distillation is a technique used to separate two different **liquids** based on their **boiling** points.
- Example: Separating water (boils at 100°C) and ethanol (boils at 78°C).
  - Heat to the lowest boiling point.
  - Once it reaches 78°C, ethanol will evaporate and move into the **condenser**.
  - The condenser is surrounded by cold water, which cools the gas down – turning it back into a **liquid.**

Chromatography:



- Draw a **line** in **pencil** (**insoluble**) and add your **ink** to the cross. Add to water and let the ink move up the chromatography paper. Remove when it reaches the top.
- One dot? You have a **pure** ink.
- More than one dot? You have a **mixture**.
- The further it moves the *more soluble*.
- To calculate the retention factor, divide the **distance the ink moves** by the **distance the solvent moves**.

Section 10: Acids, Bases and Indicators:

A. Ions and the effect of indicators on acids and bases:

Substance:	pH range:	Ions:	Colour in phenolphthalein:	Colour in Methyl orange:	Colour in Litmus:
Acid	1-6	H^+	Colourless	Red	Red
Alkali/Base	8-12	OH^-	Pink	Yellow	Blue
Neutral	7	None	Colourless	Orange	Purple

- An alkali is any base that is soluble. All alkalis are bases, but **not** all bases are alkalis.

B. Ions and pH:

- Every time the pH decreases by one, the concentration of H^+ ions **increases** by 10 times.
- If an acid is diluted from pH 4 to pH 2, it is 10x10 = 100 times more concentrated.

C. Dilute, Concentrated, Strong and Weak Acids:

- **Dilute:** The more dilute an acid is, the less H^+ ions there are per dm^3
- **Concentrated:** The more concentrated an acid is, the more H^+ ions there are per dm^3
- **Strong Acids:** Strong acids dissociate / ionise completely – producing more H^+ ions.
- **Weak Acids:** Weak acids do not fully ionise – only a small amount dissociate – fewer H^+ ions. pH = 2-6

D. What happens to the ions during neutralisation?

When any neutralisation reaction occurs, the same basic reaction occurs. Hydrogen ions react with hydroxide ions to form water:

$$H^+ + OH^- \rightarrow H_2O$$

E. Testing for Gases:

Hydrogen: Hydrogen is a flammable gas. If you take a **lit splint** and add it to a test tube containing hydrogen, you will hear a **squeaky pop**.

Carbon Dioxide: Carbon dioxide turns limewater cloudy. **Bubble** the gas through **limewater** and if it goes **cloudy/milky**, carbon dioxide is present.

F. Naming Salts:

a. If you have hydrochloric acid, HCl, you get a chloride salt
b. If you have nitric acid, HNO_3, you get a nitrate salt
c. If you have sulphuric acid, H_2SO_4, you get a sulphate salt.
d. A metal on its own will produce hydrogen gas, H_2
e. A metal oxide/hydroxide will produce water, H_2O, and
f. A carbonate will produce water **and** carbon dioxide, $H_2O + CO_2$

G. What is seen when acids and bases react together?

Look at the chemical equation below. What would you be able to **see** during this reaction?

$$CuCO_{3\,(s)} + HNO_{3\,(aq)} \rightarrow Mg(NO_3)_{2\,(aq)} + H_2O_{\,(l)} + CO_{2\,(g)}$$

The state symbols show that a solid has disappeared and that a gas has formed. Therefore:

1. The solid has disappeared/dissolved (1)
2. There will be bubbling/fizzing/effervescence proving there is a gas (1)

Don't say a gas is formed – you can't **see** that – but you can see bubbles, which **prove** there is a gas.

Section 11: Investigating pH:

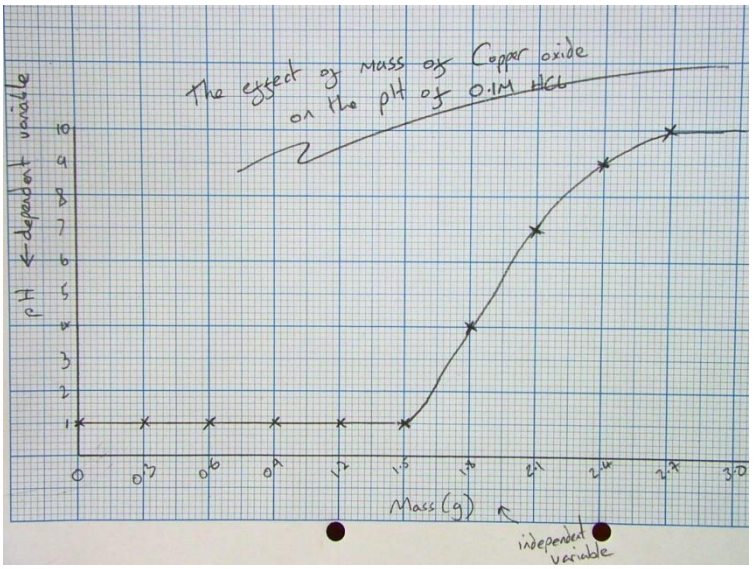
A. Core Practical: Investigating the change in pH when calcium hydroxide is added to an acid:

Calcium hydroxide	+	Hydrochloric Acid	→	Calcium chloride	+	Water
$Ca(OH)_2$	+	$2HCl$	→	$CaCl_2$	+	$2H_2O$

This is one of the big practicals that you need to know for the exam. You could be asked how to carry it out; to analyse the results; to explain the results or to evaluate the risks.

To investigate the pH of a substance, you will need to be able to do the following:

1. Measure out 50cm^3 of hydrochloric acid into a beaker.
2. Record the pH of the solution by putting some Universal Indicator paper onto a white tile and place a drop of the acid onto it.
3. Leave it for 30 seconds to make sure the colour change is complete and then record the pH.
4. Measure out 0.3g of calcium hydroxide and add it to the beaker.
5. Record the pH of the substance and repeat until 2.4g of the solid is added.
6. Plot a graph as seen on the right.

