2.2 Group 2 – The alkaline earth metals

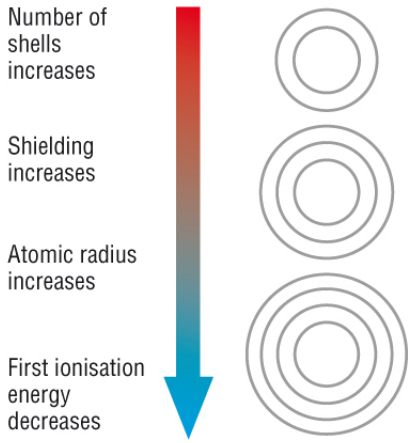
Atomic radius – Increases down the Group:

- **Shells:** **More** electron shells
- **Shielding:** **More** shielding (more inner shells)

Ionic radius – Smaller than atomic radius**:**

- **1 less shell:** **As 2e lost removing outer shell**

First ionisation - Decreases down the Group:

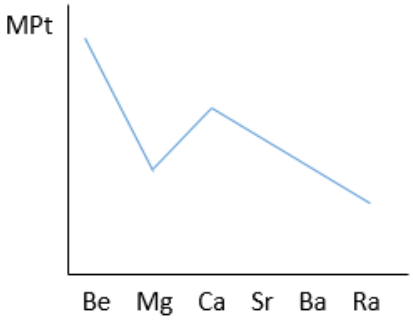


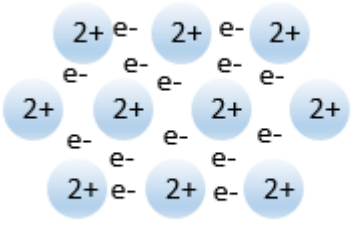
- **Shells:** **More** electron shells
- **Shielding:** **More** shielding (more inner shells)
- **No. Protons:** Number of protons **increases** but is **outweighed** by shells and shielding
- **Attraction:** Therefore attraction is **less**
- **Energy:** **Energy** required to remove an electron **decreases**

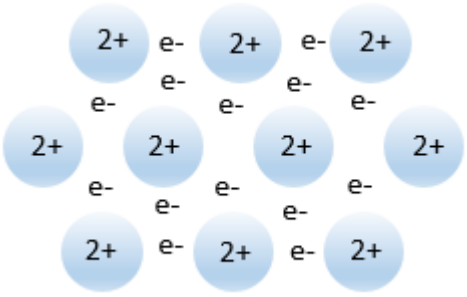
Reactivity – Increases down the Group:

- All **lose 2 electrons forming a 2+ ion** when they react
- **Ionisation energies decrease as you go down the group**
- **Electrons are lost more easily**
- **Reactivity increases** as you go **down the group**

Melting point – decreases down the group:







- As you go down Group 2 the **ionic radius increases**
- The 2+ charge from the nucleus is **further away** from the delocalised electrons
- **Attraction** is therefore **weaker**
- **Energy** required is **less**
- Mg's unusually low melting point comes from the different arrangement of the ions in the crystal structure.

Reactions of the Group 2 elements:

How the Group 2 elements react:

- Group 2 metals are reactive and all **lose 2e** when they react.
- As you go down Group 2 they become **more reactive**.
- This is due to the **decrease in Ionisation energies** as you go down the group.

$$\underset{0}{M} \rightarrow \underset{2+}{M^{2+}} + 2e^-$$

Reaction with water

- Group 2 metals react with water to give the hydroxide and hydrogen gas:

Metal + Water → Metal hydroxide + Hydrogen

$$\underset{0}{M_{(s)}} + 2H_2O_{(l)} \rightarrow \underset{+2}{M(OH)_{2(aq)}} + H_{2(g)}$$

$$\underset{0}{Ca_{(s)}} + 2H_2O_{(l)} \rightarrow \underset{+2}{Ca(OH)_{2(aq)}} + H_{2(g)}$$

Magnesium with steam:

- Magnesium reacts with steam to give the oxide and hydrogen gas:

$$\underset{0}{Mg_{(s)}} + H_2O_{(g)} \rightarrow \underset{+2}{MgO_{(aq)}} + H_{2(g)}$$

Solubility of the hydroxides and sulphates:

$Mg(OH)_2$	↓	Least soluble	$MgSO_4$	↑	Most soluble
$Ca(OH)_2$			$CaSO_4$		
$Sr(OH)_2$			$SrSO_4$		
$Ba(OH)_2$		Most soluble	$BaSO_4$		Least soluble

_____ charged negative ions tend to _____ in solubility as you go down the Group

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$Mg(OH)_2$ is said to be sparingly soluble

$BaSO_4$ is insoluble

1) Hydroxides:

- As **solubility increases, more OH^-** ions are released.
- This makes a **more alkaline** solution.
- The **pH increases** down the Group

Testing the solubility of the Group 2 hydroxides:

- This is done by **adding** hydroxide ions, **OH^-** to a solution of the Group 2 ion, **M^{2+}**:

$$M^{2+}_{(aq)} + OH^-_{(aq)} \rightarrow M(OH)_{2(s)}$$

- As $Mg(OH)_2$ is sparingly soluble, a thick white precipitate is formed.
- As $Ba(OH)_2$ is more soluble, a thin white precipitate will be formed.

