Edexcel International AS Chemistry

Your notes

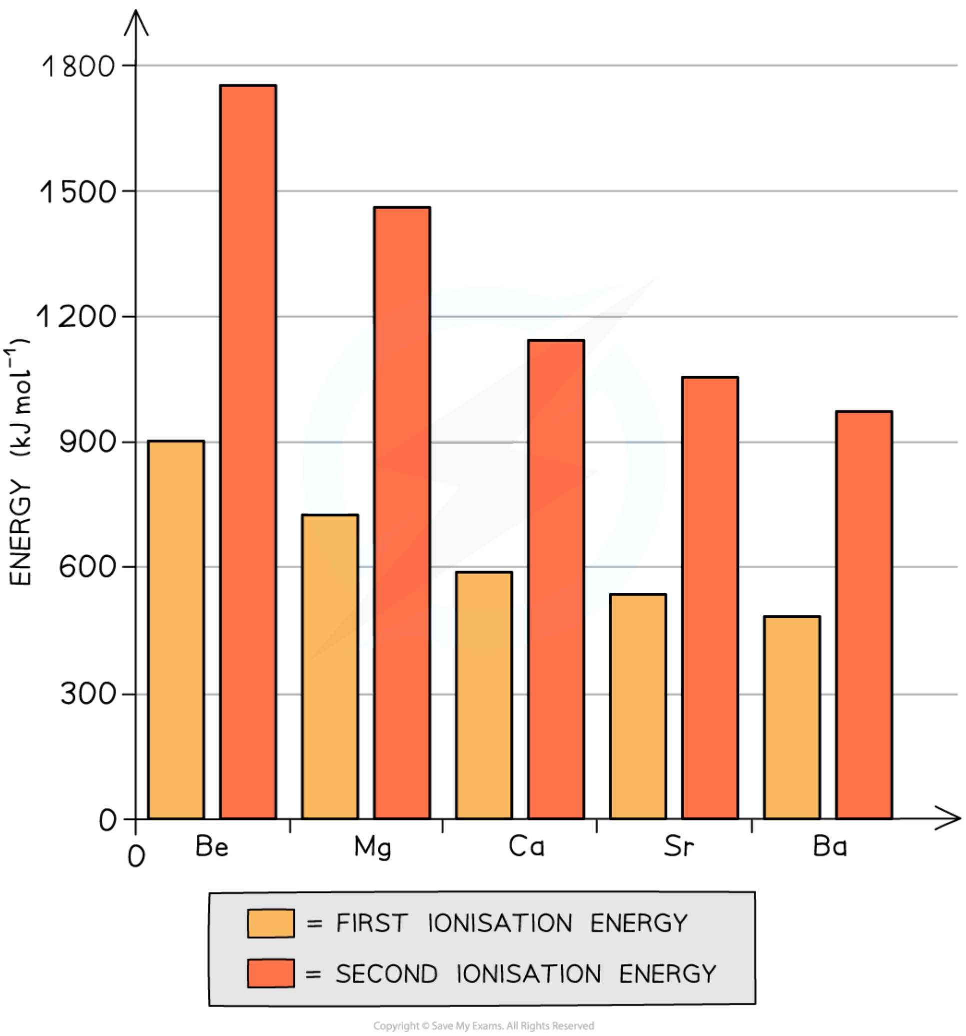
Groups 1 & 2

Contents

Your notes

Ionisation Energy Trends

- All elements in Groups 1 (also called **alkali metals)** have one electron in their **outermost principal quantum shell**
- All elements in Groups 2 (also called **alkali earth metals)** have two electrons in their **outermost principal quantum shell**
- All Group 1 and Group 2 metals can form **ionic compounds** in which they donate these **outermost electrons** (so they act as **reducing agents)** to become an ion with either a +1 or +2 charge (so they themselves become **oxidised)**
- Going down the group, the metals become more **reactive**
  - This can be explained by looking at the Group 2 ionisation energies:



The graph shows that both the first and second ionization energies decrease going down the group

Your notes

- The same trend is observed with Group 1 alkali metals
- The **first ionisation energy** is the energy needed to remove the first outer electron of an atom
- The **second ionisation energy** is the energy needed to remove the second outer electron of an atom
- The graph above shows that going down the group, it becomes easier to remove the outer two electrons of the metals
- Though the **nuclear charge** increases going down the group (because there are more protons), factors such as an **increased shielding effect** and a **larger distance** between the outermost electrons and nucleus outweigh the attraction of the higher nuclear charge

