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# 1 States of matter

## 1.1 - Solids liquids and gases

Solids	Liquids	Gas
<ul style="list-style-type: none"> <li>- Fixed shape</li> <li>- Fixed volume</li> <li>- Cannot flow</li> <li>- Cannot compress</li> <li>- Arranged so particles are tightly packed in lattice</li> <li>- Particles vibrate in place</li> </ul>	<ul style="list-style-type: none"> <li>- Takes shape of container</li> <li>- Fixed volume</li> <li>- Can flow</li> <li>- Hard to compress</li> <li>- Arranged so particles are touching</li> <li>- Particles slide over each other</li> </ul>	<ul style="list-style-type: none"> <li>- No fixed shape</li> <li>- No fixed volume</li> <li>- Can flow</li> <li>- Can compress</li> <li>- Free spread out particles</li> <li>- Move around quickly and randomly</li> </ul>

### Changes of state

**Melting:** Solid → Liquid

**Boiling:** Liquid → Gas

**Evaporating:** Liquid → Gas

**Condensing:** Gas → Liquid

**Freezing:** Liquid → Solid

### Evaporation Vs Boiling

Evaporation	Boiling
<ul style="list-style-type: none"> <li>- Surface level</li> <li>- Does not happen at a fixed temperature</li> <li>- Bubbles not produced</li> <li>- Remaining liquid is cooled</li> </ul>	<ul style="list-style-type: none"> <li>- Happens throughout liquid</li> <li>- Happens at a fixed temperature (boiling point)</li> <li>- Bubbles produced</li> <li>- Remaining liquid is not cooled</li> </ul>

### Kinetic Particle Theory:

- When a substance is changing it can either
  1. **Gain kinetic energy** from heat, which causes the particles to vibrate more vigorously. They then gain enough energy to break free of their bonds. E.g. Liquid → Gas
  2. **Lose energy** to their surroundings. The particles move closer to each other, moving less vigorously. E.g. Gas → Liquid

Factors affecting changing states:

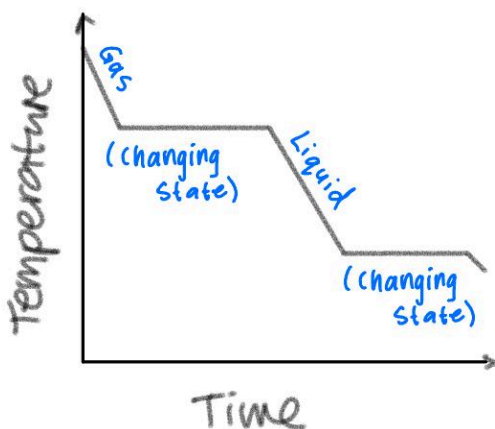
+ *Temperature:*

Increases kinetic energy of particles, particles move more rapidly.

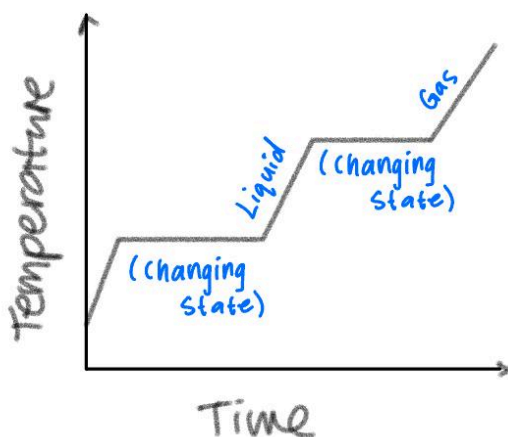
+ Pressure:

Pushes particles closer to each other. Can change the state of matter if decreased or increased sufficiently (E.g. if a gas is under very high pressures, the particles can be pushed together so densely that a liquid is formed.)

Cooling curve:



Heating curve:



## 1.2 - Diffusion

The net movement of particles from an area of high concentration to an area of low concentration down a concentration gradient as a result of random particle movement.

Factors affecting the rate of diffusion:

1. **Relative molecular mass:** The heavier the molecular mass, the slower the diffusion
2. **Temperature:** The higher the temperature, the faster the diffusion, as particles have more energy and move quicker.

## 2 Atoms, elements and compounds

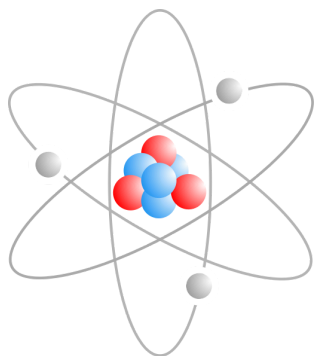
### 2.1 - Elements, compounds and mixtures

Elements	Compounds	Mixtures
<ul style="list-style-type: none"><li>- Atoms of an element are identical to each other</li><li>- They cannot be changed chemically into another element</li><li>- Cannot be broken down into simpler forms</li></ul>	<ul style="list-style-type: none"><li>- Atoms of different elements bond together in a chemical reaction to form compounds</li><li>- Can be changed chemically to a different compound</li><li>- Can be broken down into simpler forms</li></ul>	<ul style="list-style-type: none"><li>- Particles that are mixed without any chemical bonds</li><li>- Mixture retains properties of its particles</li><li>- Can be separated using techniques such as filtration and evaporation.</li></ul>

### 2.2 - Atomic structure and the Periodic Table

*Structure of an atom*

	Relative charge	Relative mass
<b>Proton</b>	+ 1	1
<b>neutron</b>	0	1
<b>Electron</b>	- 1	1/1840 (negligible)



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**Atomic number:** The number of protons in an atom of an element

**Nucleon number:** The number of protons and neutrons in an atom of an element

*Electronic configuration* - The way electrons in an atom of an element are distributed in its shells

**First shell** can hold up to 2 electrons

**2nd 3rd shells** can hold up to 8 electrons

E.g. :

Potassium (19 electrons) = 2, 8, 8, 1

Carbon (6 electrons) = 2, 4

- The number of electron shells occupied represents an element's **period**

E.g. Potassium is in period 4 and Carbon is in period 2

- The number of outer shell electrons (electron valency) is equivalent to an element's **group number**

E.g. Potassium is in group 1 and Carbon in group 4

\*Note that group VIII elements have full outer shells

### 2.3 - Isotopes

Isotopes are atoms of the same element with the same number of protons but different number of neutrons.

- Isotopes of an element have the same electronic configuration, and thus still have the same chemical properties.

**Relative atomic mass (the average mass of isotopes of an element in comparison to 1/12th of a mass of an atom of carbon-12):**

$(\text{Abundance} \times \text{mass number}) + (\text{abundance} \times \text{mass number}) / 100$

### 2.4 - Ions and ionic bonds

**Ion** - an atom or group of atom that have an electric charge (+ or -)

**Cations:** Positively charged ions

**Anions:** Negatively charged ions

- An atom loses or gains electrons to obtain a full outer shell, which will stabilise the atom.

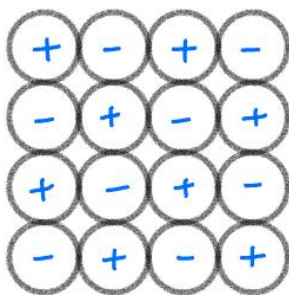
E.g.:

Potassium has an ion of 1+, it loses an electron to achieve a full outer shell 2, 8, 8, 1 → 2, 8, 8

Oxygen has an ion of 2-, it gains 2 electrons to achieve a full outer shell 2, 6 → 2, 8

**Ionic bonds:**

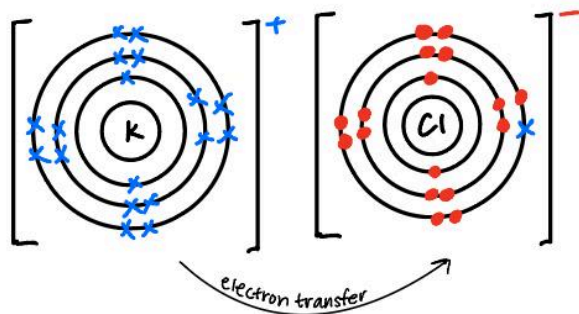
The strong electrostatic force between oppositely charged ions



- Between a metal (positive ion) and a nonmetal (negative ion)
- Ionic bonds are strong: high melting and boiling point (lots of energy is required to break the strong bonds)
- Soluble in water
- Conducts electricity when molten or aqueous only
- Brittle (positive and positive charges repel whe layers slide over each other when hit)
- E.g. Potassium oxide

Ionic bonds are represented by a dot and cross diagram:

E.g. potassium chloride :  $KCl$



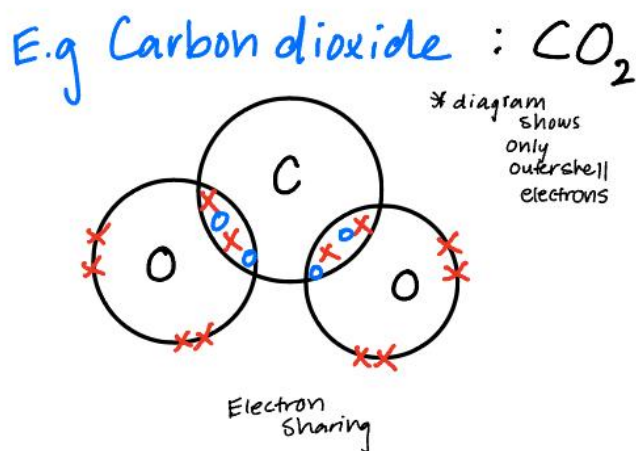
## 2.5 - Simple molecules and covalent bonds

### Covalent bonds:

The sharing of pairs of electrons between 2 atoms to achieve full electron valency.

