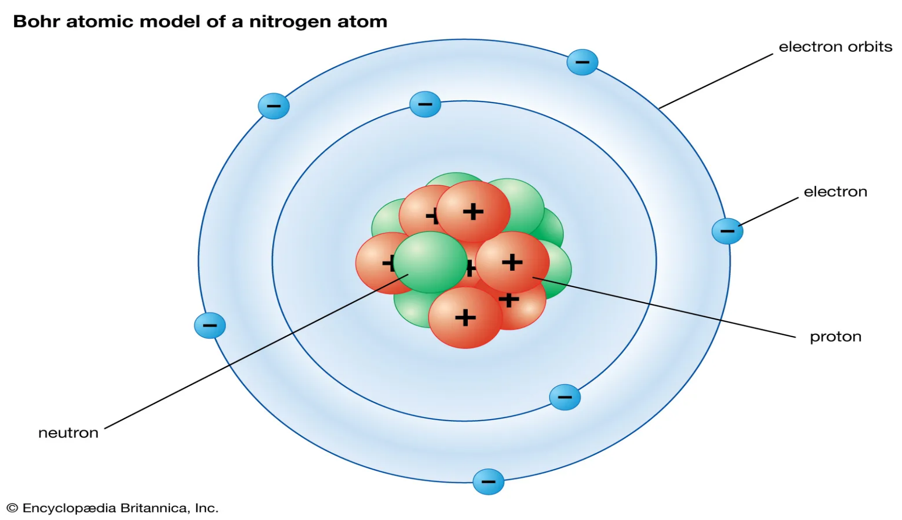
C1 and C3 – Atomic Structure



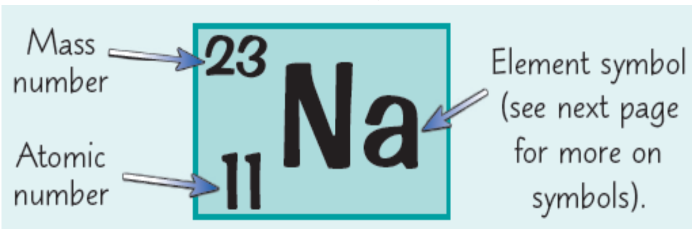
Particle	Relative Mass	Charge
proton	1	+1
neutron	1	0
electron	Very small	-1

Atomic Structure Key Points

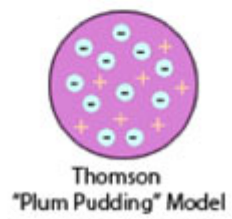
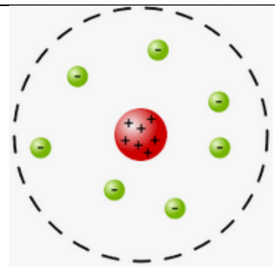
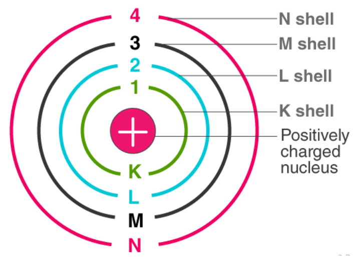
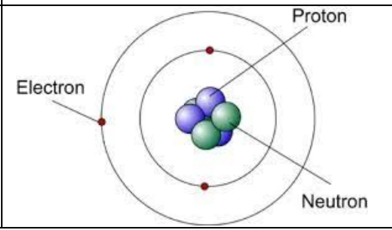
- Positive protons and neutral neutrons are found in a tiny nucleus in the centre with negative electrons orbiting in shells (energy levels around the nucleus
- Electrons have a very small mass compared to protons and neutrons
- Atoms have no overall charge as they have the same number of positive protons and negative electrons
- Size of atoms is approximately 0.1nm (1×10^{-10}m)
- Nucleus is 10,000 times smaller than the atom

Atomic Number and Mass Number

- Atomic number is the number of protons
- Mass number is the number of protons and neutrons

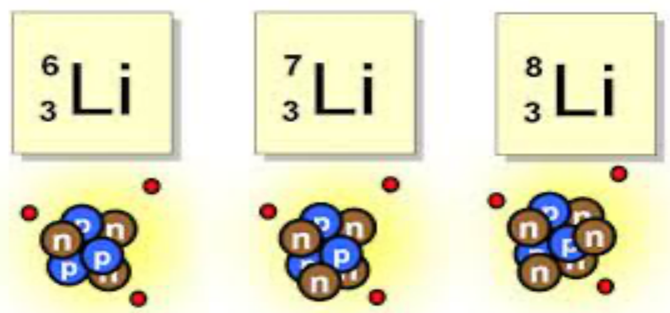


- Number of protons = atomic number e.g. 11
- Number of electrons = atomic number e.g. 11
- Number of neutrons = mass number – atomic number e.g.12

Name of Scientist and Year	Picture and Name of Model	Key Discoveries/Theory
John Dalton (1803)		Idea that atoms were solid spheres that could not be divided any further
JJ Thomson (1897)		Sphere of positive charge with negative electrons scattered within it. Disproves idea of Dalton's model of atoms being indivisible
Ernest Rutherford (1909)	 Rutherford's Nuclear Model	Tiny positive nucleus at the centre with negative electrons orbiting around it
Neils Bohr (1911)	 Bohr's Nuclear Model	Tiny positive nucleus with negative electrons orbiting in shells (energy levels) around the nucleus
James Chadwick (1932)		Discovered the neutron. Explains isotopes

Isotopes

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



Relative Formula Mass (M_r)

The relative formula mass (M_r) is the sum of all the relative atomic masses (A_r) of the atoms in the formula.

Example – H_2SO_4

Hydrogen → 2 x 1 = 2
Sulfur → 1 x 32 = 32
Oxygen → 4 x 16 = 64
Total M_r = 98

Example – $Ca(NO_3)_2$

Calcium → 1 x 40 = 40
Nitrogen → 2 x 14 = 28
Oxygen → 6 x 16 = 96
Total M_r = 164

Calculating Percentage by Mass

Example – Calculate the percentage by mass of potassium (K) in K_2O
Step 1 – Work out the total M_r first
Potassium → 2 x 39 = 78
Oxygen → 1 x 16 = 16
Total M_r = 94

Step 2 – Work out the percentage of potassium out of the total
(78/94) x 100 = 83%

Calculating Relative Atomic Mass

Relative Atomic Mass is the average mass number of all the isotopes of an element found on Earth (or a sample)