Reactivity Trend

Group 1 metals

- The reactivity of the group 1 metals **increases** as you go down the group
- When a group 1 element reacts its atoms only need to lose electron, as there is only 1 electron in the outer shell
  - When this happens, 1+ ions are formed
- The next shell down automatically becomes the outermost shell and since it is **already full,** a group 1 ion obtains **noble gas configuration**
- As you go down group 1, the number of shells of electrons increases by 1
  - This means that the outermost electron gets **further** away from the nucleus, so there are **weaker forces of attraction** between the outermost electron and the nucleus
  - **Less energy** is required to overcome the force of attraction as it gets weaker, so the outer electron is lost **more easily**
  - So, the alkali metals get more reactive as you descend the group

Group 2 metals

- The reactivity of the Group 2 metals also increases down the group for the same reasons
  - The outermost electron gets **further** away from the nucleus, so there are **weaker forces of attraction** between the outermost electron and the nucleus
  - **Less energy** is required to overcome the force of attraction as it gets weaker, so the outer electron is lost **more easily**
- This can be observed when the Group 2 metals react with water as well
  - Magnesium reacts **extremely slowly** with cold water
  - Calcium reacts **fairly vigorously** with cold water in an exothermic reaction

Reactions with Oxygen, Chlorine & Water

Group 1 metals

Reaction with oxygen

- The alkali metals react with oxygen in the air forming **metal oxides,** which is why the alkali metals **tarnish** when exposed to the air
- The metal oxide produced is a dull coating which covers the surface of the metal
- The metals tarnish more rapidly as you go down the group

Summary of the Reactions of the First Three Alkali Metals with Oxygen

Element	Reaction
Li	Lithium + Oxygen ⟶ Lithium oxide $4Li(s) + O_2(g) \longrightarrow 2Li_2O(s)$
Na	Sodium + Oxygen ⟶ Sodium oxide $4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$
K	Potassium + Oxygen ⟶ Potassium superoxide $K(s) + O_2(g) \longrightarrow KO_2(s)$

Reaction with chlorine

- The reactions of the Group 1 elements with chlorine are similar in appearance to the reactions of the Group 1 metals with oxygen
- Sodium, for example, burns with an intense orange flame in chlorine in exactly the same way that it does in pure oxygen
- The other metals also behave the same way in both gases
  - In each case, a white solid is formed, the simple chloride

$$2M\ (s) + Cl_2\ (g) \rightarrow 2MCl\ (s)$$

Reactions with water

- The reactions of the alkali metals with water get more vigorous as you descend the group

Summary of the Reactions of the First Three Alkali Metals with Water

Your notes

Element	Reaction	Observations
Li	Lithium + Water ⟶ Lithium hydroxide + Hydrogen $2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2\ (g)$	∘ Relatively slow reaction ∘ Lithium doesn't melt ∘ Fizzing can be seen and heard as the lithium reacts
Na	Sodium + Water ⟶ Sodium hydroxide + Hydrogen $2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2\ (g)$	∘ Large amounts of heat released causes the sodium to melt ∘ Hydrogen released catches fire and causes the ball of sodium to dash across the surface
K	Potassium + Water ⟶ Potassium hydroxide + Hydrogen $2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2\ (g)$	∘ Reacts more violently than sodium ∘ Enough heat released so hydrogen burns with a lilac coloured flame ∘ Melts into a shiny ball that dashes around the surface

Group 2 metals

Reactions with water and oxygen

- The reaction of Group 2 metals with oxygen follows the following general equation:

$$2M\ (s) + O_2\ (g) \rightarrow 2MO\ (s)$$

- Where M is any metal in Group 2
- Remember than Sr and Ba **also** form a peroxide, MO_2
- The reaction of all metals with water follows the following general equation:

$$M\ (s) + 2H_2O\ (l) \rightarrow M(OH)_2\ (s) + H_2\ (g)$$

- Except for, Be which does not react with water

Group 2 Metals Reacting with Water and Oxygen – Equations

	Reaction with oxygen	Reaction with water
Mg	$2Mg\ (s) + O_2\ (g) \rightarrow 2MgO\ (s)$	$Mg\ (s) + H_2O\ (g) \rightarrow MgO\ (s) + H_2\ (g)$
Ca	$2\ Ca\ (s) + O_2\ (g) \rightarrow 2CaO\ (s)$	$Ca\ (s) + 2H_2O\ (l) \rightarrow Ca(OH)_2\ (aq) + H_2\ (g)$
Sr	$2Sr\ (s) + O_2\ (g) \rightarrow 2SrO\ (s)$ $Sr\ (s) + O_2\ (g) \rightarrow SrO_2\ (s)$	$Sr\ (s) + 2H_2O\ (l) \rightarrow Sr(OH)_2\ (aq) + H_2\ (g)$