- Between nonmetal and nonmetal (2 negative ions)
- Weak intermolecular forces: low melting and boiling point (less energy is required to break the bonds)
- Insoluble in water
- Does not conduct electricity
- E.g. Carbon dioxide

Covalent bonds are represented by a dot and cross diagram in a different form:

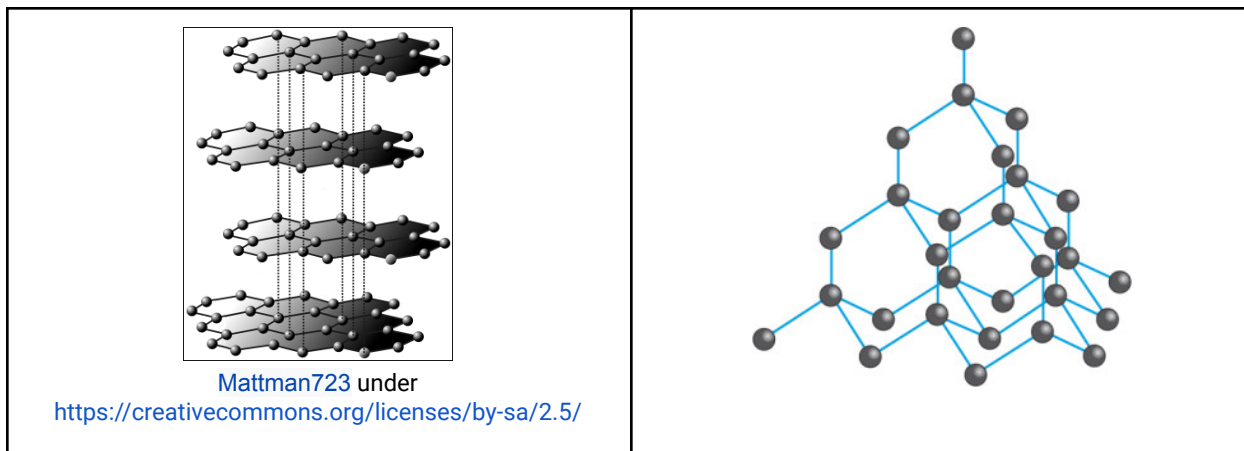


## 2.6 - Giant covalent structures

The result of many atoms bonded covalently

The 2 covered giant covalent structures:

Graphite	Diamond
<ul style="list-style-type: none"><li>- Used as a lubricant and pencils</li><li>- Hexagonal structure</li><li>- Once carbon bonded to 3 carbon: 1 delocalised</li><li>- Layers can be easily rubbed off as they are soft</li><li>- Insoluble</li><li>- Conducts electricity: delocalised electron</li></ul>	<ul style="list-style-type: none"><li>- Used in cutting tools</li><li>- Tetrahedral structure</li><li>- One carbon is bonded to 4 carbon</li><li>- Hard</li><li>- Insulator</li><li>- Do not conduct electricity</li><li>- Insoluble</li></ul>



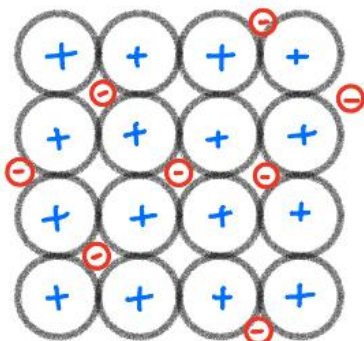
- Graphite and diamond and allotropes: They are both made up of carbon, but have different structures

\*Silicon (IV) oxide,  $\text{SiO}_2$ , is another example of a giant covalent structure. It has similar properties to diamond and also has a tetrahedral structure

- One oxygen bonded to 2 silicon
- 1 silicon bonded to 4 oxygen

## 2.7 - Metallic bonding

The electrostatic force of attraction between positive ions and a sea of delocalised electrons



- **High melting a boiling point**  
(due to strong electrostatic forces)
- **Good conductor of electricity**  
(due to the mobile electrons to carry the current)
- **Malleable and ductile**  
(positive ions are arranged in layers, when it is hit, the layers can slide over and form new bonds easily)



## 3 Stoichiometry

### 3.1 - Formulae

Key definitions:

**Molecular formula** - The number and types of different atoms in one molecule

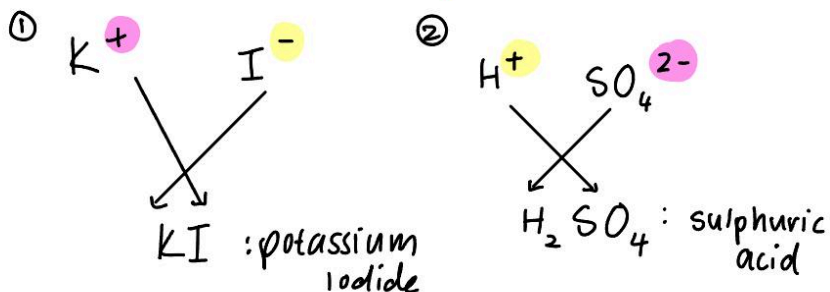
**Empirical formula** - The simplest whole number ratio of different atoms or ions in a compound

Empirical formula:

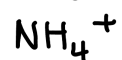
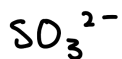
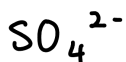
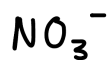
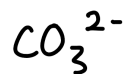


To form an ionic compound's formula,

Cross the ion's charges

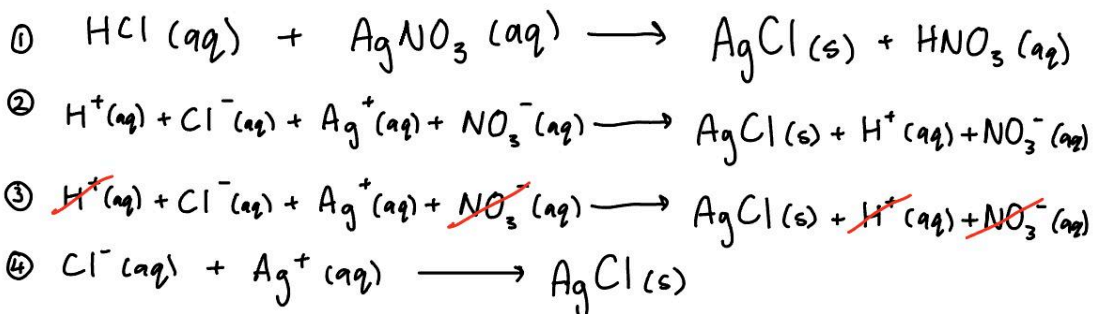


List of complex ions to remember:



How to write ionic equations:

E.g



1. Write out the full equation, including states
2. Split all **aqueous** compounds into its ions
3. Cross out any spectator ions (ions that are in both products and reactants)
4. Rewrite your final equation!

\*to know if a compound is aqueous, refer to the solubility chart (chp 7)

\*all acids are aqueous

### 3.2 - Relative masses of atoms and molecules

*Relative molecular mass:* The sum of the relative atomic masses, used for ionic compounds.  
(relative atomic mass - the average mass of isotopes of n element in comparison to 1/12th of a mass of carbon-12)

### 3.3 - The mole and the Avogadro constant

**Mole** - the unit of amount of substance where one mole contains  $6.02 \times 10^{23}$  particles  
Avogadro's constant =  $6.02 \times 10^{23}$

*Formulas used for mole calculations:*

$$\textcircled{1} \text{ moles} = \frac{\text{mass (g)}}{\text{molecular mass}}$$

$$* \text{ cm}^3 \xrightarrow{\times 1000} \text{ dm}^3$$

$$\textcircled{2} \text{ moles} = \text{mass} \times \text{concentration}$$

$$\textcircled{3} \text{ moles} = \frac{\text{volume}}{24 \text{ dm}^3}$$

OR

$$\text{moles} = \frac{\text{volume}}{24000 \text{ cm}^3}$$

} applicable  
if at  
room temp  
+  
pressure

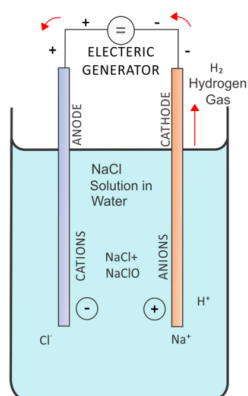
$$\textcircled{4} \text{ moles} = \frac{\text{number of ions/particles}}{6.02 \times 10^{23}}$$

## 4 Electrochemistry

### 4.1 - Electrolysis

The breakdown of an ionic compound in aqueous or molten state through passage of electricity.

An example of an electrolysis setup:



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**Anode** - positive electrode, attracts anions (negative ions), where oxidation takes place

**Cathode** - negative electrode, attract cations (positive ions), where reduction takes place

**Electrolyte** - The molten/aqueous substance that is being used for electrolysis

Why are carbon/graphite electrodes used?