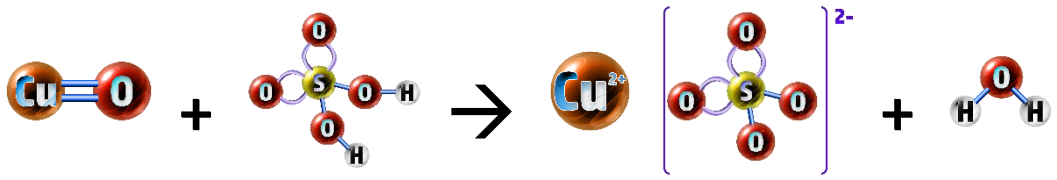


As you can see from the experiment:

- The pH stays at 1 until 1.5g of Calcium Hydroxide is added.
- After 1.5g, the pH starts to increase rapidly and continues until about 2.7g of the solid is added, at which point it stays at pH 10.
- The reaction is **neutral** at 2.1g.

Risks and management:	Improvements:	Variables:
• Calcium oxide = corrosive/ Calcium hydroxide = irritant – weak goggles and if you get it on your hands, wash it off.	• Use a pH meter instead of Universal Indicator paper. • It is more accurate (1 dp.) than UI paper.	• Independent: Mass of $Ca(OH)_2$ • Dependent: pH of substance • Control: Volume of acid, concentration of acid, type of tablet, etc.

Section 12: Preparing Copper Sulfate Crystals:



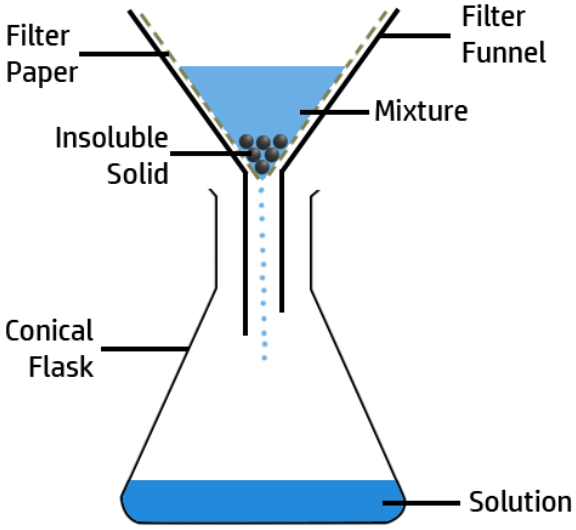
$CuO_{(s)}$ + **$H_2SO_{4\,(aq)}$** → **$CuSO_{4\,(aq)}$** + **$H_2O_{(l)}$**

Copper Oxide + Sulfuric Acid → Copper Sulfate + Water

Example: How can you produce soluble copper sulphate from sulphuric acid and **insoluble** copper oxide?

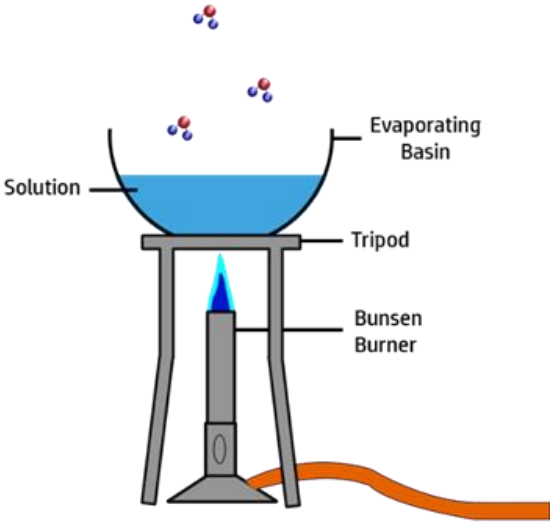
Phase One: Making the Copper Sulfate

1. **Heat the acid in a water bath** to speed up the reaction. (In a fume cupboard to reduce acidic fumes)
2. **Add the copper oxide to the acid**. A reaction will occur, producing copper sulfate – your soluble salt.
3. **Continue adding copper oxide until it is neutral**. You can check this with **pH paper**. You now have a mixture of your copper sulfate, water and **excess** (unreacted) copper oxide.



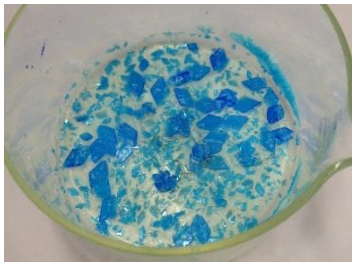
Phase Two: Filtering the solution

4. **Use a filter funnel and filter paper to remove the excess copper oxide**, leaving you with the copper sulphate and water.



Phase Three: Crystallisation

5. **Heat the solution gently**. To boil half of the water.
6. **Leave the solution to cool.** The rest of the water will **evaporate,** and **crystals** of copper sulfate will start to form.



7. **Dry the copper sulfate** between pieces of filter paper.

Hazard:	Risk:	Reducing Risk (Precaution)
Copper oxide and copper sulfate are **irritants**	Medium	Wear goggles and gloves. Wash off if you get it on hands.
Bunsen Burner flame can ignite flammable chemicals	Medium	Keep away from flammable substances and tie hair back.