Ba	$2Ba\ (s) + O_2\ (g) \rightarrow 2BaO\ (s)$ $Ba\ (s) + O_2\ (g) \rightarrow BaO_2\ (s)$	$Ba\ (s) + 2H_2O\ (l) \rightarrow Ba(OH)_2\ (aq) + H_2\ (g)$

Your notes

- Magnesium reacts **extremely slowly** with cold water:

$$Mg\ (s) + 2H_2O\ (l) \rightarrow Mg(OH)_2\ (aq) + H_2\ (g)$$

- The solution formed is **weakly alkaline** (pH 9–10) as **magnesium hydroxide** is only **slightly soluble**
- However, when magnesium is **heated in steam**, it reacts **vigorously** with steam to make magnesium oxide and hydrogen gas:

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

Reactions with chlorine

- Group 2 metals react with chlorine gas to give the metal chloride
  - For example

$$Mg\ (s) + Cl_2\ (g) \rightarrow MgCl_2\ (s)$$

Your notes

Reactions with Acid

Overall equations

- We have seen that Group 1 and Group 2 metals can react with oxygen to form oxides

metal + oxygen → metal oxide

- Oxides can react with water to form hydroxides

oxide + water → hydroxide

- If calcium oxide is added to water, you should know that calcium hydroxide is formed
  - Calcium hydroxide solution is also called **limewater**
- When metal oxides react with dilute hydrochloric acid and dilute sulfuric acid, chloride and sulfate salts are formed

metal oxide + dilute hydrochloric acid → metal chloride + water

metal oxide + dilute sulfuric acid → metal sulfate + water

- Similar reactions occur with hydroxides

metal hydroxide + dilute hydrochloric acid → metal chloride + water

metal hydroxide + dilute sulfuric acid → metal sulfate + water

Reactions of Group 1 oxides with water

- The Group 1 metal oxides will react with water to give a **colourless alkaline solution**
  - For example

$$Na_2O\ (s) + H_2O\ (l) \rightarrow 2NaOH\ (aq)$$

- The solution is alkaline as hydroxide ions are released which also occurs for the reaction of Group 2 oxides with water

$$O^{2-}\ (s) + H_2O\ (l) \rightarrow 2OH^-\ (aq)$$

Reactions of Group 1 hydroxides with dilute acid

- The Group 1 metal hydroxide is acting as an alkali when it is added to dilute acid
- When an alkali reacts with an acid, a neutralisation reaction occurs forming salt and water
  - For example

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

$$NaOH\ (aq) + H_2SO_4\ (aq) \rightarrow Na_2SO_4\ (aq) + H_2O\ (l)$$

Reactions of Group 2 oxides with water

- All Group 2 oxides are **basic**, except for BeO which is **amphoteric** (it can act both as an acid and base)
- Group 2 oxides react water to form **alkaline** solutions which get more **alkaline** going down the group

Your notes

Group 2 Oxides reacting with Water

Group 2 oxide	Reaction with water	Observations
MgO	$MgO\ (s) + H_2O\ (l) \rightarrow Mg(OH)_2\ (s)$	MgO is only slightly soluble in water, therefore a weakly alkaline solution (pH 10.0) is formed
CaO	$CaO\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (s)$	A vigorous reaction which releases a lot of energy, causing some of the water to boil off as the solid lump seems to expand and open (pH 11.0)
SrO	$SrO\ (s) + H_2O\ (l) \rightarrow Sr(OH)_2\ (aq)$	
BaO	$BaO\ (s) + H_2O\ (l) \rightarrow Ba(OH)_2\ (aq)$	

Group 2 Oxides reacting with acid

- Group 2 sulfates also form when a Group 2 oxide is reacted with sulfuric acid
- The **insoluble sulfates** form at the **surface** of the oxide, which means that the solid oxide beneath it can't react with the acid
- This can be prevented to an extent by using the oxide in **powder** form and **stirring**, in which case neutralisation can take place