1. Graphite conducts electricity
2. Graphite is inert

#### Molten electrolysis

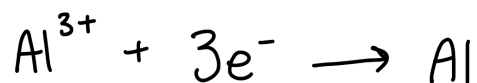
- Metal (cation) and nonmetal (anion) produced

E.g. Aluminium oxide

Anode (oxidation) = Oxygen gas produced



Cathode (reduction) = Aluminium solid formed



*Half equations:*

Equations that show reduction or oxidation

### Aqueous electrolysis rules

*Dilute solution:* OH<sup>-</sup> (oxygen) and H<sup>+</sup> (hydrogen) are produced at anode and cathode respectively

*Concentrated solution:* Halogen produced at the anode (if no halogen present, oxygen is produced), and a metal less reactive than hydrogen at the cathode (if no suitable metal is present, hydrogen is produced)

E.g.

Concentrated hydrochloric acid – Cathode: hydrogen Anode: chlorine

Concentrated aqueous copper nitrate – Cathode: copper Anode: oxygen

Dilute hydrochloric acid – Cathode: hydrogen Anode: oxygen

*Electrolysis of aqueous copper sulphate with...*

**Carbon electrodes (inert electrodes):** concentration of copper in electrolyte reduces, electrolyte **loses colour** (blue → colourless). Ions are taken from the electrolyte.

anode	cathode
$4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

**Copper electrodes (reactive electrodes):** concentration of copper in electrolyte stays the same. **Colour does not change** (stays blue). Ions move from anode to cathode, not from electrolyte. Anode loses mass, cathode gains mass.

anode	cathode
$\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

### Electroplating:

- To enhance appearance
- To prevent corrosion

Example: 1. Electroplating jewellery in silver

**Anode:** metal used for plating. Loses mass

**Cathode:** metal to be plated. Gains the lost mass

Ions move from anode → cathode

#### 4.2 - Hydrogen–oxygen fuel cells

Using hydrogen and oxygen to produce electricity, with water as the only chemical product.

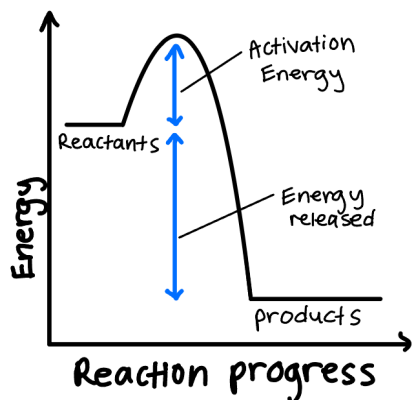
Advantages	Disadvantages
<ul style="list-style-type: none"><li>- Does not have to be recharged</li><li>- No pollution made</li></ul>	<ul style="list-style-type: none"><li>- Flammable</li><li>- It is hard to store gas</li></ul>

## 5 Chemical energetics

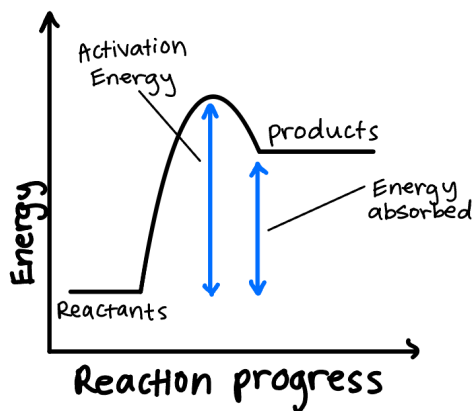
### 5.1 - Exothermic and endothermic reactions

**Activation energy:** The minimum amount of energy necessary for a reaction to occur

**Exothermic reaction:** Transfers energy to the surroundings leading to increase in surrounding temperatures. Bond making.



**Endothermic reaction:** Takes in energy from surroundings leading to decrease in the temperature of surroundings. Bond breaking.



**Enthalpy change:** the transfer of thermal energy during a reaction.

*How to calculate enthalpy change?*

Total energy of reactants – Total energy of products

\*enthalpy change is negative for exothermic reactions and positive for endothermic reactions

## 6 Chemical reactions

### 6.1 - Physical and chemical changes

*Chemical Vs Physical changes:*

Physical	Chemical
<ul style="list-style-type: none"><li>- Can be undone easily</li><li>- No new material formed: chemically the same</li><li>- Change in state</li></ul>	<ul style="list-style-type: none"><li>- Difficult to reverse</li><li>- Colour change</li><li>- Bubbles formed</li><li>- New materials formed</li></ul>

### 6.2 - Rate of reaction

*Factors that increase the rate of reaction:*

- Increasing concentration of the reactants/products
- Increasing the pressure used
- Increasing the temperature used
- Increasing the surface area of the reactants
- Using a catalyst

**Catalyst:** A substance that increases the rate of reaction while remaining unchanged (not used up) at the end of the reaction

*How to measure rate of reaction:*

1. The mass of product produced/time
2. The mass of reactant used/time

#### **Collision theory:**

For a reaction to take place, the particles must:

1. Collide with each other
2. They must collide with enough activation energy *successfully*

**Thus, to increase the chances of successful collisions you can,**

- Increase the number of particles colliding (by increasing the concentration of reactants or increasing pressure)
- Increase the kinetic energy of the particles (by increasing the temperature)
- Increase the frequency of collision (by increasing the surface area of the products)
- Add a catalyst (which would lower the activation energy required of the particles)



### 6.3 - Reversible reactions and equilibrium

A reversible reaction is represented by the symbol:  $\rightleftharpoons$

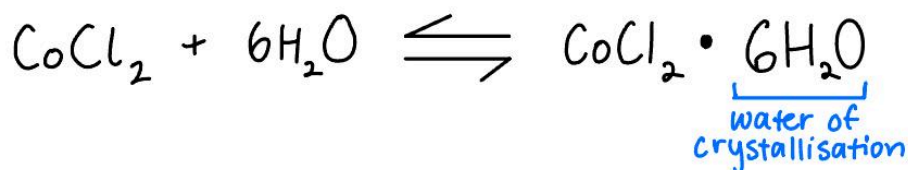
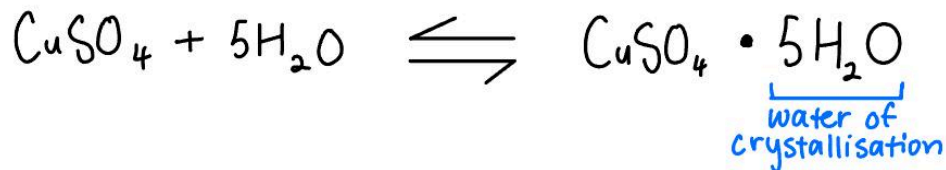
Examples of reversible reactions include hydrated and anhydrous compounds:

- When you add water to anhydrous compounds they become hydrated
- When you heat hydrated compounds they become anhydrous

**Hydrated salt:** chemically combined with H<sub>2</sub>O

**Anhydrous salt:** does not contain H<sub>2</sub>O

Anhydrous copper(II) sulphate  $\rightleftharpoons$  (WHITE) Hydrated copper(II) sulphate (BLUE)  
Anhydrous cobalt(II) chloride (BLUE)  $\rightleftharpoons$  Hydrated cobalt(II) chloride (PINK)



**Water of crystallisation:** The water present in a hydrated compound

**A reaction in equilibrium:** When the forward and the backward reaction are equal in rate, the concentration of reactants and products remain constant, and the reaction happens in a closed system.

**equilibrium shifts** (how factors affect the **yield** of forward and backward reactions):

1. *Increase in temperature*  
Favours side with endothermic reaction
2. *Decrease in temperature*  
Favours side with exothermic reaction.

1. *Increase in pressure*  
Favours side with less moles
2. *decrease in pressure*

Favours side with more moles

1. *Increase in concentration*  
Favours the production of products
2. *Decrease in concentration*  
Favours the production of reactants

1. *Adding a catalyst*

Increases rate of reaction but **does not** affect equilibrium shifts/yield

\*yield and rate of reaction are NOT equivalent

### Processes studied:

Haber process	Contact process
<ul style="list-style-type: none"><li>- Creation of <i>ammonia</i></li><li>- Reacting N<sub>2</sub> and H<sub>2</sub> to produce NH<sub>3</sub>.</li><li>- An exothermic process</li><li>- Nitrogen from air and hydrogen from methane</li></ul> <p><b>N<sub>2</sub> + 3H<sub>2</sub> ⇌ 2NH<sub>3</sub></b></p> <p><b>Conditions:</b></p> <ol style="list-style-type: none"><li>1. 200 atm (20,000 kPa)</li><li>2. iron catalyst</li><li>3. Temp of 450 degrees celsius</li></ol>	<ul style="list-style-type: none"><li>- Making of sulfuric acid</li><li>- Burn sulphur in air or roast sulphur ore to produce sulphur dioxide</li></ul> <ol style="list-style-type: none"><li>1. React with sulphur dioxide with oxygen to produce sulphur trioxide (Reversible part)</li><li>2. Sulphur trioxide react with sulphuric acid to produce Oleum (H<sub>2</sub>S<sub>2</sub>O<sub>7</sub>)</li><li>3. Oleum reacts with water to produce sulphuric acid</li></ol> <p>– have to transfer into oleum because directly it is highly exothermic and the fumes are toxic.</p> <p>S + O<sub>2</sub> → SO<sub>2</sub></p> <p><b>2SO<sub>2</sub> + O<sub>2</sub> ⇌ 2SO<sub>3</sub></b></p> <p>SO<sub>3</sub> + H<sub>2</sub>SO<sub>4</sub> →(Oleum) H<sub>2</sub>S<sub>2</sub>O<sub>7</sub></p> <p>H<sub>2</sub>S<sub>2</sub>O<sub>7</sub> + H<sub>2</sub>O → 2H<sub>2</sub>SO<sub>4</sub></p> <p><b>Conditions:</b></p> <ul style="list-style-type: none"><li>- Vanadium (V) oxide catalyst</li><li>- Temp of 450 degrees celsius</li><li>- 2 atm (200 kPa)</li></ul>

Why aren't **lower temperatures** used since the **reactions are exothermic** (exothermic reactions favour lower temperatures)?