Example

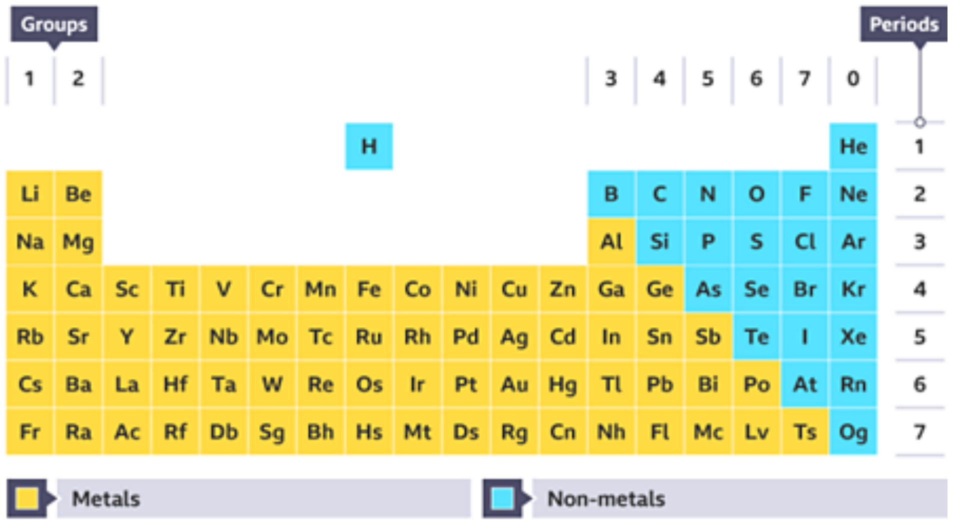
Calculate the relative atomic mass of a sample of Titanium that contains 20% Titanium-46, 20% titanium-47 and 60% Titanium-48

Step 1 – Multiply mass numbers by percentages
Titanium-46 → 46 x 20 = 920
Titanium-47 → 47 x 20 = 940
Titanium-48 → 48 x 60 = 2880

Step 2 – Add all answers together
920 + 940 + 2880 = 4740

Step 3 – Divide by 100
4740 ÷ 100 = 47.4

C1 – Periodic Table



Metals and Non-metals

They are found at the **left** part of the periodic table. Non-metals are at the **right** of the table.

Metals

Are strong, malleable, good conductors of electricity and heat. They bond metallically.

Non-Metals

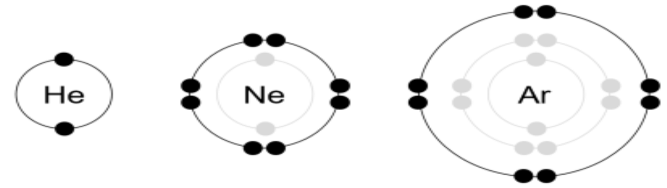
Are dull, brittle, and not always solids at room temperature.

Groups 0 – Noble Gases

Group 0 elements are found on the right hand of the periodic table and either have 2 or 8 electrons in the outer shell.

Properties

- They are all gases with very low boiling points
- They are unreactive – full outer shell, so do not need to gain or lose electrons
- They are colourless gases



Trends

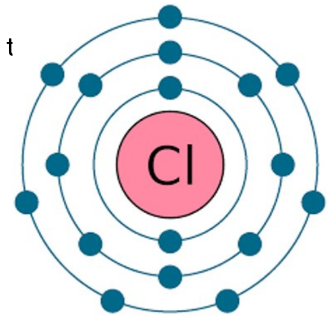
- Boiling point increases down the group (bigger atom, so stronger intermolecular forces)
- Density increases down the group

Groups and Periods

The modern periodic table is arranged by atomic (proton) number.

A group is a column on the periodic table and tells you the number of electrons in the outer shell of the atom. Elements in the same group will have the same number of electrons in the outer and also will have similar physical and chemical peroperties

A period is a row on the periodic table and t electron shell in the atom



For example, chlorine on the right has 7 electrons in the outer shell, so has 7 electrons in the outer shell. It also has 3 electron shells, so it is period 3

Group 7 – Halogens

Group 7 elements have 7 electrons in the outer shell, so will react to gain one electron to have a full outer shell.

Properties

- Coloured gases
- Poisonous and smelly
- Very reactive
- For -1 ions

Trends

- Boiling point increases down the group (bigger atom, so stronger intermolecular forces)
- Colour gets darker down group (chlorine – pale yellow gas, bromine – red liquid, iodine – grey solid)
- Gets less reactive down the group – more shells, bigger distance between positive nucleus and incoming electron, so weaker attraction and harder to gain electron

Group 1 Alkali Metals

Group 1 elements are found on the left hand side of the periodic table and all have 1 electron in the outer to shell. This means they all react to lose 1 electron to have a full outer shell

Properties

- Soft – can be cut with a knife
- Low density – floats on water
- Low melting points for metals
- Dull – reacts with oxygen in the air
- Very reactive

Trends

- Melting point decreases down the group (bigger atom, so stronger intermolecular forces)
- Gets more reactive down the group - more shells, bigger distance between positive nucleus and outer electron, so weaker attraction and easier to lose outer electron

Reactions with Water

They are called alkali metals as they all react with water to produce a metal hydroxide, which is an alkali. They all also fizz within the water (hydrogen released), move around the surface of the water and turn into a molten ball. Potassium will also set alight into a lilac flame

Development of the Periodic Table

Newlands and Mendeleev both arranged the elements by atomic weight. No protons discovered at the time, so could not be in order of atomic number.

Newland's realised that every 8^{th} element had similar properties, and therefore his theory was called Law of Octaves. However, he ended up putting elements such as oxygen and iron together when they do not have similar properties. Therefore. His periodic table was accepted.

Mendeleev also arranged the elements by atomic weight, but did 3 things different that allowed his ideas to become accepted

- Left gaps for undiscovered elements, and later these elements got discovered and fit into these gaps
- Mendeleev predicted the properties of these future elements, and he was correct in terms of his predictions
- Mendeleev switch some elements around (like Iodine and Tellurium) so they fit into groups with similar properties

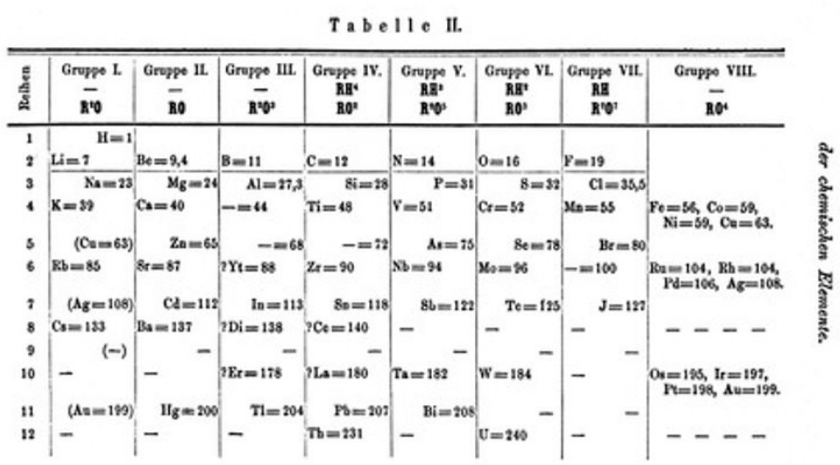


Tabelle II.

Reihen	Gruppe I. — R^2O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

der chemischen Elemente.