Reactions of Group 2 hydroxides with dilute acid

- The Group 2 metal hydroxides form **colourless solutions of metal salts** when they react with a dilute acid
- The sulfates decrease in **solubility** going down the group (barium sulfate is an insoluble white precipitate)

Group 2 Hydroxide Reactions with Dilute Acids

Group 2 hydroxide	Reaction with dilute HCl	Reaction with dilute H_2SO_4
$Mg(OH)_2$	$Mg(OH)_2\ (s) + 2HCl\ (aq) \rightarrow MgCl_2\ (aq) + 2H_2O\ (l)$	$Mg(OH)_2\ (s) + H_2SO_4\ (aq) \rightarrow MgSO_4\ (aq) + 2H_2O\ (l)$

$Ca(OH)_2$	$Ca(OH)_2$ (s) + 2HCl (aq) → $CaCl_2$ (aq) + $2H_2O$ (l)	$Ca(OH)_2$ (s) + H_2SO_4 (aq) → $CaSO_4$ (aq) + $2H_2O$ (l)
$Sr(OH)_2$	$Sr(OH)_2$ (s) + 2HCl (aq) → $SrCl_2$ (aq) + $2H_2O$ (l)	$Sr(OH)_2$ (s) + H_2SO_4 (aq) → $SrSO_4$ (s) + $2H_2O$ (l)
$Ba(OH)_2$	$Ba(OH)_2$ (s) + 2HCl (aq) → $BaCl_2$ (aq) + $2H_2O$ (l)	$Ba(OH)_2$ (s) + H_2SO_4 (aq) → $BaSO_4$ (s) + $2H_2O$ (l)

Your notes

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Solubility Trends

Group 2 hydroxides

- Going down the group, the solutions formed from the reaction of Group 2 oxides with water become more **alkaline**
- When the oxides are dissolved in water, the following ionic reaction takes place:

$$O^{2-}(aq) + H_2O(l) \rightarrow 2OH^-(aq)$$

- The higher the **concentration** of OH^- ions formed, the more **alkaline** the solution
- The **alkalinity** of the formed solution can therefore be explained by the **solubility** of the Group 2 **hydroxides**

Solubility of the Group 2 Hydroxides Table

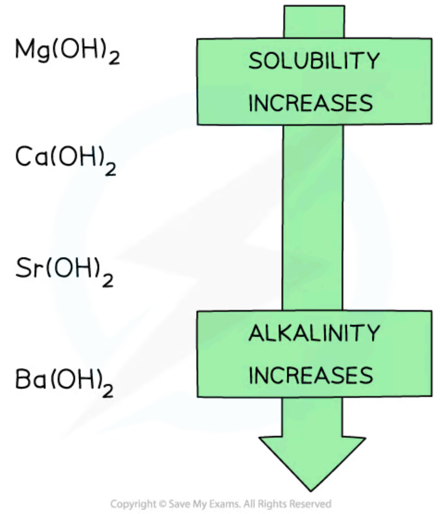
Group 2 hydroxide	Solubility at 298 K (mol / 100 g of water)
$Mg(OH)_2$	2.0×10^{-5}
$Ca(OH)_2$	1.5×10^{-3}
$Sr(OH)_2$	3.4×10^{-3}
$Ba(OH)_2$	1.5×10^{-2}

- The hydroxides dissolve in water as follows:

$$X(OH)_2(aq) \rightarrow X(aq) + 2OH^-(aq)$$

Where X is the Group 2 element

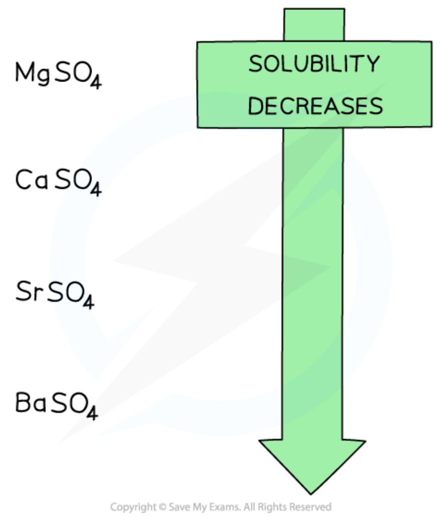
- When the metal oxides react with water, a Group 2 hydroxide is formed
- Going down the group, the **solubility** of these hydroxides **increases**
- This means that the **concentration** of OH^- ions **increases,** increasing the pH of the solution
- As a result, going down the group, the **alkalinity** of the solution formed increases when Group 2 oxides react with water