2) Sulphates:

- Most are actually soluble ranging from sparingly soluble → soluble.
- **Barium sulphate** however is **insoluble**
- This therefore is used as the **chemical test** for the presence of **sulphate ions, SO_4^{2-}**

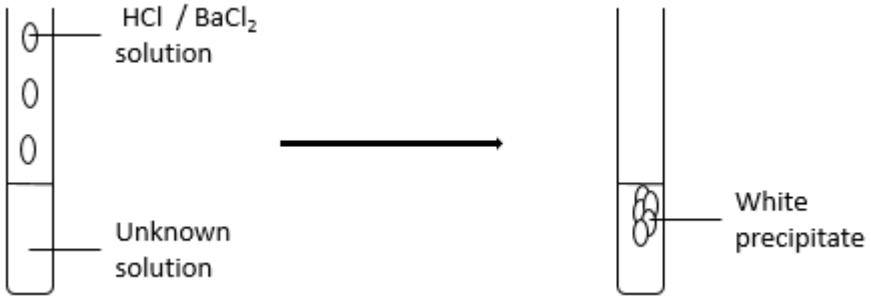
Testing the solubility of the Group 2 sulphates:

- This is done by **adding** sulphate ions, **SO_4^{2-}** or sulphuric acid, **H_2SO_4** to a solution of the Group 2 ion, **M^{2+}**:

$$M^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow MSO_{4(s)}$$

- As $MgSO_4$ is soluble, no precipitate will form.
- As $BaSO_4$ is insoluble, a white precipitate will be formed.

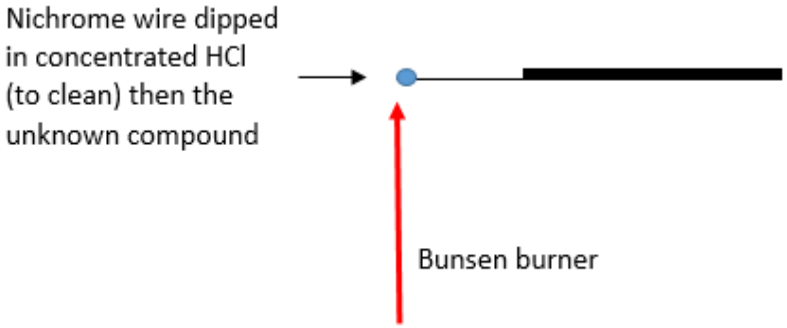
<u>Test for sulphate ions, SO_4^{2-}</u> (Part of Required practical 4)



- Add HCl first – This reacts and removes any sulphites or carbonates that may also give a white precipitate.
- Add $BaCl_2$ solution: If sulphates are present, a **white precipitate of $BaSO_4$** will form.

$$Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$$

Test for Group 2 metal ions - Flame tests (Part of Required practical 4)



Metal ion	*Flame colour*
Calcium, Ca^{2+}	**Brick red**
Strontium, Sr^{2+}	**Red**
Barium, Ba^{2+}	**Pale green**

Uses of Group 2 compounds:

1) Barium meals – X – Rays:

- Barium sulphate, $BaSO_4$, does not allow X – rays to pass through.
- Drinking a **suspension** of $BaSO_4$ coats the oesophagus, stomach or intestines (Barium meal)
- These now show up on an X – Ray allowing you to see any problems.
- Other Barium compounds are poisonous.
- Other Group 2 metal compounds are soluble so cannot be used.

2) Extraction of Titanium:

- TiO_2 is converted to $TiCl_4$ by heating with carbon and chlorine
- $TiCl_4$ is then reduced by Mg:

$$\underset{+4}{TiCl_{4(g)}} + \underset{0}{2Mg_{(l)}} \rightarrow \underset{0}{Ti_{(s)}} + \underset{+2}{2MgCl_{2(l)}}$$

- Ti has been **reduced** from **+4 → 0**
- Mg has been **oxidised** from **0 → +2**
- This makes magnesium a **reducing agent**

3) Removal of SO_2 from flue gases:

- SO_2 is produced burning fossil fuels to make electricity.
- This can be removed by reacting with an alkali such as CaO or $CaCO_3$ slurry (mixed with water)
- The process is called **wet scrubbing**:

$$CaO_{(s)} + 2H_2O_{(l)} + SO_{2(g)} \rightarrow CaSO_{3(aq)} + 2H_2O_{(l)}$$

$$CaCO_{3(s)} + 2H_2O_{(l)} + SO_{2(g)} \rightarrow CaSO_{3(aq)} + 2H_2O_{(l)} + CO_{2(g)}$$

4) Neutralising acids:

- Group 2 hydroxides are alkaline and therefore can be used to neutralise acids.
- $Ca(OH)_2$ is used to neutralise acidic soils.
- $Mg(OH)_2$ is used to neutralise excess stomach acids.