- Below 450 degrees celsius, the rate reaction is too slow
- Above 450 degrees celsius, the yield of forward reaction decreases

Why aren't **higher pressures** used (since higher pressure favours the side with less moles)?

- Higher pressures are dangerous
- Higher pressures are more expensive to maintain

## 6.4 - Redox

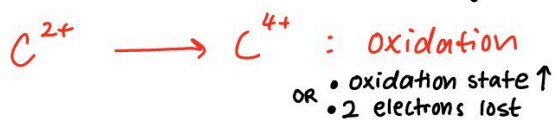
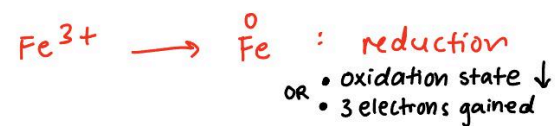
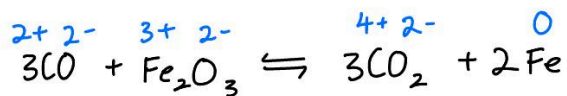
A reaction that undergoes both oxidation AND reduction

What is oxidation and reduction?

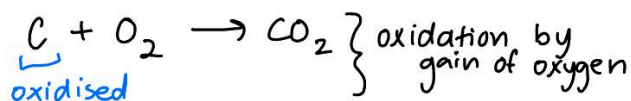
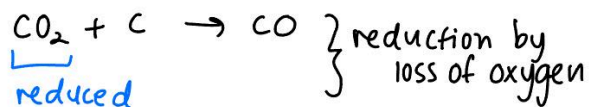
Oxidation	Reduction
<ul style="list-style-type: none"><li>• Loss of hydrogen</li><li>• Gain of oxygen</li><li>• Loss of electrons (OIL)</li></ul> Oxidation is loss of electrons	<ul style="list-style-type: none"><li>• Gain of hydrogen</li><li>• Loss of oxygen</li><li>• Gain of electrons (RIG)</li></ul> Reduction is gain of electrons
<ul style="list-style-type: none"><li>• Increase of oxidation state</li></ul>	<ul style="list-style-type: none"><li>• Decrease in oxidation state</li></ul>

1. Elements not combined (alone) have an oxidation state of 0 e.g H<sub>2</sub>, Mg
2. Oxidation state = charge (it is represented as roman numerals)  
E.g. cobalt(II) has an oxidation state of +2
3. Sum of all oxidation states in a a molecule is 0
4. In complex ions (e.g SO<sub>4</sub>), sum of oxidation numbers equal to charge

Examples:



} redox by change in oxidation state/electrons



**Oxidising agent:** The substance that causes another substance to be oxidised in a reaction. It itself is reduced

**Reducing agent:** The substance that causes another substance to be reduced in a reaction. It itself is oxidised

*Potassium (VII) manganate* : oxidising agent  
Purple to colourless when it is reduced

*Potassium iodide*: reducing agent  
Colourless to yellow/brown when it is oxidised

## 7 Acids, bases and salts

### 7.1 - The characteristic properties of acids and bases

#### Acids:

They are proton donors

- They have a pH of less than 7
- Oxides or hydroxides of nonmetals are acidic
- Aqueous solutions of acids contain H<sup>+</sup> ions

#### Acid reactions:

Metal + acid → salt + hydrogen

Metal Oxide + acid → salt + water (neutralisation)

Metal carbonates + acid → salt + water + carbon dioxide

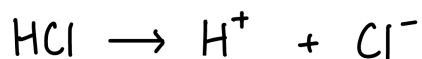
Metal hydroxide + acid → salt + water (neutralisation)

**Weak acid:** partially dissociates in water (e.g ethanoic acid): CH<sub>3</sub>COOH (pH 4-6)

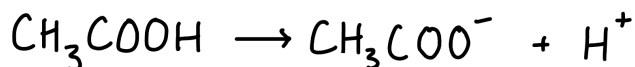
**Strong acid:** fully dissociates with water (means that the hydrogen ions are broken off fully)  
e.g. HCl (pH 1-3)

Dissociation equations:

*Strong acid (hydrochloric acid)*



*Weak acid (ethanoic acid)*



*Identification of acids:*

- Turns damp blue litmus paper red
- Turns thymolphthalein colourless
- Turns methyl orange red

#### Bases:

They are proton acceptors

- They have a pH above 7
- alkalis are soluble bases (all alkalis are bases but not all bases are alkalis)
- Oxides or hydroxides of metals are basic
- Aqueous solutions of alkalis contain OH<sup>-</sup> ions

**Base reactions:**

Base + acid → salt + water

Base + ammonium salt → ammonia + water + salt (decomposition reaction)

E.g. NaOH + (NH<sub>4</sub>)<sub>2</sub>SO<sub>4</sub> → NH<sub>3</sub> + H<sub>2</sub>O + Na<sub>2</sub>SO<sub>4</sub>

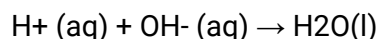
**Weak bases:** partially dissociates in water (pH 8-10)

**Strong bases:** fully dissociates with water (pH 11-14)

*Identification of bases:*

- Turns damp red litmus paper blue
- Thymolphthalein turns blue
- Methyl orange turns yellow

Neutralisation reaction between acids and alkali's ionic equation:



## 7.2 - Oxides

- metallic oxide = base E.g CuO and CaO
- Non metallic oxide = acid E.g SO<sub>2</sub> and CO<sub>2</sub>

**Amphoteric oxide:** Oxides that react with both base and acid to produce salt and water. E.g. Zinc oxide and aluminium oxide

**Neutral oxides:** oxides that do not react with either acid nor base. E.g carbon monoxide and nitric oxide

## 7.3 - Preparation of salts

Solubility rules of salts:

<b>Salts</b>	<b>Soluble</b>	<b>not soluble</b>
--------------	----------------	--------------------

Sodium, potassium and ammonium	All	none
Nitrates	All	None
Chlorides	Most	Silver and lead
Sulphates	Most	Barium, calcium and lead
Carbonates	sodium, potassium and ammonium	Most

Methods of preparing salts:

<b>Insoluble salt prep - soluble + soluble (precipitation)</b>	<b>Soluble salt prep - metal + acid (not containing K, NH<sub>4</sub>, Na)</b>	<b>Soluble salt prep - acid + alkali (titration) salt containing K, NH<sub>4</sub>, Na</b>
<ol style="list-style-type: none"> <li>2 aqueous solutions of soluble substances mixed and reacted</li> <li>Filtrate filtered out</li> <li>Rinse with distilled water</li> <li>Dry residue in oven</li> </ol>	<ol style="list-style-type: none"> <li>React insoluble base and acid</li> <li>Filter out excess solid</li> <li>Heat solution in evaporating dish over a water bath</li> <li>Stop when crystals start to appear around the edge</li> <li>Leave solution at room temp for crystals to form</li> </ol>	<ol style="list-style-type: none"> <li>Prepare acid in a burette</li> <li>Prepare a known amount of alkali in a beaker</li> <li>Add and suitable indicator into alkali</li> <li>Slowly drip the acid into the alkali until it reaches its end point (when the indicator changes colour)</li> <li>Record the volume of titrant used</li> <li>Repeat experiment without the indicator</li> <li>Solution is evaporated and cooled to form crystals</li> </ol>

## 8 The Periodic Table

### 8.1 - Arrangement of elements

The elements in the periodic table are arranged in order of increasing proton number

Group ►	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18														
Period ▼																																
Nonmetals	1 H																	2 He														
Metals	3 Li	4 Be	Transition metals (sometimes excl. group 12)										5 B	6 C	7 N	8 O	9 F	10 Ne														
	11 Na	12 Mg	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr														
	37 Rb	38 Sr	La to Yb		71 Y	72 Zr	73 Nb	74 Mo	75 Tc	76 Ru	77 Rh	78 Pd	79 Ag	80 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe												
	55 Cs	56 Ba	Ac to No		87 Fr	88 Ra	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn										
	s-block (incl. He)		f-block		d-block										p-block (excl. He)																	
Lanthanides	<table border="1"> <tr> <td>57 La</td><td>58 Ce</td><td>59 Pr</td><td>60 Nd</td><td>61 Pm</td><td>62 Sm</td><td>63 Eu</td><td>64 Gd</td><td>65 Tb</td><td>66 Dy</td><td>67 Ho</td><td>68 Er</td><td>69 Tm</td><td>70 Yb</td> </tr> </table>																		57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb
57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb																			
Actinides	<table border="1"> <tr> <td>89 Ac</td><td>90 Th</td><td>91 Pa</td><td>92 U</td><td>93 Np</td><td>94 Pu</td><td>95 Am</td><td>96 Cm</td><td>97 Bk</td><td>98 Cf</td><td>99 Es</td><td>100 Fm</td><td>101 Md</td><td>102 No</td> </tr> </table>																		89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No
89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No																			