Going down the group, the solubility of the hydroxides increases which means that the solutions formed from the reactions of the Group 2 metal oxides and water become more alkaline going down the group

Group 2 sulfates

- The solubility of the Group 2 sulfates decreasing going down the group



Going down the group, the solubility of the sulfates decreases

Examiner Tips and Tricks

You may be wondering why there are no trends here for the solubility of Group 1 hydroxides and sulfates. You should recall from previous studies that Group 1 compounds are all soluble in water. They will therefore not produce any precipitates when testing for cations, so to identify them you need to use flame tests.

Group 1 hydroxides will be more soluble than Group 2 hydroxides. Even though we say the solubility of the Group 2 hydroxides increases down the group barium hydroxide is less soluble than a Group 1 hydroxide such as potassium hydroxide.

At 25 °C the solubility of **$Ba(OH)_2$** is 4.68 g / 100 cm^3

At 25 °C the solubility of **KOH** is 121 g / 100 cm^3

Your notes

Thermal Stability

- **Thermal decomposition** is the breakdown of a compound into two or more different substances using **heat**

Thermal decomposition of carbonates

- In Group 1, lithium carbonate when heated will decompose producing lithium oxide and carbon dioxide

$$Li_2CO_3\,(s) \xrightarrow{\text{HEAT}} Li_2O\,(s) + CO_2\,(g)$$

- The rest of the Group 1 carbonates don't decompose at Bunsen temperatures
    - The decomposition temperatures increase as you go down the Group
- The Group 2 carbonates break down (**decompose**) when they are heated to form the **metal oxide** and give off **carbon dioxide gas**
- The general equation for the decomposition of Group 2 carbonates is:

$$XCO_3\,(s) \xrightarrow{\text{HEAT}} XO\,(s) + CO_2\,(g)$$

X = Group 2 element

- Going **down** the group, more heat is needed to break down the carbonates

$$MgCO_3 \xrightarrow{\text{HEAT}} MgO + CO_2$$

Thermal decomposition of nitrates

- The only Group 1 nitrate that will decompose to produce nitrogen dioxide (which is a brown toxic gas) and oxygen is lithium nitrate $LiNO_3$

$$4LiNO_3\,(s) \xrightarrow{\text{HEAT}} 2Li_2O\,(s) + 4NO_2\,(g) + O_2\,(g)$$

- The rest of the Group don't decompose so completely producing the metal **nitrite** (NO_2^-) and oxygen, but no nitrogen dioxide

$$2XNO_3\,(s) \xrightarrow{\text{HEAT}} 2XNO_2\,(s) + O_2\,(g)$$

- All the nitrates from sodium to caesium decompose in this same way, the only difference being how hot they have to be to undergo the reaction.
- Down Group 1, the decomposition gets more difficult, and you have to use higher temperatures.
- The Group 2 nitrates break **decomposed** when they are heated to form the **metal oxide, oxygen gas** and **nitrogen dioxide gas**

- Since the formed nitrogen dioxide gas is toxic, the decomposition of nitrates is often carried out in a **fume cupboard**
- The general equation for the decomposition of Group 2 nitrates is:

$$Mg(NO_3)_2 \xrightarrow{\text{HEAT}} MgO + \frac{1}{2}O_2 + 2NO_2$$

X = Group 2 element

- Going **down** Groups 1 and 2, more heat is needed to break down the carbonate and nitrate ions
- The thermal stability of the Group 1 and 2 carbonates and nitrates therefore **increases** down the group
  - The smaller positive ions at the top of the groups will polarise the anions more than the larger ions at the bottom of the group
    - The small positive ion attracts the delocalised electrons in the nitrate or carbonate ion towards itself
    - The higher the charge and the smaller the ion the higher the polarising power
  - The more polarised they are, the more likely they are to thermally decompose as the bonds in the carbonate and nitrate ions become weaker