Section 13: Titrations and Precipitates:

A: How to prepare soluble salts from an acid and a soluble reactant:

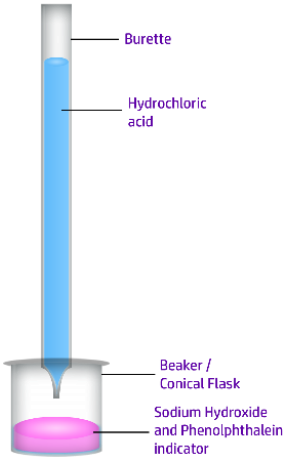
If you want to make a salt from a soluble reactant you **cannot** use filtration.

As you add the **acid** to your soluble **alkali**, the pH will not stop at **pH 7 (neutral)**. This is because the alkali will dissolve and dissociate into OH^- ions, increasing the pH and turning the solution alkaline.

To get around this, you need to use a **titration**. You need to add the acid to the soluble alkali until it is **exactly neutral** (you can do this using an **indicator**) – at which point you can boil the solution and then leave the water to evaporate, leaving your **crystals** behind:

Method:

1. Fill your **burette** with acid.
2. Use a **pipette** to put your alkali into a **conical flask**.
3. Add an **indicator** such as **phenolphthalein** (which will turn pink in the alkali).
4. **Turn the tap** adding the hydrochloric acid drop by drop until it reaches the end point (**pink** → **colourless**).
5. **Repeat *without* the indicator**, this gives you just sodium chloride and water.
6. **Heat the solution gently**. Boil the solution until half of the water has evaporated.
7. Leave the solution to **cool**, and once the rest of the water has evaporated, you will be left with **crystals** of your salt which can be **dried** using filter paper.



B. How to prepare insoluble salts:

A **precipitate** is an **insoluble product** formed when two solutions are mixed together. The reaction where this occurs is called a **precipitation reaction**.

In a **precipitation reaction** the endings of your two products swap around:

Silver nitrate + copper chloride → silver chloride + copper nitrate

Salts:	Soluble:	Insoluble:
Nitrates	**All soluble**	-
Chlorides	Mostly soluble	Silver chloride and lead chloride
Sulfates	Mostly soluble	Lead, barium & calcium sulfate
Carbonates / hydroxides	Sodium/potassium/ammonium	Mostly insoluble

You can see from the solubility table above that most chlorides are soluble **except** for **silver chloride** and lead chloride. Therefore, silver chloride is insoluble = **your precipitate**!

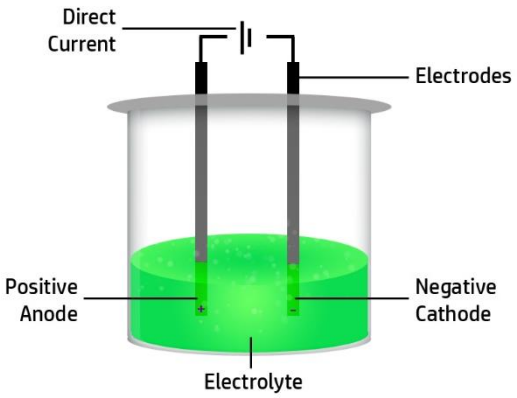
If you wanted to prepare a pure dry precipitate of **silver chloride** from the above reaction, there are 5 steps you need to follow:

1. **Dissolve** the solid (copper chloride and silver nitrate).
2. **Mix** them together (to produce your copper nitrate and silver chloride).
3. **Filter** the solution (to give you silver chloride in the filter paper).
4. **Wash** the filter paper (with deionised water to remove the impurities).
5. **Dry** the precipitate (between pieces of filter paper or in an oven).

Section 14: Electrolysis:

A. What is Electrolysis?

- Electrolysis is the *breaking down* of an **electrolyte** using **electrical energy**.
- An electrolyte is any liquid that contains **ions**. The liquid can be *molten* (melted) or *aqueous* (dissolved).
- The electrical energy needed is **direct current**.
- Normally it involved **inert** (unreactive) electrodes which aren't part of the reactions.



B. How do you remember the name and charge of the electrodes?

If you're in the exam and are trying to remember the name of the electrodes, don't forget to PANIC!

P⁺ositive
Anode
N⁻egative
Is
Cathode

The anode is always positive, and the cathode is always negative.
To remember that, use the acronym on the left.

Another way to remember it is that:

- The cations will go to the cathode.
- Cations are 'paw'sitive.
- Opposites attract, so the cathode must be negative.

C. What forms at the electrodes?

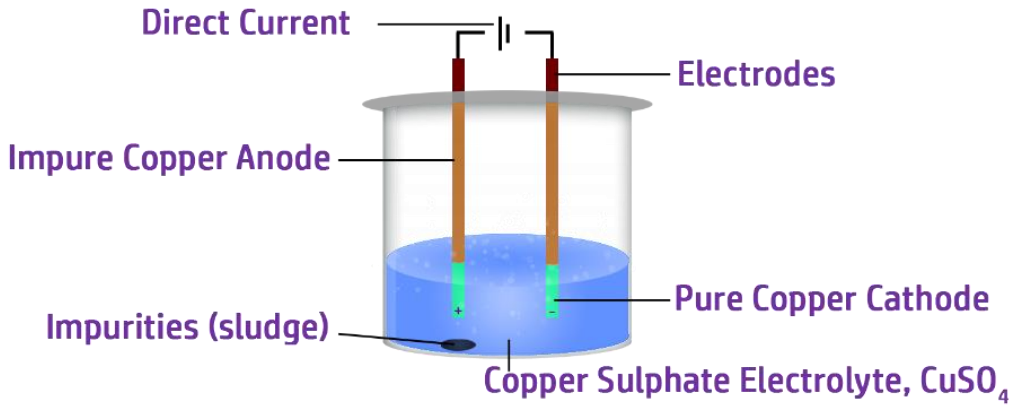
1. Molten electrolytes: Example – Lead Bromide, $PbBr_{2\,(l)}$
   - When you have a molten electrolyte, you only have two ions in the liquid – the metal and the non-metal.
   - At the cathode, lead ions (Pb^{2+}) will gain electrons and turn back into lead atoms.
   - Half Equation: $Pb^{2+} + 2e^- \rightarrow Pb$
   - At the anode, bromide ions will lose electrons and turn back into bromine molecules, Br_2.
   - Half Equation: $2Br^- \rightarrow Br_2 + 2e^-$