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- Metals are found to the left of the periodic table
- Non metals are found to the right of the periodic table
- Members of the same group have the same electron valency
- Elements of the same group have the same chemical properties (due to same electron valency)
- The group number of an element also determines its ionic charge E.g Group 2 has charge +2
- The period of a group represents the number of shells an element occupies

### 8.2 - Group I properties

- They are soft metals that can be cut with a knife
- They are known as alkali metals
- They react with water to form an alkali solution

As you go down the group:

- Melting point decreases
- Density increases
- Reactivity increases

### 8.3 - Group VII properties



- They exist as diatomic molecules E.g Br<sub>2</sub>, Cl<sub>2</sub>
- They are known as halogens
- More reactive halogens can displace less reactive halogens in a reaction

As you go down the group:

1. Density increases
2. Reactivity increases

Appearance of halogens:

The colour gets darker down the group

<b>Chlorine</b>	Pale yellow-green gas
<b>Bromine</b>	Red-brown liquid
<b>Iodine</b>	Grey-black solid

Displacement chart of halogens: more reactive halogens can displace less reactive halogens in a reaction

	<b>Potassium chloride</b>	<b>Potassium bromide</b>	<b>Potassium iodide</b>
<b>Chlorine</b>	x	Potassium chloride and Bromine produced	Potassium chloride and Iodine produced
<b>Bromine</b>	No reaction	x	Potassium bromide and Iodine produced
<b>Iodine</b>	No reaction	No reaction	x

#### 8.4 - Transition elements

Found around the middle of the periodic table

Properties:

- High density
- High melting and boiling point
- Form coloured compounds
- They act as catalysts
- Their ions have variable oxidation states E.g Iron(II) and Iron(III)

#### 8.5 - Noble gases (group VIII)

- Noble gases have full electron shells so they are inert (they do not react) and stable
- They exist as monatomic gases E.g. Ar, He

## 9 Metals

### 9.1 - Properties of metals

1. Metals are conductors of heat and electricity
2. they are malleable and ductile
3. High melting and boiling points

#### Metal reactions:

Metal + acid → salt + hydrogen

Metal + cold water → metal hydroxide + hydrogen

Metal + steam → metal oxide + hydrogen

Metal + oxygen → metal oxide

### 9.2 - Uses of metals

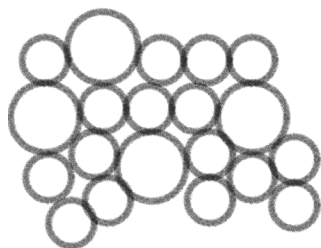
1. Aluminium
  - Used in aircrafts as it has a low density
  - It is used in electrical cables due to its low density and conductivity
  - Used in food storage as it has corrosion resistance
2. Copper
  - Used in electrical wiring due to copper being ductile and a conductor

### 9.3 - Alloys and their properties

Alloys are mixtures of metals

E.g.

1. Brass; mixture of zinc and copper
  2. Stainless steel; mixture of iron, carbon, chromium and nickel
- Alloys are often stronger than pure metals due to the differently sized atoms, that stops the layers from sliding over each other



\*stainless steel is used in cutlery due to its rust resistance

#### 9.4 - Reactivity series

Most reactive

1. Potassium K
2. Sodium Na
3. Lithium Li
4. Calcium Ca
5. Magnesium Mg
6. Aluminium Al: seems less reactive due to the oxide layer it has
7. Carbon C
8. zinc Zn
9. Iron Fe
10. Hydrogen H
11. copper Cu
12. silver Ag
13. Gold Au

Least reactive

More reactive metals can displace less reactive metals in a reaction.

*Reactivity series reactions:*

	<b>Steam</b>	<b>Water</b>	<b>Acid</b>
Potassium K	Reacts with steam vigorously to form hydroxide	Reacts with cold water vigorously to form hydroxide	Reacts explosively to form hydrogen gas
Sodium Na			
Lithium Li			
Calcium Ca			

Magnesium Mg	Reacts with steam vigorously to form hydroxide	Reacts slowly to form hydroxide	Reacts to form hydrogen
zinc Zn Iron Fe	Reacts slowly to form hydroxide	Reacts slowly to form hydroxide	Reacts slowly to form hydrogen
copper Cu silver Ag Gold Au	Does not react	does not react	does not react

### 9.5 - Corrosion of metals

**Sacrificial protection:** The reaction of a more reactive metal in place of another metal to protect it from corrosion

E.g. Galvanising iron using zinc

- Zinc is more reactive than iron. Reacts instead of iron. Zinc oxidises (loses electrons)

When Iron is in the presence of **water** and **oxygen**, RUST is formed (hydrated iron(III) oxide)

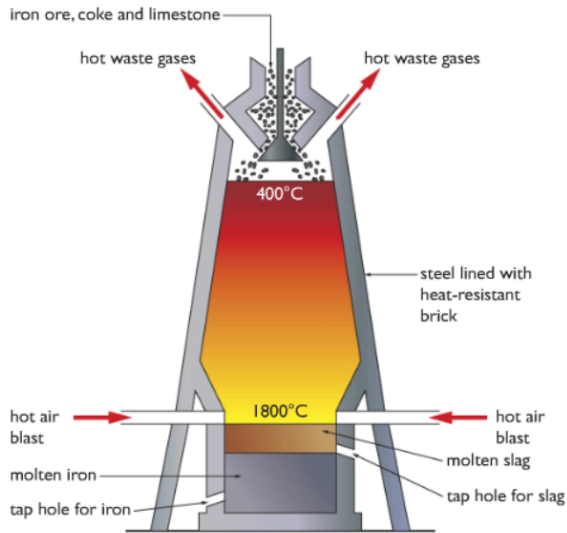
**rust prevention:** painting, greasing, coating with plastic, galvanising

### 9.6 - Extraction of metals

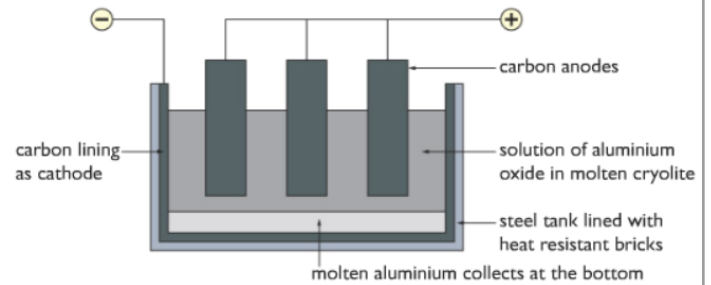
2 processes studied:

Extraction of Iron (blast furnace)	Extraction of Aluminium oxide (electrolysis)
------------------------------------	--

- Substances added: (iron ore) Hematite, coke, limestone, air
- $C + O_2 \rightarrow CO_2$  (produces heat)
- $CO_2 + C \rightarrow CO$  (reduction)
- $3CO + Fe_2O_3 \rightleftharpoons 3CO_2$  and  $2Fe$  (redox)
- $CaCO_3 \rightarrow CaO + O_2$  (thermal decomposition)
- $CaO + SiO_2 \rightarrow CaSiO_3$  (neutralisation: removes impurities - produces **slag**)



- Substances added: (aluminium ore) Bauxite purified to  $Al_2O_3$  and Cryolite
- $Al^{3+} + 3e^- \rightarrow Al$  (cathode)
- $2O_2 \rightarrow O_2 + 4e^-$  (anode)
- Anode needs to be replaced continuously as carbon reacts with the  $O_2$  produced to produce  $CO_2$
- Cryolite usage:
  1. Used as solvent
  2. Lowers working temp
  3. Increases conductivity of electrolysis



## 10 Chemistry of the environment

### 10.1 - Water

#### Test for water:

1. Cobalt(II) chloride: blue to pink
2. Copper(II) sulphate: white to blue

**Test for water purity:** Heat to boiling point (pure water boils at 100 degrees celsius)

- Distilled water is used in experiments as they have the least impurities

*What is found in natural water?*

- Dissolved oxygen (good for aquatic life)
- Metal compounds
- Plastic
- Sewage
- Harmful microbes
- Nitrates and phosphates from fertilisers (deoxygenation of water due to eutrophication)

*How is water treated?*

- Treated water is used for: drinking
1. Screen filter (large particles removed)
  2. Coagulate and skim off (removal of smaller insoluble particles)
  3. Filtration (small particles)
  4. Use of carbon to remove taste and odour
  5. Chlorination to kill pathogens

### 10.2 - Fertilisers

Fertiliser used to improve plant growth and health:

- NPK fertilisers – Nitrogen, phosphorus, potassium for plant growth
- Ammonium (NH<sub>4</sub>) salts

### 10.3 - Air quality and climate

#### Air composition -

78% Nitrogen

21% Oxygen

0.04% CO<sub>2</sub>

<1% water vapour and other noble gases

**Pollutants found in air:**

	<b>Source</b>	<b>Affect</b>
<b>CO<sub>2</sub> (also found in clean air)</b>	Combustion	Global warming + climate change
<b>CO</b>	Incomplete combustion	Toxic gas Acid rain
<b>CH<sub>4</sub> (methane)</b>	Decomposition of vegetation Waste gases from animals	Global warming + climate change
<b>Oxides of nitrogen</b>	High temp from car engines react O <sub>2</sub> and N <sub>2</sub>	Acid rain, photochemical smog and respiratory issues
<b>Sulphur dioxide</b>	Sulphur impurities in crude oil and combustion of fossil fuels	Acid rain