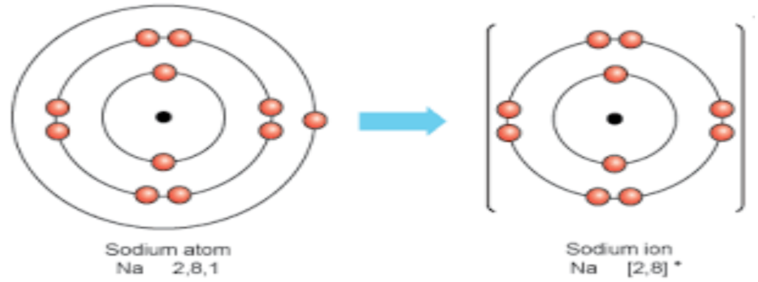
C2 – Ionic and Covalent

Atoms and Ions

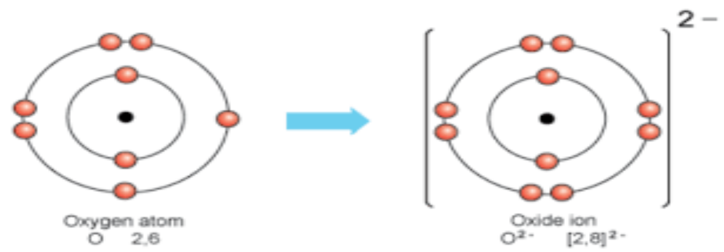
Atoms have no overall charge as they have the same number of positive protons and negative electrons.

Ions are atoms that have gained electrons (negative ion) or lost electrons (positive ion).

Metals lose electrons to become positive ions and have a full outer shell.



Non-metals gain electrons to become negative ions and have a full outer shell.



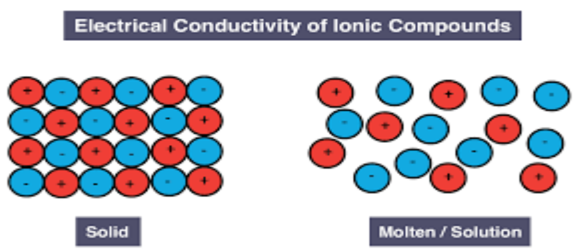
Properties of Ionic Compounds

Ionic compounds have high melting points because….

- They are giant lattices.
- Have strong electrostatic forces of attraction between oppositely charged ions.
- They need lots of energy to break forces.

Ionic compounds do not conduct electricity when solid as the ions are fixed and cannot move.

Ionic compounds do conduct electricity when molten or dissolved in water as the ions can now move and carry charge through whole structure.

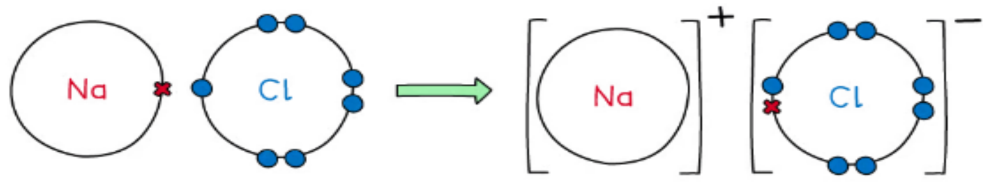


Ionic Bonding

In an ionic bond, the metal atom(s) lose electrons to become a positive ion. The non-metal atom(s) gain electrons to become a negative ion.

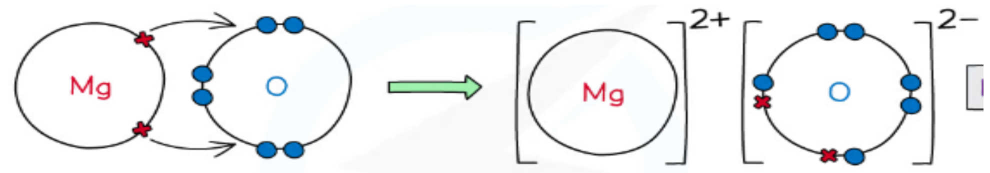
An ionic bond is a strong electrostatic attraction between the oppositely charged ions.

Example – Sodium Chloride (NaCl)



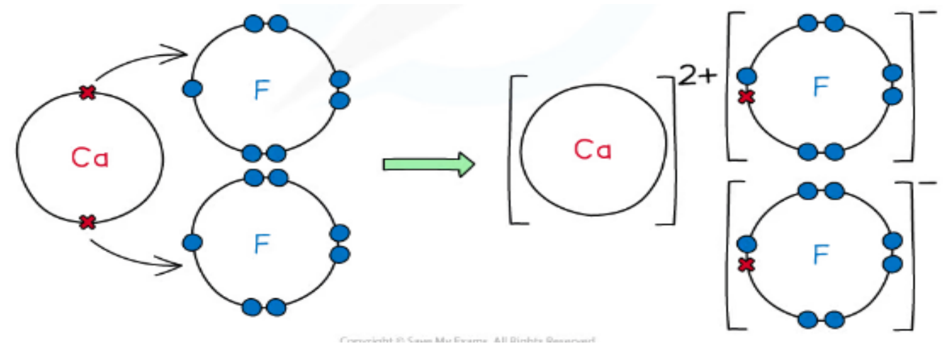
- One sodium atom loses 1 electron to form a Na^+ ion.
- One chlorine atom gains 1 electron to form a Cl^- ion.

Example – Magnesium Oxide (MgO)



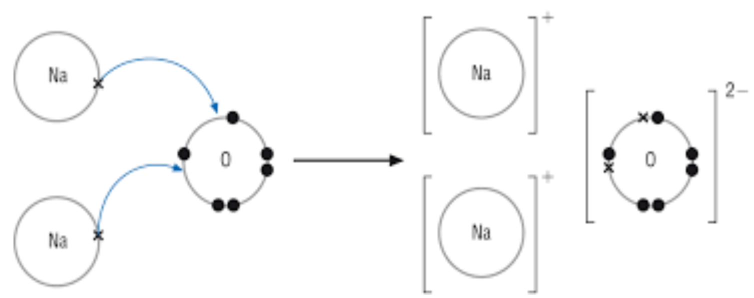
- One magnesium atom loses 2 electrons to become a Mg^{2+} ion.
- One oxygen atom gains 2 electrons to become a O^{2-} ion.

Example – Calcium Fluoride (CaF_2)



- One calcium atom loses 2 electrons to become a Ca^{2+} ion.
- Two fluoride atoms gain 1 electron each to form F^- ions.

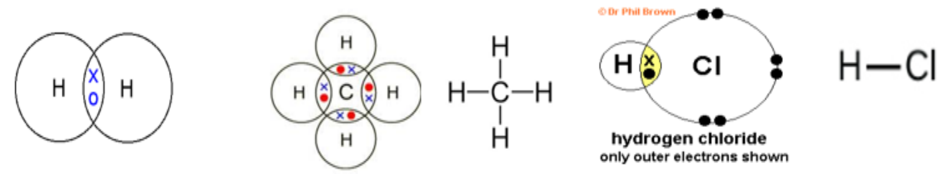
Example – Sodium Oxide (Na_2O)



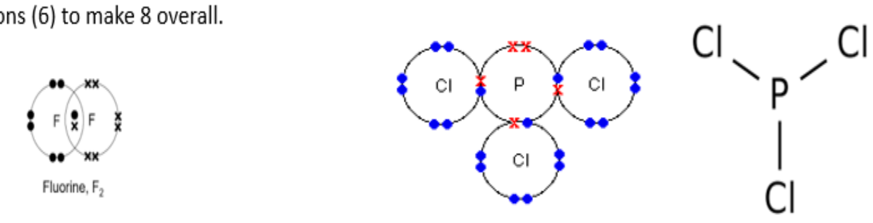
- Two sodium atoms lose 1 electron each to become a Na^+ ion.
- One oxygen atom gains 2 electrons to become a O^{2-} ion.