Testing for gases

- Carbon dioxide
  - When bubbled through limewater, carbon dioxide gas will turn the limewater milky
  - This is a fine white precipitate of calcium carbonate, $CaCO_3$

$$Ca(OH)_2\,(aq) + CO_2\,(g) \rightarrow CaCO_3\,(s) + H_2O\,(l)$$

- Oxygen
  - If oxygen gas is present, it will relight a glowing splint

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

- Nitrogen dioxide
  - NO_2 is a toxic brown - orange gas and if dissolved in water it would give an acidic solution

$$2NO_2\,(g) + H_2O\,(l) \rightarrow 2HNO_3\,(aq) + HNO_2\,(aq)$$

Your notes

Characteristic Flame Colours

- Metal ions produce a **colour** if heated strongly in a flame
- Ions from **different** metals produce **different colours**
- The flame test is thus used to identify metal ions by the colour of the flame they produce
- Dip the loop of an **unreactive** metal wire such as nichrome or platinum in concentrated acid, and then hold it in the blue flame of a Bunsen burner until there is no colour change
- This cleans the wire loop and avoids **contamination**
  - This is an important step as the test will only work if there is just **one type** of ion present
  - Two or more ions means the colours will mix, making identification erroneous
- Dip the loop into the solid sample and place it in the edge of the **blue** Bunsen flame
- Avoid letting the wire get so hot that it glows red otherwise this can be confused with a flame colour

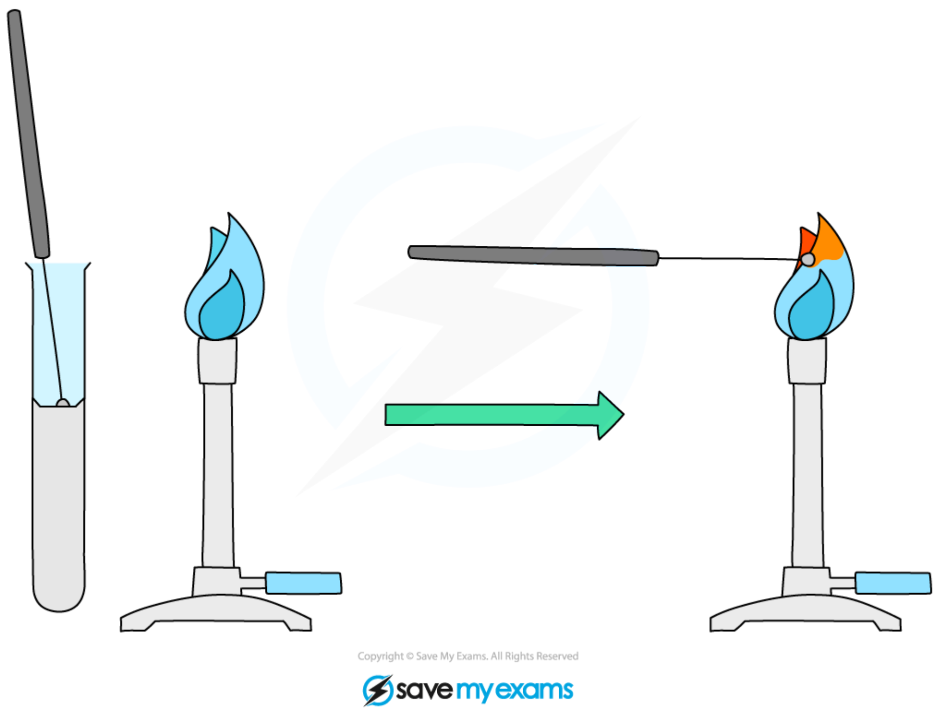


Diagram showing the technique for carrying out a flame test

Explanation for the occurrence of the flame

- In a flame test the heat causes the electron to move to a higher energy level

- The electron is unstable at this energy level so falls back down
- As it drops back down from the higher to a lower energy level, energy is emitted in the form of visible light energy with the wavelength of the observed light

Colours Observed in Flame Tests

Metal ion	Colour observed
Li^+	Scarlet red
Na^+	Yellow
K^+	Lilac
Rb^+	Red
Cs^+	Blue
Mg^{2+}	No flame colour
Ca^{2+}	Brick red
Sr^{2+}	Red
Ba^{2+}	Apple green

- Mg^{2+} does not have an observed colour because the energy emitted during a flame test involving magnesium is outside the visible spectrum

Your notes

Reactions & Ionic Equations

Test Tube Reactions

- Simple test tube reactions can be done to identify the following ions:
  - Ammonium ions (NH_4^+)
  - Carbonate ions (CO_3^{2-}) and hydrogencarbonate ions (HCO_3^-)
  - Sulfate ions (SO_4^{2-})
- If the sample to be tested is a solid, then it must be dissolved in deionised water and made into an aqueous solution