*Greenhouse gases: CO<sub>2</sub> and CH<sub>4</sub> –*

- They absorb heat energy from the earth, and re-emits it back into our atmosphere
- Heat is trapped instead of released into space, causing **global warming**

**How to reduce the effects of pollutants:**

1. Plant trees (reduces CO<sub>2</sub> through plants photosynthesizing)
2. Reduce livestock (reduce methane)
3. Decrease fossil fuel use (reduce CO<sub>2</sub>)
4. Use hydrogen and renewable energy (reduce CO<sub>2</sub>)
5. Flue gas desulfurization using calcium oxide (reduces creation of sulphur dioxide)
6. Use catalytic converters (reverses the creation of CO and NO)

The use of catalytic converters in an engine:  $2\text{CO} + 2\text{NO} \rightarrow 2\text{CO}_2 + \text{N}_2$

*Photosynthesis in plants:*

Carbon dioxide + water → glucose + oxygen

Conditions:

- Light
- Chlorophyll

## 11 Organic chemistry

### 11.1 - Formulae, functional groups and terminology

*Key definitions:*

**Saturated compound:** A compound where all carbon-carbon bonds are single

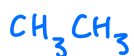
**Unsaturated compound:** A compound where at least one carbon-carbon bond is not single (alkenes)

**Functional group:** An atom or group of atoms that determine the chemical properties of a homologous series

**Homologous series:** a family of similar compounds with the same chemical properties due to the same functional group

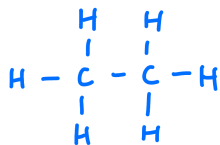
**Condensed formula:** The formula that shows the way a molecule's atoms are arranged

ethane :



**Displayed formula:** The drawn out structure of the molecule and its arranged atoms

ethane :



*Members of the same homologous series have:*

1. Same chemical properties
2. Trend in physical properties
3. Same functional group
4. Same general formula
5. Differ from one member to another by one -CH<sub>2</sub>- unit

Homologous series studied:

1. Alkanes
2. Alkenes
3. Alcohols
4. Carboxylic acids



Homologous series	Alkane	Alkene	Alcohol	Carboxylic acid
General formula	$C_nH_{2n+2}$	$C_nH_{2n}$	$C_nH_{2n+1}OH$	$C_nH_{2n+1}COOH$
Functional group	C-C	C double bond C	OH	COOH

E.g.

An alkane with 2 carbon's formula is  $C_2H_6$  (ethane)

An alcohol with 4 carbon's formula is  $C_4H_9OH$  (butanol)

### 11.2 - Naming organic compounds

Number of carbons	Prefix
1	meth-
2	eth-
3	prop-
4	but-

Homologous series	Suffix
Alkane	-ane
Alkene	-ene
Alcohol	-ol
Carboxylic acid	-oic acid

E.g.

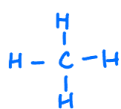
An alkene with 3 carbons = propene

A carboxylic acid with 2 carbons = ethanoic acid

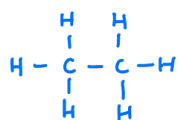
All displayed formulas:

**Alkanes:**

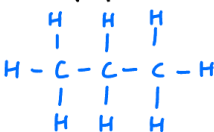
methane



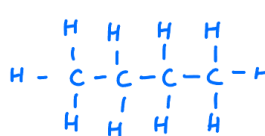
ethane



propane

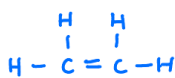


butane



**Alkenes:**

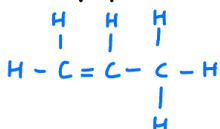
ethene



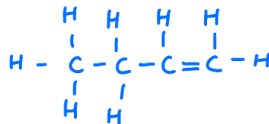
no H?

Carbon can only form 4 bonds.

propene

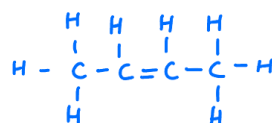


but-1-ene



why not  
but-3-ene?  
Start counting from the carbon closest to the double bond

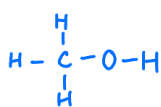
but-2-ene



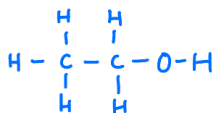
isomers

## Alcohols :

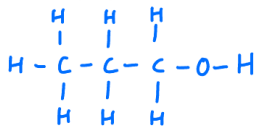
methanol



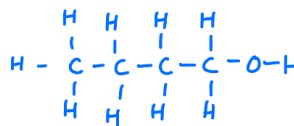
ethanol



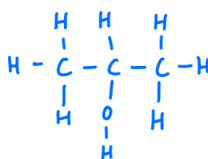
propan-1-ol



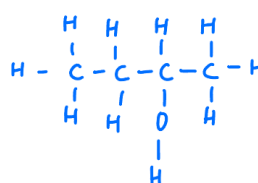
butan-1-ol



propan-2-ol



butan-2-ol

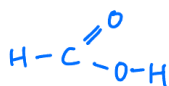


isomers

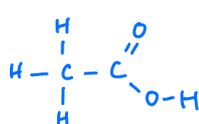
isomers

## carboxylic acids:

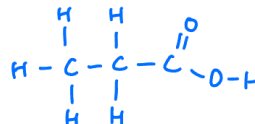
methanoic acid



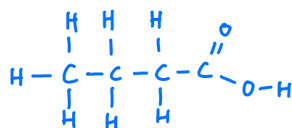
ethanoic acid



propanoic acid



Butanoic acid



### \*Alkenes:

- The number represents the location of the double carbon bond

### \*Alcohols:

- The number represents the location of the -OH bond

**Structural isomers:** compounds with the same molecular formula but different structural formula

## 11.3 - Fuels

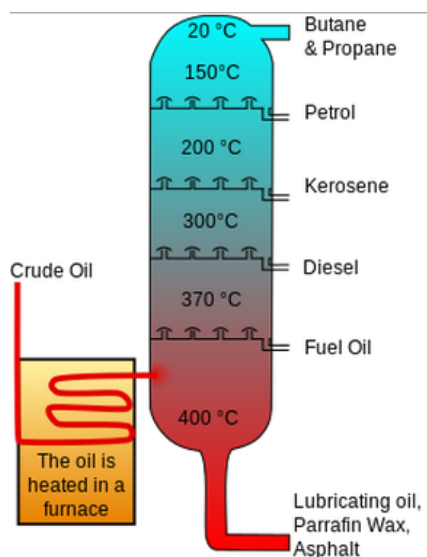
**Hydrocarbons:** compounds that contain hydrogen and carbon ONLY

**Petroleum:** a mixture of hydrocarbons

**Fossil fuels:** Coal, Natural gas and petroleum

\*Methane is mainly made out of natural gas

Petroleum can be separated into different products with different uses. This is done using fractional distillation:



Wikimedia commons

## 11.4 - Alkanes

General formula:  $C_nH_{2n+2}$

- They are saturated hydrocarbons (single C-C bonds)

### Reactions of alkanes:

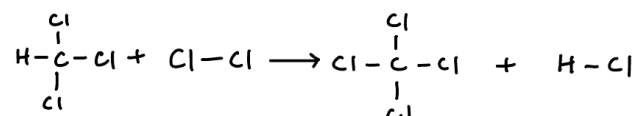
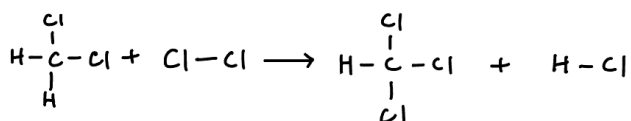
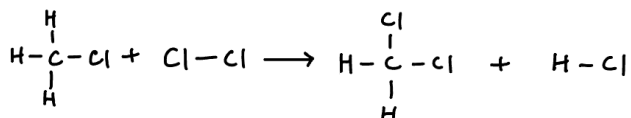
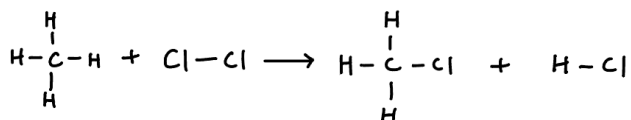
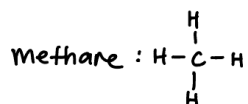
- generally unreactive except:

**Combustion** AND/OR

**Substitution by chlorine** (photochemical reaction, UV light needed to provide activation energy)

Substitution: where an atom is replaced by another atom

*Substitution by chlorine in methane:*



## 11.5 - Alkenes

General formula:  $\text{C}_n\text{H}_{2n}$

- They are unsaturated hydrocarbons

*How are they formed?*

- Cracking

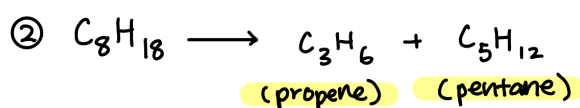
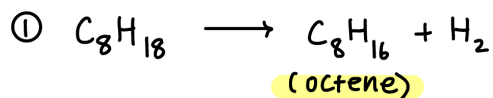
Taking a long chain alkane and heating it to form the products of

1. An alkene and hydrogen

OR

2. An alkene and an alkane

Cracking of Octane:



The equation should be balanced

E.g.