Covalent Bonding

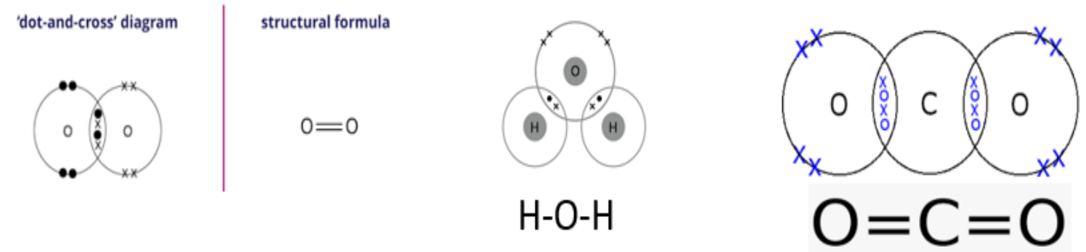
- This is when non-metal atoms share pair(s) of electrons
- Hydrogen will always share one pair of electrons in an overlap. It will also only have 2 electrons on its shell



- Group 7 elements will always share one pair of electrons in an overlap. However, remember to draw the rest of the electrons (6) to make 8 overall.



- Oxygen (and sulfur) will always share 2 pairs of electrons overall (could be with different atoms)

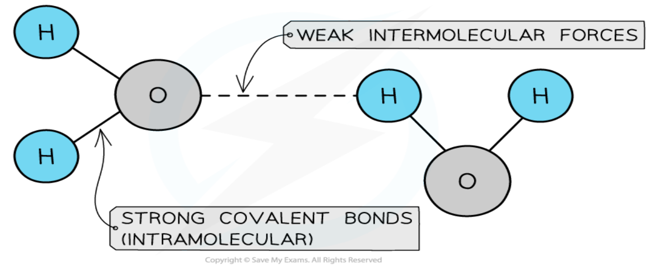


- Nitrogen will always share 3 pairs of electrons overall (could be with different atoms)

Properties of Simple Covalent Substances

Simple covalent substances have low melting points because…..

- They are small molecules.
- With weak intermolecular forces.
- They need little energy to break forces.



Simple covalent substances do not conduct electricity as they do not have any charges that can move and carry charge through the whole structure.

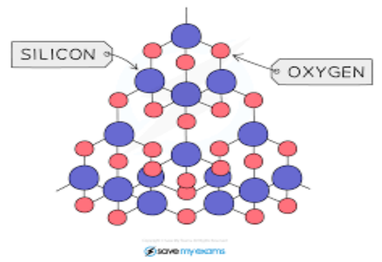
C2 – Giant Covalent, Metallic and Polymers

Giant Covalent Substances

These are macromolecules made of a giant lattice composed of lots of strong covalent bonds (non-metal atoms that share pairs of electrons)

Examples include allotropes (same element but different 3D structures) of carbon, such as diamond, graphite, graphene, and fullerenes.

Silicon dioxide, also known as silica, (SiO_2) is also a giant covalent substance, which has a similar structure to diamond, and therefore similar properties.



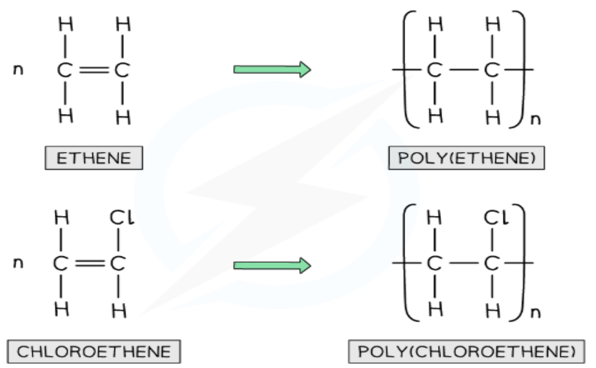
Polymers

Polymers are long chains of monomers (small units) joined together in a polymerisation reaction (kind of like a chain of paper clips joined together).

The backbone of the chain is joined together by strong covalent bonds.

There are intermolecular forces between the chains. AS the chains are long, the intermolecular forces are strong, and therefore require lots of energy to break. So, polymers are solids with generally high melting points, but lower than ionic compounds and giant covalent substances.

Writing and Naming Polymers

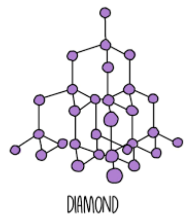
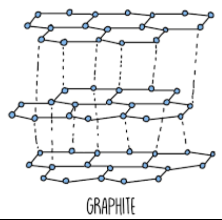


Diamond and Graphite

Both diamond and graphite are made of carbon elements.

They have different structures due to the conditions in which there were formed.

Due to their different structures, they have some different properties.

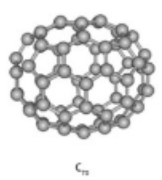
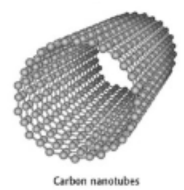
	Diamond	Graphite
Picture		
Structure	Each C atom is connected 4 other carbon atoms in a large tetrahedral structure	Each C atom is connected to 3 other carbon atoms in layers of hexagonal rings. This leaves an electron per carbon atom that is delocalised.
Melting Point	Very high, as it is a giant molecule with lots of strong covalent bonds, that require a lot of energy to break.	Very high, as it is a giant molecule with lots of strong covalent bonds, that require a lot of energy to break.
Hardness	Tetrahedral structure with no layers, so atoms cannot slide over each other.	Layers of hexagonal rings. Weak intermolecular forces between layers, so layers can slide over each other
Electrical Conductivity	No delocalised electrons to move and carry charge through whole structure	Delocalised electrons move in the between the layers and carry charge through whole structure.

Fullerenes

- Can be in the shape of a hollow sphere or a hollow (nano)tube.
- First fullerene discovered was C_{60} called buckminsterfullerene.
- The spherical structure is usually made of hexagonal rings but can be pentagons/heptagons.

Uses of Spherical Fullerenes

- **Catalysts** – large surface area (hollow), and therefore less catalyst needed for same effect.
- **Deliver Drugs in Body** – drug trapped inside the hollow structure, so it has fewer side effects on the other cells of the body.
- **Lubricant** – can roll due to its shape to reduce friction.

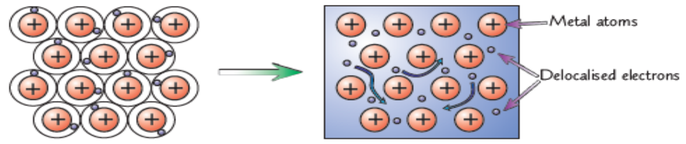
Uses of Nanotubes

- Conducts heat and electricity.
- They have high length to diameter ratio (very long compared to thickness

Tennis Rackets/Golf Clubs – have a very high tensile strength (doesn't break easily when stretched) without adding much weight.

Electronics – has delocalised electrons.

Metals



Metals are giant lattices with layers of positive metal atoms surrounded by a sea of delocalised electrons.

The delocalised electrons are the electrons on the outer shell.