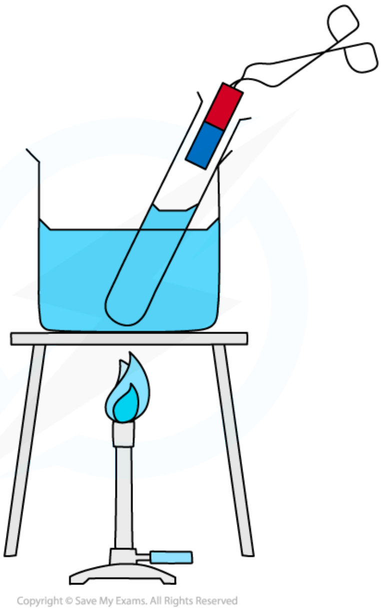
Testing for Ammonium Ions

- About 10 drops of a solution containing ammonium ions, such as ammonium chloride, should be added to a clean test tube
- About 10 drops of sodium hydroxide should be added using a pipette

Overall equation: **NH_4Cl (aq) + NaOH (aq) → NH_3 (g) + H_2O (l) + NaCl (aq)**

Ionic equation: **NH_4^+ (aq) + OH^- (aq) → NH_3 (g) + H_2O (l)**

- The test tube should be swirled carefully to ensure that it is mixed well
- The test tube of the solution should then be placed in a beaker of water, and the beaker of water should be placed above a Bunsen burner, so that it can become a water bath
- As the solution is heated gently, fumes will be produced
- A pair of tongs should be used to hold a damp piece of red litmus paper near the mouth of the test tube, to test the fumes
- The red litmus paper will change colour and become blue in the presence of ammonia gas



Damp red litmus paper turning blue in the presence of ammonia gas

- Alternatively NH_3 can be tested for as it will form a white smoke if reacted with fumes of hydrogen chloride (from concentrated hydrochloric acid)
- The white smoke formed is ammonium chloride

$$NH_3\,(g) + HCl\,(g) \rightarrow NH_4Cl\,(s)$$

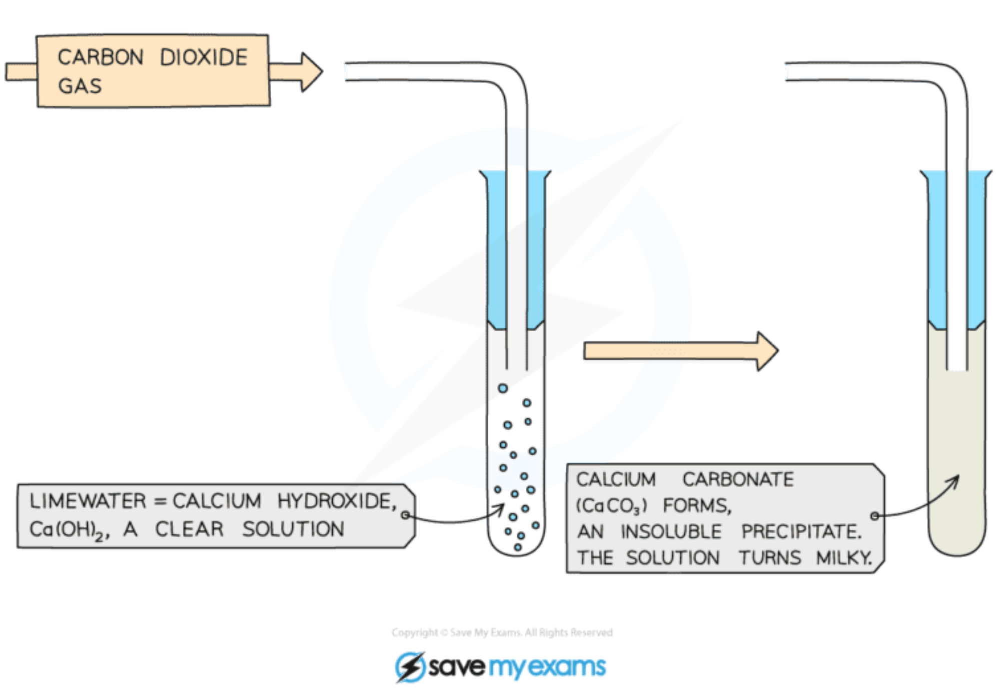
Testing for Carbonate Ions

- A small amount (around 1 cm^3) of dilute hydrochloric acid should be added to a test tube using a pipette
- An equal amount of sodium carbonate solution should then be added to the test tube using a clean pipette