## Reactions of alkenes:

### Alkenes

#### 1. Test for unsaturated hydrocarbons (alkenes):

Add bromine water (turns orange brown to colourless)

2. **Addition reaction:** adding on new atoms to the molecule  
double bond is broken to free up bonds for the addition

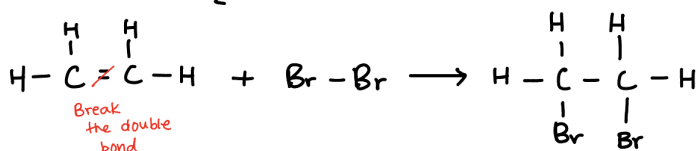
- Element is added

\*The addition of hydrogen turns an alkene back to an alkane (this requires a NICKEL catalyst)

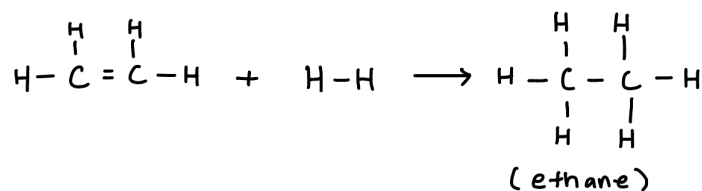
\*The addition of oxygen and hydrogen (steam) turns an alkene to an alcohol (this requires a PHOSPHORIC ACID catalyst)

#### Examples: (using ethene)

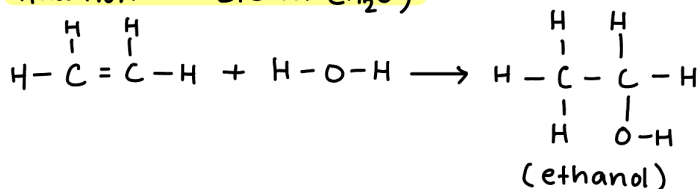
##### Addition of Br<sub>2</sub>



##### Addition of H<sub>2</sub>



##### Addition of Steam (H<sub>2</sub>O)



## 11.6 - Alcohols

General formula:  $C_nH_{2n+1}OH$

*How are they formed?*

1. Addition reaction of alkenes with steam (catalytic addition; a large scale process)
2. Fermentation of aqueous glucose

<b>Fermentation</b>	<b>Catalytic addition of steam to ethene</b>
<p><i>Conditions:</i></p> <ol style="list-style-type: none"><li>1. 25-35 degrees celsius</li><li>2. Yeast</li><li>3. Anaerobic respiration; no oxygen</li></ol> <p>It is a batch process</p> <ul style="list-style-type: none"><li>+ The reactants are renewable</li><li>+ Does not require a lot of energy</li><li>- It is a slow process</li><li>- The alcohol produced is not pure and must be separated</li></ul>	<p><i>Conditions:</i></p> <ol style="list-style-type: none"><li>1. 300 degrees celsius</li><li>2. 60 atm (6000 kPa)</li><li>3. Phosphoric acid catalyst</li></ol> <ul style="list-style-type: none"><li>+ It is a continuous process</li><li>+ The alcohol produced is pure</li><li>- It is expensive</li><li>- It is not renewable (requires ethene, sourced from petroleum)</li></ul>

### Reactions of alcohols:

Combustion

*Uses of ethanol:*

1. As a solvent
2. As a biofuel

## 11.7 - Carboxylic acids

General formula:  $C_nH_{2n+1}COOH$

*How are they formed?*

1. Oxidising an alcohol with an oxidising agent (E.g Potassium manganate(VII))
2. Bacterial oxidation of an alcohol in the production of vinegar

### Reactions of carboxylic acids:

1. Reacts like a normal acid (weak acid; pH 4-6)

Metal + acid → salt + hydrogen

Metal Oxide + acid → salt + water (neutralisation)

Metal carbonates + acid → salt + water + carbon dioxide

Metal hydroxide + acid → salt + water (neutralisation)

2. Reacts with alcohol using sulfuric acid to form an ester

- Esters are sweet smelling liquids often used in flavourings and perfumes

### Naming an ester:

\_\_yl \_\_ate

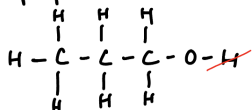
(name of alcohol)yl (name of carboxylic acid)ate

E.g.

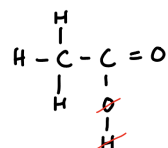
Methanol + Propanoic acid → Methyl propanoate

Propanol + Ethanoic acid → Propyl ethanoate

propanol :

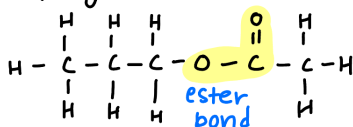


Ethanoic Acid :



- Carboxylic acid loses OH
- Alcohol loses H

propyl ethanoate :



## 11.8 - Polymers

Polymers are large molecules formed by the joining of many monomers

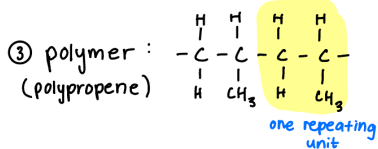
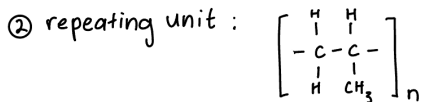
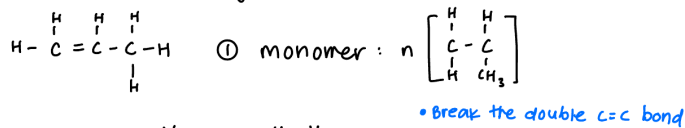
E.g. the monomer propene turns into the polymer **polypropene**

There are 2 methods of forming polymers:

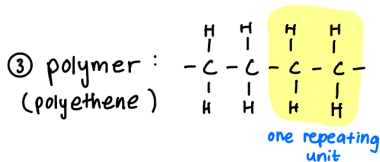
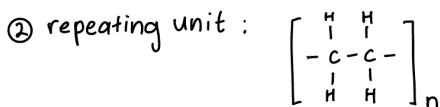
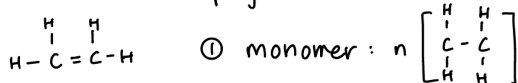
Addition polymerisation	Condensation polymerisation
<ul style="list-style-type: none"><li>- only for alkenes</li></ul> Monomer → repeating unit → polymer  E.g.	Monomer → repeating unit → polymer  There are 2 products of this polymerisation <ol style="list-style-type: none"><li>1. Polyesters: made with diols and dicarboxylic acids</li><li>2. Polyamides: made with diamides and</li></ol>



propene → polypropene



ethene → polyethene



dicarboxylic acids

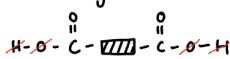
**Polyamide:**

- Have amide linkages

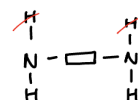
A synthetic polyamide includes

Nylon

dicarboxylic acid

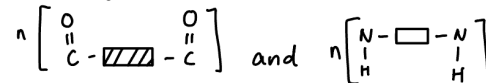


diamine

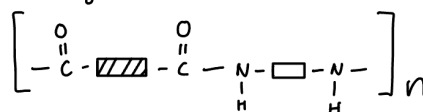


H<sub>2</sub>O formed

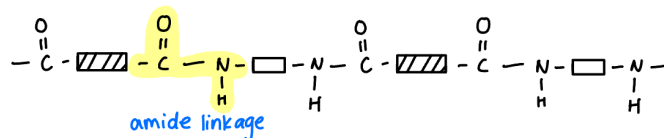
monomers:



repeating unit:



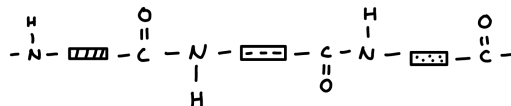
polymer (nylon):