2. Aqueous electrolytes:
   - When you electrolyse an aqueous solution, you have H^+ and OH^- ions **as well** as the metal and non-metal ions.
   - For example: copper chloride, $CuCl_2$, will have Cu^{2+} and H^+ cations as well as Cl^- and OH^- anions.
   - You need to be able to work out what will form. There are two rules to know:
     - Cathode: The **least** reactive ion will form.
     - Anode: If there is a **halide** (group 7), that will form – **if not, the hydroxide will go there, and form water and bubbles of oxygen** will.

Aqueous Solution:	Form at cathode:				Form at anode:			
Copper chloride, $CuCl_{2\,(aq)}$	Cu^{2+}	✓	H^+	✗	Cl^-	✓	OH^-	✗
Sodium chloride, $NaCl_{(aq)}$	Na^+	✗	H^+	✓	Cl^-	✓	OH^-	✗
Sodium sulphate, $Na_2SO_{4\,(aq)}$	Na^+	✗	H^+	✓	SO_4^{2-}	✗	OH^-	✓
Sulfuric acid, $H_2SO_{4\,(aq)}$	NA		H^+	✓	SO_4^{2-}	✗	OH^-	✓

Section 15: Electrolysis of copper sulfate:

Method:



You will need to be able to explain how to investigate the change of current on the mass of the electrodes in the above set up.

1. Add some **copper sulfate** to a beaker.
2. **Clean** both copper electrodes
3. Measure the initial **mass** of your copper cathode and copper anode.
4. Turn on the power and use a **variable resistor** to change the current to 0.2A.
5. Leave the power on for **20 minutes**.
6. Wash and **dry** the electrodes using propanone.
7. Record the final mass of the electrodes and calculate the **change in mass** for the anode and cathode.

Results:

- The mass of the **anode** will **decrease**.
- The mass of the **cathode** will **increase**.
- The solution will remain **blue**.
- The change in mass between the electrodes will ***not*** be the same.

Explanation:

- **Anode:** The copper atoms **lose two electrons** and turn into **Cu^{2+}** ions (**oxidation**). This causes the mass of the anode to **decrease**.
- **Cathode:** The copper ions from the anode will move to the cathode and **gain two electrons**, turning back into copper atoms (**reduction**). This causes the mass of the cathode to **increase.**
- **The change in mass** is never the same because there are **impurities** on the anode. These impurities fall to the bottom – this is called **sludge**.
- The solution remains **blue** because the **same amount of ions** enter and leave the solution.

Risks and management:

Hazard:	Risk:	Reducing Risk (Precaution)
Copper Sulfate is an **irritant**	Medium	Wear goggles and gloves. Wash off if you get it on hands.
Electrodes can cause an **electric shock**	Low	Make sure electrodes don't touch. Turn off electricity before handling.

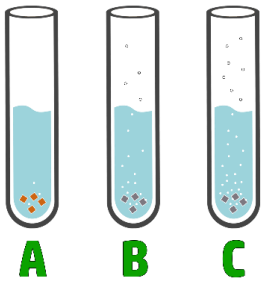
Variables:

- **Independent variable (the thing you change!):** The **current**, A.
- **Dependent variable (the thing you measure!):** The **change in mass** of the electrodes.
- **Control variables (the things you keep the same!):** The **volume** of copper sulfate, the **concentration** of copper sulfate, how far the electrodes are dipped into the copper sulfate.

Section 15: Extracting Metals and Reactivity:

A. *How do we know how reactive a metal is?*

When you react a metal with water and different acids, you can use the observations to work out how reactive they are:



- The **more bubbles** / fizzing / effervescence there is, the **more reactive** the metal.
- On the right, you can see that **'C'** is the *most* reactive metal because it produces the ***most*** bubbles.
- A is the least reactive as it has the ***least*** bubbles.
- You can also look at **temperature changes**. The **bigger** the temperature change, the **more reactive** the metal.

B. Extracting Metals:

An ore is a rock that contains enough metal to make it financially worthwhile to extract. If it isn't going to give a profit to extract it – it is not an ore!

- Most metals are found in ores, such as bauxite – aluminium oxide, Al_2O_3
- Some metals are found uncombined – these are the metals that don't react.

There are three ways of removing metals from the ground:

- **Dig them out** – this can only be done to the **unreactive metals** silver, gold and platinum
- **Heat them with carbon** – this can only be done with metals less reactive than carbon – the transition metals.
- **Electrolysis** – this is for elements more reactive than carbon as it is expensive.

Potassium
Sodium
Calcium
Magnesium
Aluminium
Zinc
Iron
Tin
Copper
Silver
Gold

C. Biological Methods of Extracting Metals: (Higher)

There are two other methods of extracting metals which are used to extract rocks with small amounts of metals in – called **low grade ores:** Bioleaching and Phytoextraction

Process:	Method:	Advantages:	Disadvantages:
Bioleaching	• Bioleaching uses bacteria to separate metals from low grade ores.	• Does not require higher temperatures • No harmful gases • Less damage to landscape than mining • Conserves supplies of higher-grade ores	• Very slow • Toxic substances and sulphuric acid can be produced – damaging the environment
Phytoextraction	• Phytoextraction involves growing plants that absorb metal compounds, which are then burnt to produce the metal.	• Can extract metals from contaminated soil • No harmful gases • Less damage to landscape than mining • Conserves supplies of higher-grade ores	• Very slow • More expensive than mining some ores • Growing plants depends on the weather/climate

D. Oxidation and Reduction

Iron oxide + carbon → iron + carbon dioxide

- Carbon has been oxidised because it has had **oxygen** added (to form carbon dioxide)
- Iron oxide has been reduced because it has had oxygen removed (to form iron).