High melting point

- Giant lattices
- With strong electronstatic forces of attraction between positive metal ions and delocalised electrons. This is a metallic bond.
- Lots of energy requried to break forces.

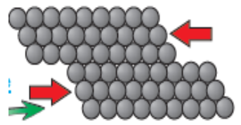
Conducts Electricity

- Delocalised electrons
- That move and carry charge through whole structure

Conducts Heat

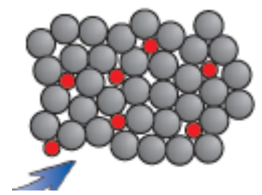
- Delocalised electrons
- That move and transfer thermal energy

Malleable



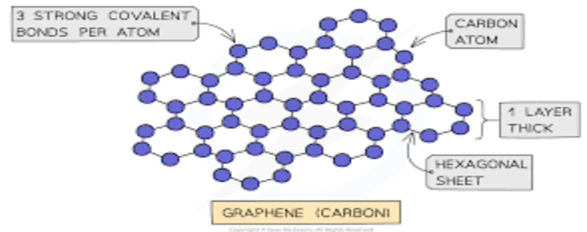
- Atoms are the same size, in layers.
- Atoms can slide over each other.

Alloys



- Alloys are mixtures of 2 or more metals.
- As a result, there are atoms of different sizes.
- This distorts the layered structure.
- Atoms cannot sldie over each other.

Graphene



Graphene is a single layer of graphite.

Used in electronics – has delocalised electrons to move and carry charge through whole structure. It is also transparent and very lightweight.

Used in composites - (made of 2 different materials) as it is lightweight and can add a lot of strength without adding to the weight of the material.

C3 – Quantitative Chemistry

Conservation of Mass

This is the idea that the total mass of reactants before a reaction is equal to the total mass of products after the reaction. The atoms simply rearrange themselves to form new chemicals.

Example 1

2Cu	+	O_2	→	CuO
15g	+	x	→	23g

In the case above, the total mass of the products after the reaction is 23g. So, the mass of oxygen is 23-15= 8g.

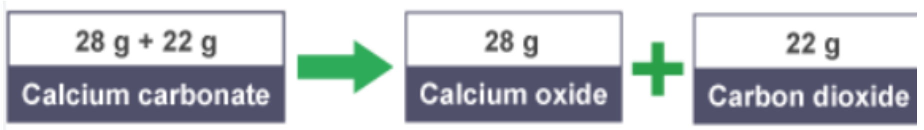
Example 2

$TiCl_4$	+	2Mg	→	Ti	+	$2MgCl_2$
40kg	+	20kg		x	+	33kg

In the case above, the total mass of the reactants before the reaction is 60kg. So, the mass of Ti is 60-33 = 27kg

Apparent Changes in Mass

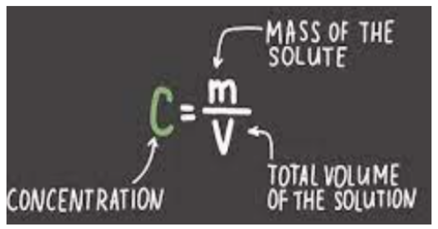
In oxidation reactions as shown above, the oxygen is a gaseous reactant that comes in from the air. The mass of oxygen is not measured, and so the mass of reaction appears to increase.

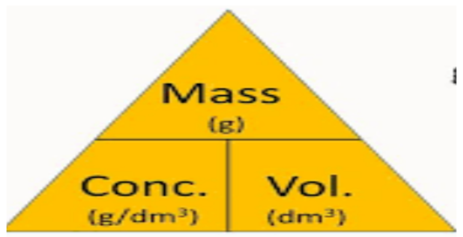


In thermal decomposition reactions as show above, carbon dioxide is a gaseous product which escapes into the air. The mass of carbon dioxide is not measured, and so the mass of the reaction appears to increase.

Concentration

Concentration is the amount of mass dissolved in a certain volume of liquid (usually water). It can be calculated using the equation and triangle below.





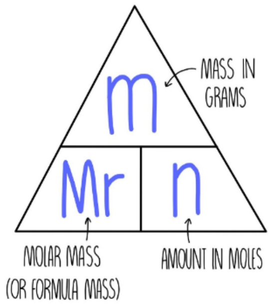
One thing to note here is that the volume has to be in dm^3 but is usually given in cm^3. To convert cm^3 into dm^3, divide by 1000.

Avogadro's Constant and Moles (HT only)

A mole of any substance contains the same number of particles (Avogadro's Constant) and is linked by the equation below.

Number of particles = moles x Avogadro's constant

Avogadro's constant is 6.022×10^{23}



This means that 1 mole of substances will have the same number of particles but different masses in grams. The mass of one mole of any substance is the M_r in grams.

Examples

1 mole of H2O is 18g.
1 mole of Cu is 63.5g.
1 mole of Mg is 24g.

Limiting Reactants (HT only)

In any reaction, there is usually one reactant that is limiting (less moles) and one in excess (more moles). In your exam, you maybe asked to work out the limiting reactant (as well as use the the limiting reactant to work out reacting masses.

Use the example below to work out the limiting reactant.

115 g of sodium was burnt in 80 g of oxygen. Which reactant is the limiting reactant?

$$4Na + O_2 \longrightarrow 2Na_2O$$

	4Na	O_2
Work out the M_r	23	32
Work out moles for each reactant	**115/23** **5 moles**	**80/32** **2.5 moles**
Divide the moles by the number in front. Smallest number is limiting reactant	**5/4 = 1.25**	**2.5 = 2.5**

Sodium is the limiting reactant

10 moles of sodium would have been needed for 2.5 moles of oxygen

Reacting Masses Examples (HT only)

- Aluminium is made by the electrolysis of aluminium oxide as shown in the equation below. What mass of aluminium can be formed from 1kg of aluminium oxide? Give your answer to 3 significant figures.

$$2Al_2O_3 \rightarrow 4Al + 3O_2$$

Step 1 – work out the M_r for the 2 substances in the question

- M_r of Al_2O_3 is 102
- M_r of Al is 27

Step 2 – calculate the moles of the substance for which the mass is given

- In this case, this is Al_2O_3
- Moles = mass/M_r
- Moles of Al_2O_3 = 1000g/102 = 9.80moles

Step 3 – use the symbol equation, and ratios to work out the moles of the substance for which you need to find the mass

- Looking at the equation $2Al_2O_3$: 4Al
- So, to get the moles of Al, you double the moles of Al_2O_3
- Moles of Al = 9.80 x 2 = 19.6moles

Step 4 – work out the mass of the substance using mass = moles x Mr

- Mass of Al = 19.6 x 27 = 529g

Magnesium reacts with iron chloride solution.

As soon as you see a question like this, think table!!

$$3\,Mg + 2\,FeCl_3 \rightarrow 2\,Fe + 3\,MgCl_2$$

(c) 0.120 g of magnesium reacts with excess iron chloride solution.