A natural polyamide includes

Proteins

Made of the monomer amino acids

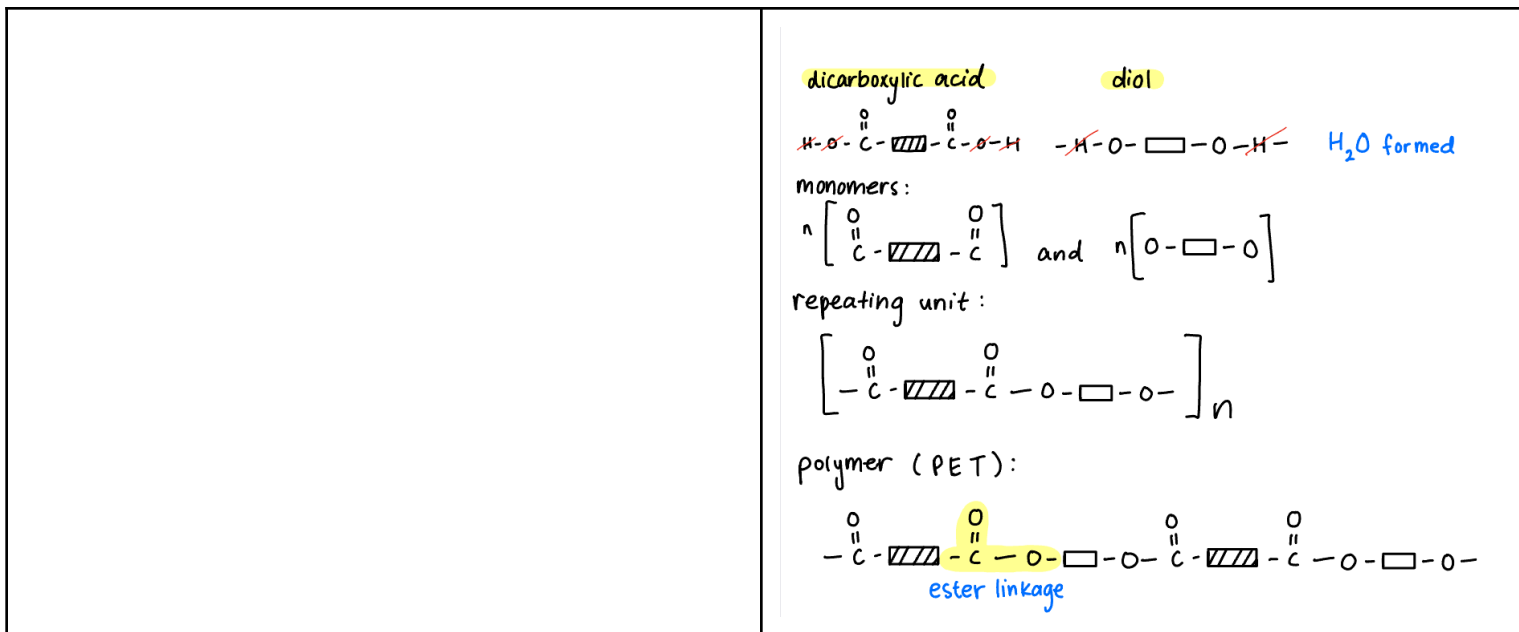


**Polyesters:**

- Have ester linkages

A synthetic polyamide includes

PET (terylene)



**Difference between condensation and addition polymerisation:**

- Addition polymerisation involves one monomer repeated, and only the polymer forms (one product)
- Condensation involves multiple monomers repeated, a small molecule is removed as a byproduct (usually water, so more than one product)

**Plastics:**

- made from polymers

**Implications for their disposals:**

1. Accumulation in landfills  
Takes a long time to biodegrade, smelly
  
2. Combustion of plastics  
Releases toxic gases
  
3. Water pollution  
Kills aquatic life and releases toxins

## 12 Experimental techniques and chemical analysis

### 12.1 - Experimental design

Apparatus:

- Stopwatches - measuring time
- Thermometers - measuring temperature
- Balances - measuring mass
- Burretes - measuring liquid volume
- Volumetric pipettes - measuring a **fixed** volume of liquid
- Measuring cylinders - measuring liquid volume
- Gas syringe - measuring gas volume

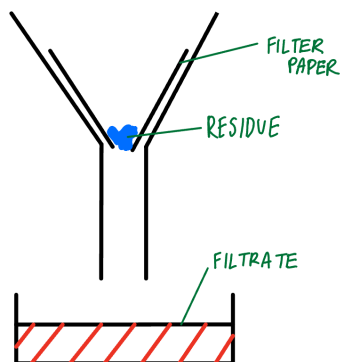
**Solvent:** dissolves a solute

**Solute:** the substance getting dissolved

**Saturated solution:** solution with the maximum concentration of dissolved solute at a temperature

**Residue:** the substance left after filtration of a solution

**Filtrate:** the substance produced after successfully passing through the filter



### 12.2 - Acid-base titrations

Titration is the neutralisation of a soluble base (potassium, ammonium and sodium hydroxides) and acid

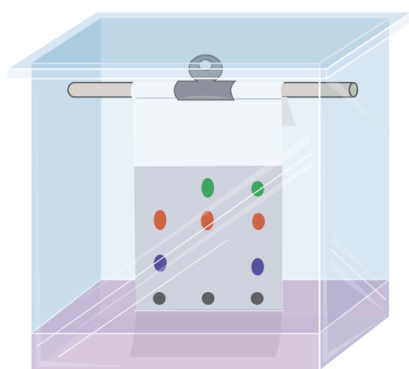
Equipment:

- Burette
- Volumetric pipette
- Indicator

1. Fill a burette with acid
2. Fill a conical flask with a known volume of an alkali (using volumetric palette)
3. Add a few drops of indicator into the alkali (e.g thymolphthalein)
4. Drop the acid from the burette into the alkali, swirling the solution as you go
5. Stop when the colour of the indicator changes (it has reached its end-point)
6. Record the volume of acid used
7. Repeat the experiment without the indicator, using the known volume of acid

### 12.3 - Chromatography

Chromatography is used to separate soluble inks and smaller molecules (e.g amino acids)



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Rf value equation:

Distance travelled by substance/Distance travelled by solvent

Rf value < 1

- If chromatography is used to separate colourless substances, a **locating agent** is required to identify where the substance has travelled to

### 12.4 - Separation and purification

Methods:

- Filtration
- Crystallisation
- Simple distillation
- Fractional distillation

\*distillation is a separation technique that takes advantage of the different boiling points of substances

***If the wanted substance is insoluble:***

- filter

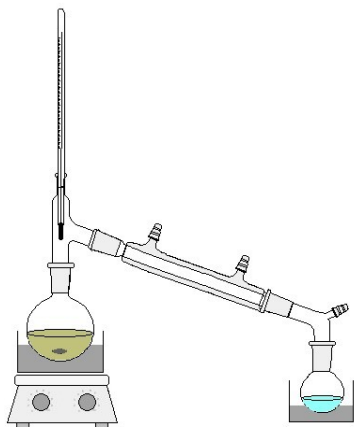
- wash the residue
- dry the residue in an oven

**If the wanted substance is soluble:**

- filter
- crystallise the filtrate

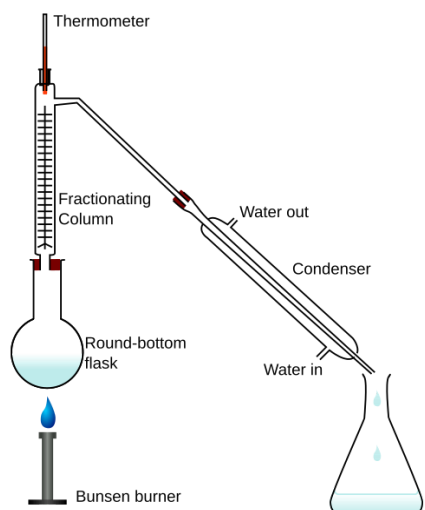
**Separation of 2 liquids:**

- simple distillation



**Separation of liquids with similar boiling points:**

- fractional distillation



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Uses a fractionating column to separate the liquids

## 12.5 - Identification of ions and gases

### Anion tests

	test	result
<b>Carbonate</b>	React with dilute acid then test for carbon dioxide using limewater	Bubbling and limewater turns cloudy/milky
<b>Chloride</b>	Acidify with dilute nitric acid and add aqueous silver nitrate	White precipitate forms
<b>Bromide</b>	Acidify with dilute nitric acid and add aqueous silver nitrate	Cream precipitate forms
<b>Iodide</b>	Acidify with dilute nitric acid and add aqueous silver nitrate	Yellow precipitate forms
<b>Nitrate</b>	Add aqueous sodium hydroxide and then aluminium foil. Warm the mixture.	Bubbling, gas produced that turns damp red litmus paper blue (ammonia)
<b>Sulfate</b>	Acidify with dilute nitric acid and add aqueous barium nitrate	White precipitate formed
<b>Sulfite</b>	Add acidified aqueous potassium manganate (VII)	The potassium manganate (VII) changes from purple to colourless

### Cation tests

	Aqueous sodium hydroxide	Aqueous ammonia
<b>Aluminium</b>	White precipitate formed, soluble in excess forming a colourless solution	White precipitate formed, insoluble in excess
<b>Ammonium</b>	On warming, a gas is produced that turns damp red litmus paper blue (ammonia)	No reaction
<b>Calcium</b>	White precipitate formed, insoluble in excess	No precipitate/very little white precipitate
<b>Chromium(III)</b>	Green precipitate formed, soluble in excess	Green precipitate, insoluble in excess
<b>Copper(II)</b>	Light blue precipitate formed, insoluble in excess	Light blue precipitate, soluble in excess forming a dark blue

		solution
<b>Iron(II)</b>	Green precipitate formed, insoluble in excess	Green precipitate formed, insoluble in excess
<b>Iron(III)</b>	Red-brown precipitate formed, insoluble in excess	Red-brown precipitate formed, insoluble in excess
<b>Zinc</b>	White precipitate formed, soluble in excess forming a colourless solution	White precipitate formed, soluble in excess forming a colourless solution

### Test for gases

	test	result
<b>Ammonia</b>	Damp red litmus paper	Turns blue
<b>Carbon dioxide</b>	Limewater	Turns milky
<b>Chlorine</b>	Damp litmus paper	Bleaches
<b>Hydrogen</b>	Lighted splint	Goes off with a squeaky pop
<b>Oxygen</b>	Glowing splint	Relights
<b>Sulphur dioxide</b>	Bubble through acidified potassium manganate (VII)	Changes from purple to colourless

### Flame test to identify cations:

Using a nichrome wire loop, dip it into some HCl and hold it to the flame, this cleans the loop. Dip the loop into the cation sample and hold it to a blue flame. Observe the colour changes.

<b>Lithium</b>	Red flame
<b>Sodium</b>	Yellow flame
<b>Potassium</b>	Lilac flame
<b>Calcium</b>	Orange-red flame
<b>Barium</b>	Yellow-green flame
<b>Copper</b>	Blue-green flame

Notes by: Gisele