When metals are oxidised, they **corrode**. When iron **rusts**, it isn't just reacting with oxygen, but water as well.

Section 16: Recycling:

A. Reasons to recycle metals:

- Conserves Earth's natural resources
- Less mining of ores is needed. This is good because:
  - It doesn't damage the landscape
  - It doesn't create noise/dust pollution
- It can take less energy to recycle than to extract from the ore
- Metals don't end up wasted in landfill sites.
- However – it can cost time and money to recycle 😟

B. Life Cycle Assessments

A lifestyle assessment looks at the **total environmental cost** of a product by looking at each stage of the life of a product:

Stage:	Description:
Choice of material	• When metals are extracted from their ores, it needs a lot of energy and can produce a lot of pollution. • Raw materials can come from crude oil, which is non-renewable. • Crude oil also gives out greenhouse gases when combustion occurs.
Manufacture	• Manufacturing needs a lot of energy • It *can* cause a lot of pollution • Waste products need disposing of safely – some can be recycled • The water used in lots of manufacturing needs to be safe/unpolluted when put back into the environment

Product use	• The products can themselves be harmful – such as:
	  - Toxic fumes from paint,
	  - Toxic gases from combustion and
	  - Fertilisers draining into lakes/rivers causing eutrophication

Disposal	• Lots of products are disposed of in landfills, which takes up space and can pollute the land/water • Products can also be burnt – which can give off toxic gases or greenhouse gases

Each of these factors need to be considered when manufacturing a product. If there are multiple ways of producing a product, the one that has the least environmental cost will be chosen.

Example: A decision needs to be made about whether to produce a cabinet from two different sources:

Source:	Material	Waste solid produced	Water used (m^3)	Expected lifespan (years)
A	Iron from an iron ore	15,000 kg	8.2	20
B	Recycled iron	5,400 kg	6.5	40
C	Recycled steel	9,000kg	4.5	14

- You would not choose A because the iron comes from an ore – which takes more energy to extract than recycling. It also produces the most waste and uses more water.
- You would choose B because it is recycled, which uses less energy, and produces the least waste. It also will last the longest and therefore has the least environmental cost.

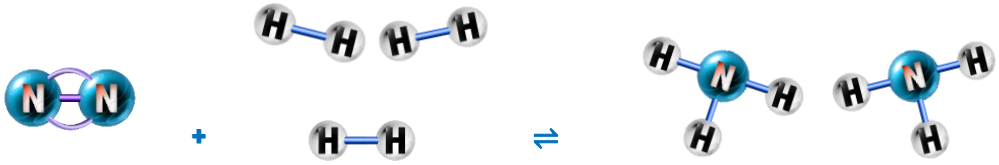
Section 17: Reversible Reactions:

A. What is a reversible reaction?

A **reversible reaction** is any reaction that can go in both the **forward direction** and **backwards direction**.

- It has a different symbol: '⇌' instead of '→
- A reversible reaction can occur at the same time or happen at different times.

An example of a reversible reaction is the **Haber Process**:



Nitrogen	Hydrogen	Ammonia
N_2	$3H_2$	$2NH_3$

In this reaction, you can see that nitrogen, N_2, and hydrogen, H_2, react to form ammonia, NH_3. At the same time, ammonia is decomposing to form nitrogen and ammonia. To make it easier to write, we use the ⇌ symbol.

B. Investigating Reversible Reactions

You could be asked to investigate the reversible reaction involving copper sulfate:

Hydrated Copper Sulfate	⇌	Anhydrous Copper Sulfate	+	Water
$CuSO_4.5H_2O$		$CuSO_4$		$5H_2O$

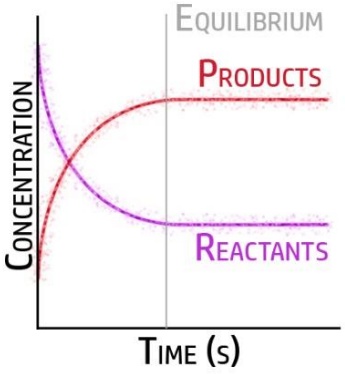
Blue copper sulfate is hydrated – meaning it contains water and has the formula of $CuSO_4.5H_2O$.

When **heated**, the water *escapes* and forms anhydrous copper sulfate, $CuSO_4$, which is a white powder.

If you then add the water back into the white powder, it turns back into your **blue** hydrated copper sulfate.

What is dynamic equilibrium?

- **Dynamic equilibrium** is when the forward reaction and backward reaction are occurring at the **same time and the same rate**, meaning the concentration of reactant and product stay the same.
- On the graph on the right, you can see where equilibrium is reached because the concentration has stayed the **same** – both lines are flat.
- Dynamic equilibrium only occurs in a **closed system**, where nothing is allowed to enter or leave the environment.



When dynamic equilibrium is met in the Haber Process:

1. The concentration of N_2, H_2, and NH_3 does **not** change.
2. N_2 and H_2 are reacting to form NH_3 at the **same time and rate** as NH_3 is breaking down into N_2 and H_2.

- There may be 80% NH_3 and 20% N_2 and H_2, but as long as the percentage stays the **same**, it has reached **equilibrium**. **Iron** is the catalyst used in the Haber Process.
- It does not change the position of the equilibrium, but it does speed up the production – so used producing ammonia is quicker.