Relative atomic masses (A_r): Mg = 24 Fe = 56

Calculate the mass of iron produced, in mg

3:2 ratio

	3Mg	2Fe
Mass	0.120g	0.186g, so 186mg
Mr	24	56
Moles	0.005	0.00333

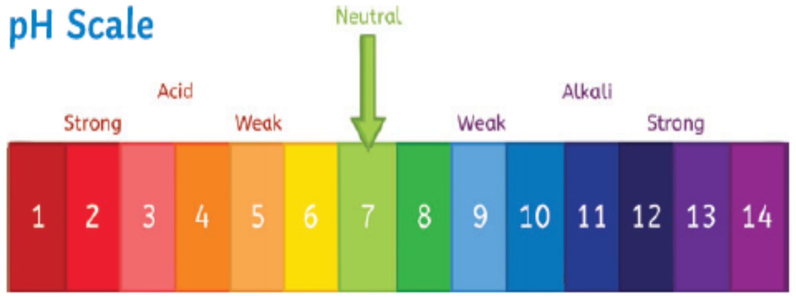
Balancing Equations Using Moles (HT only)

In a reaction, 9.2 g of sodium (Na) reacted with 3.2 g oxygen gas (O_2) to give 12.4 g sodium oxide (Na_2O).

	Na	O_2	Na_2O
1. Write down the symbol equation	Na	+ O_2 →	Na_2O
2. Write down the masses	9.2 g	3.2 g	12.4 g
3. Write down the M_r	23	32	62
4. Calculate the number of moles. Moles = mass / M_r	0.4	0.1	0.2
5. Convert to the simplest ratio	4	1	2
6. Write down the balanced equation	4Na	+ O_2 →	$2Na_2O$

C4 – Acids, Bases and Neutralisation

pH Scale



In aqueous solutions, acids produce H^+ ions and alkalis produce OH^- ions.

Neutral solutions are pH7 and are neither acids or alkalis.

For example, in neutralisation reactions, hydrogen ions from an acid react with hydroxide ions from an alkali to produce water:

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

Strong and Weak Acids (Higher Tier Only)

A **strong** acid **completely dissociates** in a solution. For example: $HCl \longrightarrow H^+ + Cl^-$

Hydrochloric acid is able to completely dissociate in solution to form hydrogen and chloride ions.

Examples of strong acids include nitric acid (HNO_3) and sulfuric acid (H_2SO_4).

Weak acids in comparison only partially dissociate.

For example acetic acid **partially dissociates** to form a hydrogen and acetate ion.

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

The **double arrow** symbol indicates that the reaction is **reversible**. Both the forward and reverse reaction occur at the same time and the reaction never goes to completion.

Reactions of Acids

The general formula for the reaction between an acid and a metal is: acid + metal ⟶ salt + hydrogen

For example: hydrochloric acid + sodium ⟶ sodium chloride + hydrogen

$2HCl + 2Na \longrightarrow 2NaCl + H_2$

When an acid reacts with an alkali, a neutralisation reaction takes place and a salt and water are produced.

The general formula for this kind of reaction is as follows:

acid + alkali ⟶ salt + water

hydrochloric acid + sodium hydroxide ⟶ sodium chloride + water

$HCl + NaOH \longrightarrow NaCl + H_2O$

Naming Salts

The first part comes from the metal in the metal carbonate, oxide or hydroxide. The second part of the name comes from the acid that was used to make it.

For example, sodium chloride.

Acid Used	Salt Produced
hydrochloric	chloride
nitric	nitrate
sulfuric	sulfate

Reactions with Carbonates

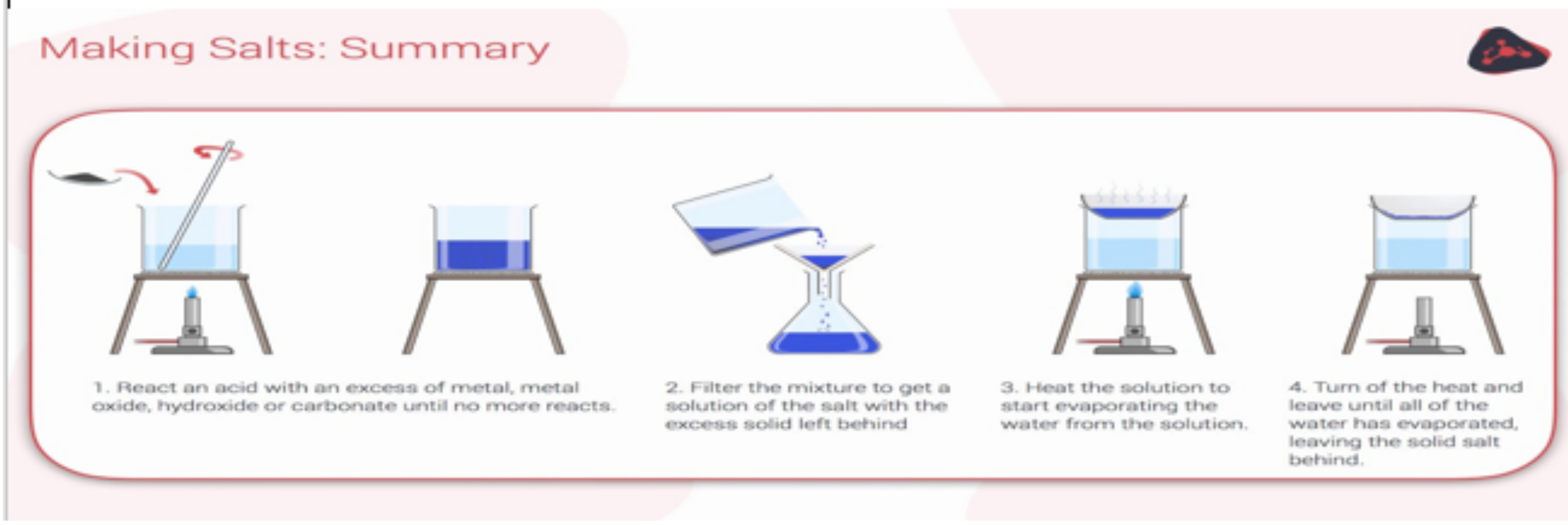
The general formula for the reaction between an acid and a carbonate is: acid + carbonate ⟶ salt + water + carbon dioxide

hydrochloric acid + calcium carbonate ⟶ calcium chloride + water + carbon dioxide

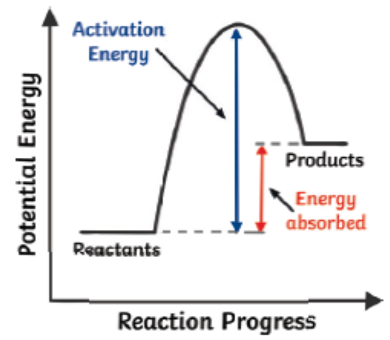
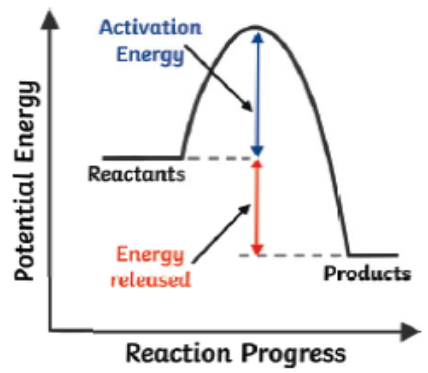
Making a Soluble Salt (Required Practical)