Overall equation: $\mathbf{2HCl\,(aq) + Na_2CO_3\,(aq) \rightarrow 2NaCl\,(aq) + CO_2\,(g) + H_2O\,(l)}$

Ionic equation: $\mathbf{2H^+\,(aq) + CO_3^{2-}\,(aq) \rightarrow CO_2\,(g) + H_2O\,(l)}$

- As soon as the sodium carbonate solution is added, a bung with a delivery tube should be attached to the test tube
  - The delivery tube should transfer the gas which is formed into a different test tube which contains a small amount of limewater (calcium hydroxide solution)
- Carbonate ions will react with hydrogen ions from the acid to produce carbon dioxide gas
- Carbon dioxide gas will turn the limewater milky



When carbon dioxide gas is bubbled into limewater it will turn cloudy as calcium carbonate is produced

- A similar reaction is seen with sodium hydrogen carbonate, but the equations are:

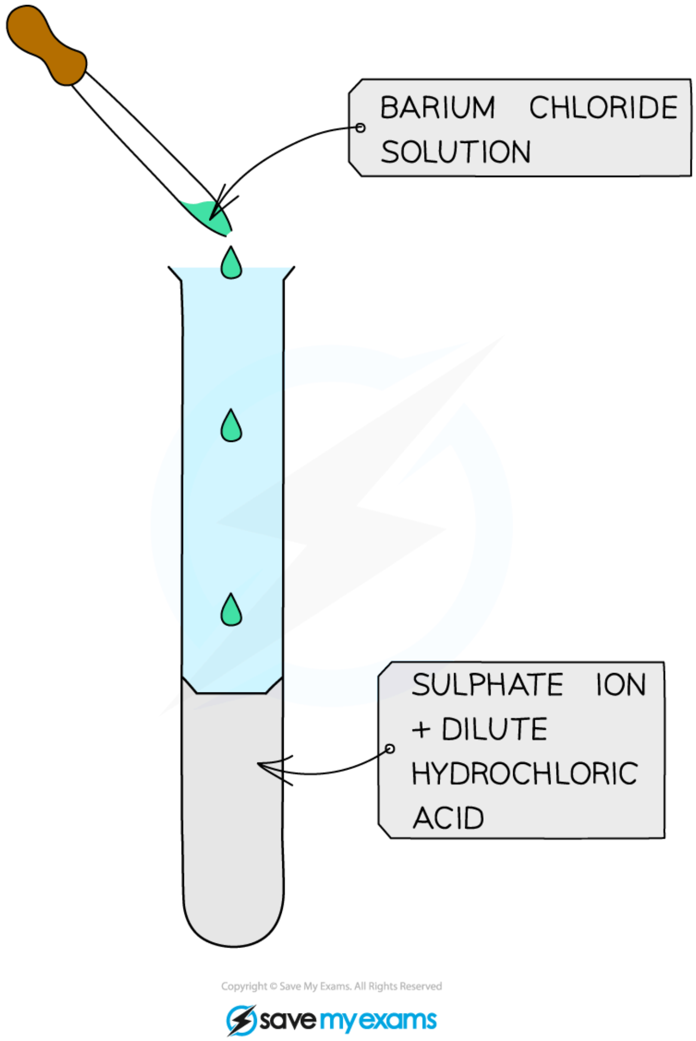
Overall equation: **$HCl\ (aq) + NaHCO_3\ (aq) \rightarrow NaCl\ (aq) + CO_2\ (g) + H_2O\ (l)$**

Ionic equation: **$H^+\ (aq) + HCO_3^-\ (aq) \rightarrow CO_2\ (g) + H_2O\ (l)$**

Testing for Sulfate Ions

Acidify the sample with dilute hydrochloric acid and then add a few drops of aqueous barium chloride If a sulfate is present then a **white** precipitate of barium sulfate is formed:

$Ba^{2+}\ (aq) + SO_4^{2-}\ (aq) \rightarrow BaSO_4\ (s)$



A white precipitate of barium sulfate is a positive result for the presence of sulfate ions

Examiner Tips and Tricks

HCl is added first to remove any carbonates which may be present and would also produce a precipitate and interfere with the results.