Section 18: Transition Metals, Alloys and Corrosion:

A. The Transition Metals

The transition metals are found in the centre of the Periodic Table and contain the most used metals used in wiring, jewellery, vehicles, etc...

- The transition metals share the same **physical properties** as most metals, but they do have **higher melting points** and **densities** than other metals
- One of the main **chemical properties** of the transition metals is that they form ***colourful compounds*** and can be used as catalysts (which speed up a chemical reaction without being changed chemically)

B. Corrosion vs Rusting

Corrosion is the **weakening of a metal over time** through **rusting**, **corrosion,** or **chemical reactions**.

The more reactive a metal, the more quickly it corrodes. This is because they **lose their electrons faster**. Some metals, such as gold, do not corrode at all.

Rusting ***only*** occurs in **iron**. Iron reacting with oxygen is the same as corrosion.
For iron to rust it needs oxygen **and** water.

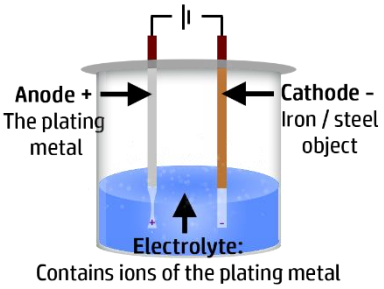
iron + oxygen + water → hydrated iron oxide

C. Preventing Corrosion:

1. Removing Oxygen: Storing the metal in argon / nitrogen
2. Physical Barriers: Coating the metal with paint / oil or plastic.
3. Sacrificial Protection: Putting a more reactive metal over the iron. The oxygen will then react with the more reactive metal, instead of the iron.
4. Galvanizing: is a combination of a physical barrier **and** sacrificial protection. The iron/steel is coated in a thin layer of zinc.
5. Electroplating: Electroplating uses **electrolysis** to put a thin layer of a **less reactive** metal onto the object you want to protect.

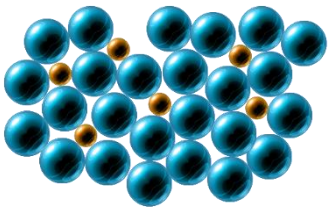
D. How does electroplating work:

- The **cathode** (**negative** electrode) should be the metal you want to protect.
- The **anode** (**positive** electrode) should be the metal you are using to protect it.
- The **electrolyte** (**ionic** liquid) should contain ions of the protecting metal.
- The metal on the anode will turn into an ion (losing electrons) and move to the cathode, coating it as it turns back into a metal.



E. Alloys:

An **alloy** is a mixture of a metal and another element – usually another metal.
Pure metals have particles all the same size, so the **layers** can slide past each other – making them soft.
Alloys have different sized particles, so the layers cannot slide past each other – this makes them stronger.



F. Uses of Metals and Alloys:

Name:		Use:	Information:
Gold	**Pure Metal**	**Memory Chips**	Excellent conductor electricity and malleable
Jewellery Gold	**Alloy**	**Jewellery**	Stronger than gold but still attractive.
Copper	**Pure Metal**	**Wires / Pipes**	Good conductor of electricity and ductile
Brass	**Alloy**	**Plug Pins**	Stronger than copper and good conductor.
Aluminium	**Pure Metal**	**Overhead Cables**	Good conductor of electricity and low density.
Magnalium	**Alloy**	**Aircraft Parts**	Stronger than aluminium but still low density.

Section 19: Concentration and Yields

A. How to find an unknown concentration using a titration:

Example: It takes 20cm^3 of hydrochloric acid to completely neutralise 25cm^3 of 0.1moldm^{-3} sodium hydroxide. Calculate the concentration of the hydrochloric acid.

$HCl + NaOH \rightarrow NaCl + H_2O$

The Calculation

- **Step 1:** Convert the volumes to dm^3: HCl = 20/1000 = 0.02dm^3, NaOH = 25/1000 = 0.025dm^3
- **Step 2:** Find out the moles of sodium hydroxide = conc x volume = 0.1 x 0.025 = 0.0025mol
- **Step 3:** The ratio between HCl and NaOH = 1:1, so moles = the same. You can therefore work out the concentration for HCl = mol / vol = 0.0025mol / 0.02dm^3 = 0.125moldm^{-3}

B. Yields:

Percentage Yield = Actual Yield ÷ Theoretical Yield x 100

Example: *In a chemical reaction sodium carbonate is reacted with excess hydrochloric acid. The maximum mass of sodium chloride is* ***5.85g****, however only* ***3.7g*** *of* ***sodium chloride*** *is actually produced. Calculate the percentage yield and give your answer to 3 significant figures.*

- **Step 1: 3.7g** (actual yield) ÷ **5.85g** (theoretical yield) x 100 = **63.247%**
- **Step 2:** Give the answer to 3 Significant Figures: **63.2**47% → **63.2%**

Why is the yield less than 100%?

It is very rare that a reaction produces 100% yield. There are three main reasons for this:

- Some of the products is lost when transferred between containers
- Side reactions may have occurred
- The reaction may not have finished. (One of the reactants may have run out.)

C. Atom Economy:

Atom Economy is the percentage of mass converted into useful products and is calculated using the equation →

$$\text{Atom Economy} = \frac{M_r \text{ of useful product}}{\text{Total } M_r \text{ of products}} \times 100$$

Example: Ammonium chloride reacts with calcium hydroxide to form ammonia, calcium chloride and water. Work out the atom economy for the formation of **calcium chloride** in this reaction.