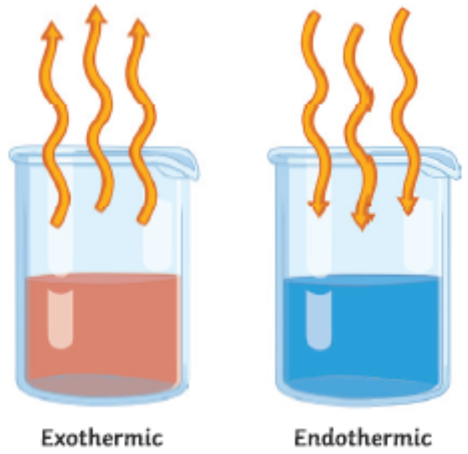
Method for making copper sulfate crystals

1. Add solid copper oxide to sulfuric acid until there is an excess of copper oxide that does not dissolve. This is to make sure all acid has reacted
2. Mixture is heated to speed up the rate of reaction and make as much product as possible and make sure all acid has reacted
3. Filter the mixture to remove excess copper oxide using filter paper and a funnel. Collect the filtrate (copper sulfate solution) in an evaporating basin
4. Once some of the solvent has evaporated, or when you see crystals start to form (the point of crystallisation), remove the dish from the heat and leave the solution to cool and dry in a warm place
5. The salt should start to form crystals as it becomes insoluble in the cold, highly concentrated solution.



C5 – Energy Changes

	Endothermic	Exothermic
Temperature increase or Temperature decrease?	Temperature decrease	Temperature increase
Heat energy released to or absorbed from surroundings?	Heat energy absorbed from surroundings	Heat energy released to surroundings
What has more energy? Reactants or products?	Products	Reactants
Energy Profile Diagram		
Examples	Sports injury packs Thermal decomposition Photosynthesis	Neutralisation reactions Combustion reactions Respiration

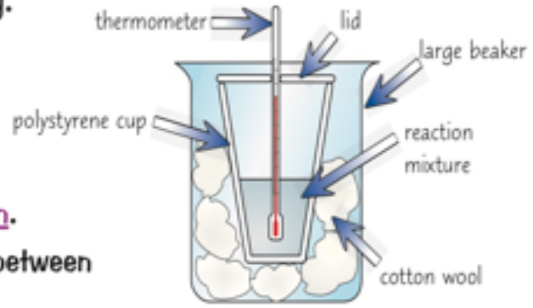


Activation Energy – the minimum amount of energy required for a chemical reaction to take place.

Catalysts – increase the rate of a reaction. Catalysts provide an alternative pathway for a chemical reaction to take place by **lowering** the activation energy.

Energy Changes – Required Practical

1) You can measure the amount of energy released by a chemical reaction (in solution) by taking the temperature of the reagents, mixing them in a polystyrene cup and measuring the temperature of the solution at the end of the reaction. Easy.
2) The biggest problem with energy measurements is the amount of energy lost to the surroundings.
3) You can reduce it a bit by putting the polystyrene cup into a beaker of cotton wool to give more insulation, and putting a lid on the cup to reduce energy lost by evaporation.
4) This method works for neutralisation reactions or reactions between metals and acids, or carbonates and acids.



5) You can also use this method to investigate what effect different variables have on the amount of energy transferred — e.g. the mass or concentration of the reactants used.
6) Here's how you could test the effect of acid concentration on the energy released in a neutralisation reaction between hydrochloric acid (HCl) and sodium hydroxide (NaOH):

1) Put 25 cm³ of 0.25 mol/dm³ of hydrochloric acid and sodium hydroxide in separate beakers.
2) Place the beakers in a water bath set to 25 °C until they are both at the same temperature (25 °C).
3) Add the HCl followed by the NaOH to a polystyrene cup with a lid — as in the diagram above.
4) Take the temperature of the mixture every 30 seconds, and record the highest temperature.
5) Repeat steps 1-4 using 0.5 mol/dm³ and then 1 mol/dm³ of hydrochloric acid.

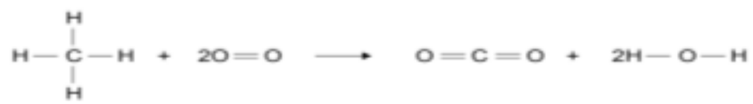
Calculating Energy Changes – HT Only

0 6 . 2 Combustion is an exothermic reaction.

Calculate the overall energy change for the complete combustion of one mole of methane in oxygen.

[3 marks]

$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$



Bond	Bond energy in kJ/mol
C—H	413
O=O	498
C=O	805
O—H	464

Bonds broken
4 C-H 4 x 413 = 1652
2O=O 2 x 498 = 996
Total Bonds Broken = 2648

Bonds Formed
2 C=O 2 x 805 = 1610
4 O-H 4 x 464 = 1856
Total Bonds Formed = 3466

Energy change = 2468 – 3466 = -818kJ/mol
This is an exothermic reaction as more energy released when forming bonds than needed to break bonds

C6 – Rate of Reaction

Rate of Reaction

- Rate of Reaction is how fast a reaction is occurring.
- Can be calculated by measuring the formation of products over time or how quickly reactants are used up.
- Equations for rate of reaction are….

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

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

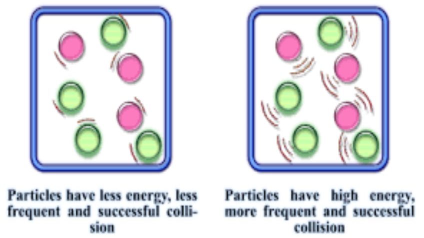
- Units for rate of reaction are either g/s, cm^3/s or mol/s (HT only)
- The following factors affect the rate of a reaction…..
  - Temperature
  - Concentration (for solutions)
  - Pressure (for gases)
  - Surface area
  - Catalyst

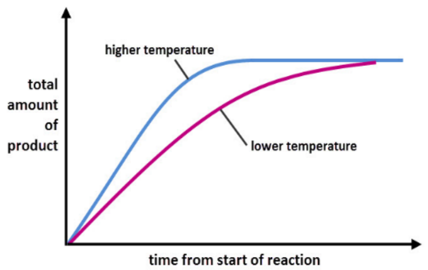
Collision Theory

- In order for a reaction to occur, particles must collide and collide with enough energy called activation energy.
- Activation energy is the minimum energy needed for a reaction to occur.
- Increasing temperature, increasing concentration/pressure and increasing the surface area all increase the frequency of collisions.
- Increasing the temperature also gives more particles activation energy, so more successful collisions.

Temperature

- Increasing the temperature, increases the kinetic energy of the particles.
- Particle move faster.
- This increases the frequency of collisions.
- This increases the rate of reaction.
- However, same amount of product is made, just faster.

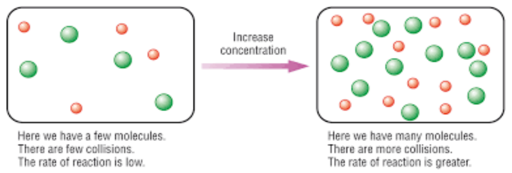


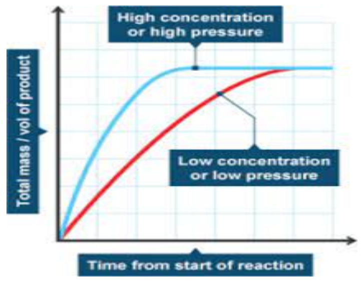


- Increasing temperature also gives more activation energy, so more collisions are successful.