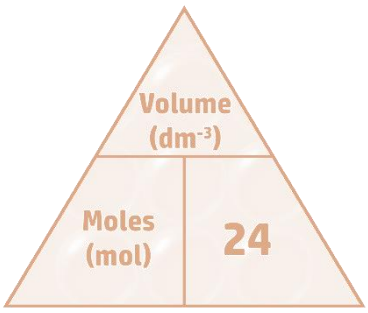
$2NH_4Cl + Ca(OH)_2 \rightarrow 2NH_3 + CaCl_2 + 2H_2O$

- **Step 1:** Find out the formula mass (M_r) for calcium chloride (40 + 71 = **111**)
- **Step 2:** Find out the **total** formula mass of all the products combined:
  - $2NH_3$: 2 x (14 + 3) = 34
  - $2H_2O$: 2 x (2 + 16) = 36
  - Total formula mass of products = 34 + 36 + **111** = **181**
- **Step 3:** Use these two values to calculate the percentage of useful products:
  - **(111 / 181) x 100 = 61.3%**

Section 20: Molar Volumes of Gases

A. What is the Molar Volume?

The **molar volume** is the volume of a gas when you have **one mole** of molecules.



At room temperature (20°C) and pressure (1atm), the molar volume for 1 mole is always **24dm³mol⁻¹**.

B. How do you calculate Molar Volume?

You can use this to work out masses and volumes for any balanced equation:

Example 1: *Calculate the volume of 0.50 mol of oxygen at room temperature & pressure. (Molar volume = 24dm³mol⁻¹)*

- **Volume = moles x molar volume**
- **Volume =** 0.5 x 24 = **12dm³**

Example 2: *In an experiment, hydrogen gas is made by reacting 20.0g of sodium with excess water. Calculate the maximum **volume** of **hydrogen** that could be formed. (A_r: Na = 23, H = 1, O = 16)*

$$2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(s)} + H_{2(g)}$$

- **Step 1:** Find out the formula mass for **ONE** mole of sodium = **23**
- **Step 2:** Find out the **moles** for sodium = mass ÷ M_r = 20g / 23 = **0.870** moles
- **Step 3:** By looking at the balanced equation, you can see that you have **2Na** to **1H_2**. This is a ratio of **2:1**, so to work out the moles of hydrogen, divide 0.870 by 2: **0.870÷2** = **0.435 moles of hydrogen**.
- **Step 4:** Now you have the moles, you can calculate the volume. Volume of hydrogen = **moles of H_2** x **24** = **0.435 moles** x **24** = **10.4dm³**

How can you calculate volumes from balanced equations?

Avogadro's law states that if the **temperature** and **pressure** are the **same**, the *volume will be the same* if you have the *same number of molecules*. You can use this to work out the volumes from balanced equations:

Example: **200dm³** of **hydrogen** reacts with **oxygen** to form water vapour. What volume of **oxygen** would completely react with the hydrogen?

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O(g)$$

The ratio of moles between **hydrogen** and **oxygen** is **2:1** – so the volume of oxygen is half of the volume of hydrogen – **200dm³** ÷ 2 = **100dm³**.

The volume of water produced would be 200dm³ because the ratio is 2:2.

Section 21: Fertilisers

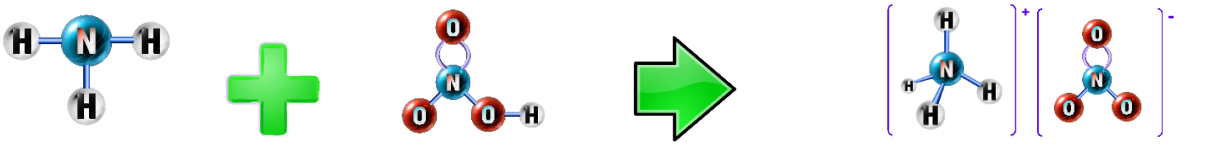
A. Fertilisers

When plants grow, they absorb minerals from the soil. Over time, the soil can become mineral deficient which can stop plants from growing – or cause mineral deficiency diseases in the plants. Farmers get round this by using **fertilisers**.

Fertilisers contain three key minerals: **Nitrogen**, **Phosphorus** and **Potassium**:

#	Name:	Formula:	Elements:	From:
1	Ammonium Nitrate	NH_4NO_3	**Nitrogen**	Nitrate ions – NO_3^- Ammonium ions – NH_4^+
2	Ammonium Sulfate	$(NH_4)_2SO_4$	**Nitrogen**	Nitrate ions – NO_3^-
3	Ammonium Phosphate	$(NH_4)_2PO_4$	**Nitrogen** **Phosphorus**	Nitrate ions – NO_3^- Phosphate ions – PO_4^{2-}
4	Potassium Nitrate	KNO_3	**Potassium** **Nitrogen**	Potassium ions – K^+ Nitrate ions – NO_3^-

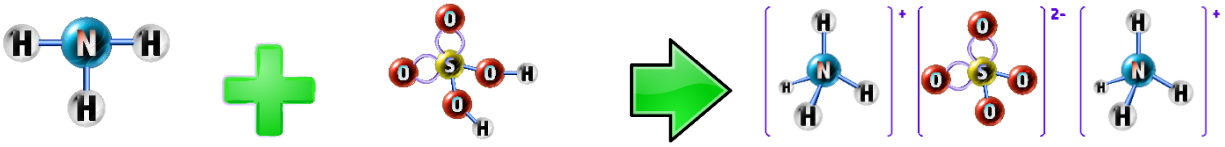
Making Ammonium Nitrate:



Ammonia, NH_3 Nitric acid, HNO_3 Ammonium nitrate, NH_4NO_3

Making Ammonium Sulfate:

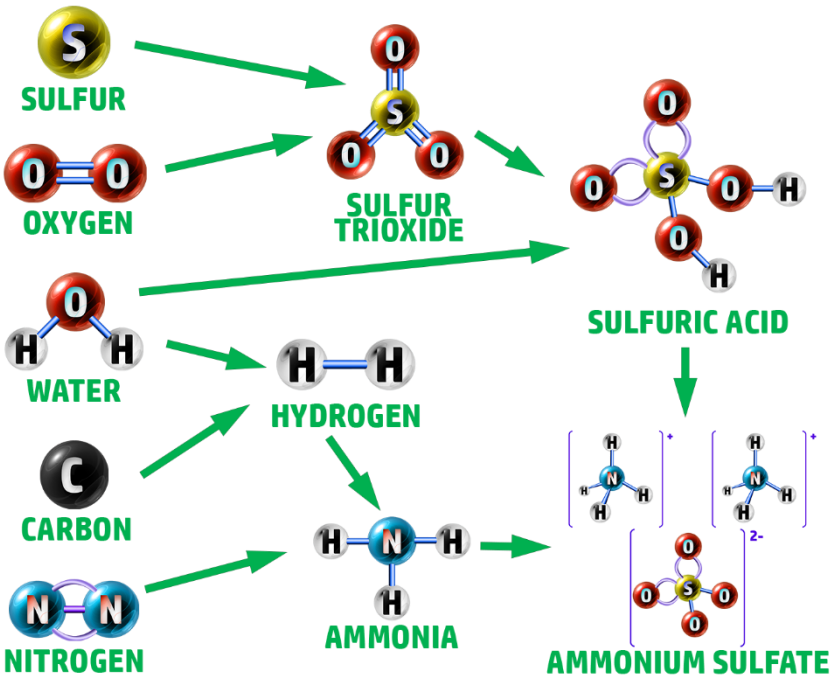
Ammonium sulfate can be produced both in the Science labs and in industry. In the lab, it is as simple as a titration between ammonia solution and sulfuric acid, followed by crystallisation:



Ammonia, NH_3 Sulfuric acid, H_2SO_4 Ammonium sulfate, $(NH_4)_2SO_4$

This is a **batch** process, and is much slower than industry, so is not good for making large amounts of the fertiliser.

In industry, there are several steps:



- The raw materials are all added continuously:
  - Sulfur and oxygen react to form **sulfur trioxide**.
  - Water and sulfur trioxide then form **sulfuric acid**.
  - Water and carbon are used to make **hydrogen**.
  - Hydrogen and nitrogen react to form **ammonia** (Haber Process).
  - Ammonia and sulfuric acid react to form **ammonium sulfate**.
- This is a **continuous** process and can be made on a large scale.

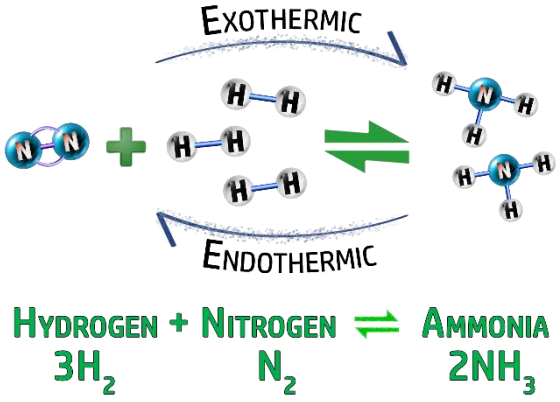
Section 22: The Position of Equilibrium & Fuel Cells

What factors affect the position of equilibrium? (Higher)

In a reversible reaction, changing the concentration, pressure and temperature can affect the position of equilibrium. For example, in the Haber Process, the backwards reaction is endothermic.

Increasing Temperature:

- **Increasing the temperature** moves equilibrium in the direction of the **endothermic** reaction.
- **Equilibrium shifts to the left**, **decreasing** the yield of ammonia.
- You will, however, reach equilibrium **quicker**, as the reaction occurs faster - so can get your yield faster.



Concentration:

- **Increasing the concentration** of reactants **moves equilibrium to the right** ($2NH_3$) and **decreases the time** to reach equilibrium.

Pressure:

- **Increasing the pressure** moves equilibrium towards the side with the **fewest molecules** – in this case the $2NH_3$
- **Equilibrium shifts to the right** and reaches equilibrium faster and giving you a larger yield of NH_3.

Catalyst:

Using a catalyst **doesn't** affect the position of equilibrium, but because it lowers the activation energy, making more collisions successful, it means you will **reach equilibrium faster**.

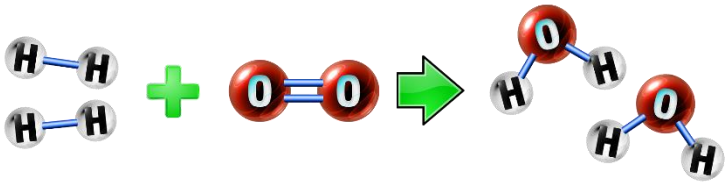
C. Fuel Cells

Chemical cells – such as everyday batteries – produce a voltage until one of the reactants is used up. They contain two parts:

- Two different metals with solutions of their salts
- A salt bridge to move the ions between the two parts of the cell.

A fuel cell is made up of **hydrogen** (from fuels) and **oxygen** (from air):

- During the reaction, hydrogen loses electrons (oxidation) to form hydrogen ions:
  $2H_2 \rightarrow 4H^+ + 4e^-$



- The hydrogen ions then react with oxygen, gain electrons (reduction) and form water:
  $4H^+ + 4e^- + O_2 \rightarrow 2H_2O$

Advantages of Fuel Cells:	Disadvantages of Fuel Cells:
More efficient than using petrol	Hydrogen is a gas – so it takes more space to store
No moving parts means less energy lost due to friction	Hydrogen is explosive – so it is harder to store
Less reaction steps means less heat lost	Hydrogen is produced from hydrocarbons or electrolysis – both of which use fossil fuels and give off carbon dioxide.
No pollutants such as CO_2, NO_2, SO_2 and CO	