Concentration/Pressure

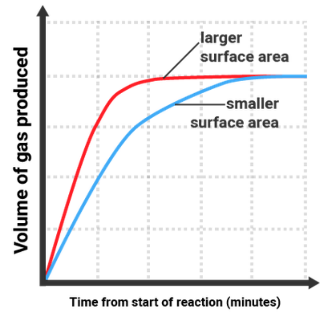
- Increasing the concentration, increases the number of particles in a given volume.
- This increases the frequency of collisions.
- This increases the rate of reaction.
- Changing the concentration, also affects the amount of product.





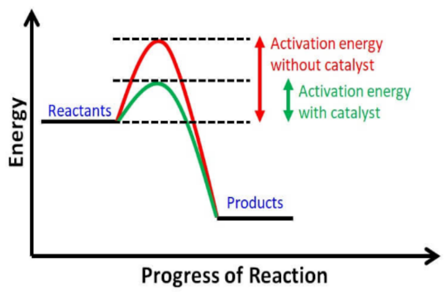
Surface Area

- Breaking a large solid into smaller pieces, increases the surface area, so more particles exposed.
- This increases the frequency of collisions.
- This increases the rate of reaction.
- However, same amount of product made, just faster.



Catalyst

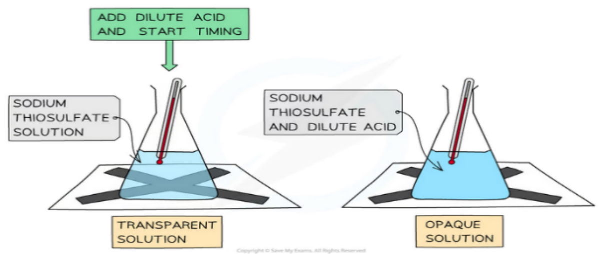
- Provides an alternative route with a lower activation energy.
- This increases the rate of reaction.
- However, same amount of product made, just faster.



Measuring Rate of Reaction When a Solid is a Product

$Na_2S_2O_{3(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + S_{(s)} + SO_{2(g)} + H_2O_{(l)}$

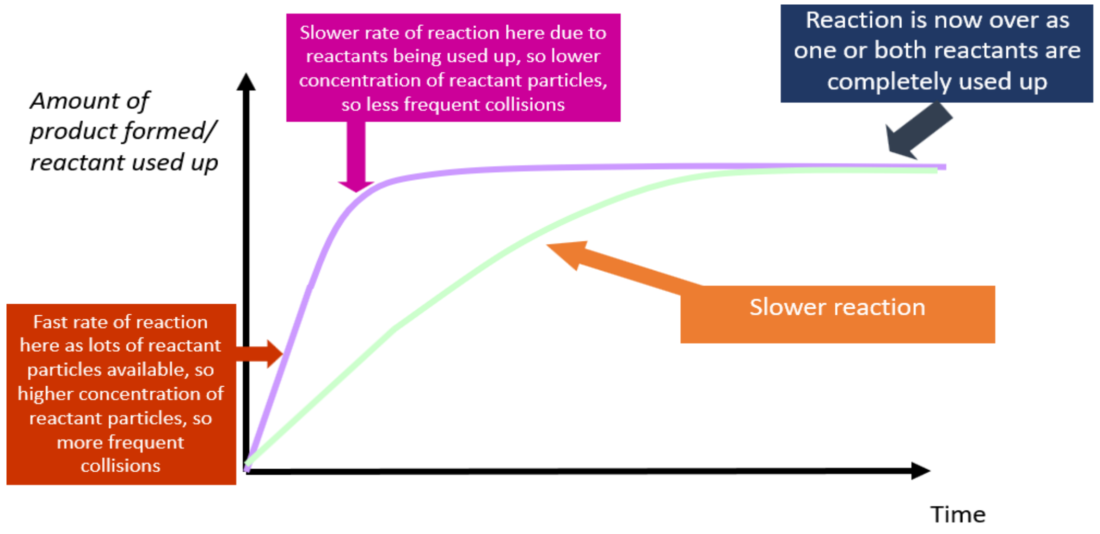
For the reaction above, sulfur is a solid precipitate, which causes a change in colour.



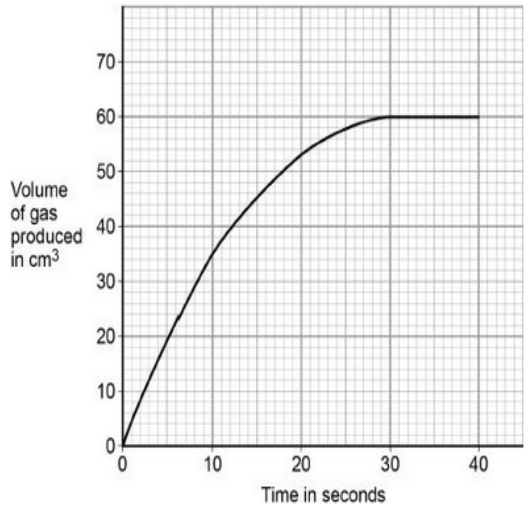
Method

1. Measure $50cm^3$ of sodium thiosulphate using a measuring cylinder and put into a conical flask placed on a black cross.
2. Measure $10cm^3$ of hydrochloric acid using a different measuring cylinder (prevent reaction from starting)
3. Pour the hydrochloric acid into the conical flask and start timer.
4. Stop the timer when the cross is no longer visible.
5. Repeat by changing the independent variable.
6. Mention 2 control variables (could be volumes of the solutions, concentration of solutions or temperature)

Rate of Reaction Graph



Calculating Rate of Reaction from a Graph



Calculate mean rate of reaction between 10s and 30s.

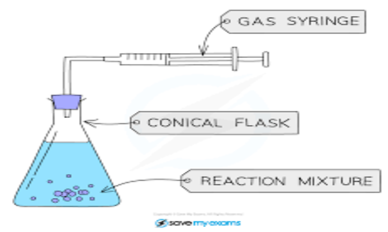
- Using lines on the graph, work out gas given off at 10s and 20s.
- 10s is $35cm^3$ and 20s is $53cm^3$.
- Rate of reaction is $18cm^3/20cm^3$, which is $0.9cm^3/s$.

Calculate rate of reaction at 20s.

- Draw a tangent to the curve at 20s.
- Then, calculate gradient of the tangent.

Measuring Rate of Reaction When a Gas is a Product

$Mg_{(s)} + HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$



Hydrogen is a gaseous product, so can be collected in a gas syringe.

For other gases, such as carbon dioxide, can also measure the rate of mass loss using a mass balance.

Method

1. Measure $25cm^3$ of hydrochloric acid using a measuring cylinder and put into a conical flask.
2. Measure 3cm length piece of magnesium ribbon
3. Place the magnesium into the flask, and immediately place bung on top, and start timer.
4. Measure the time taken for 20cm3 of gas to be given off or measure the volume of gas given off in 30s.
5. Repeat by changing the independent variable.
6. Mention 2 control variables (could be volumes of the solutions, concentration of solutions, mass of solid or temperature)