

*Topic 2.6*

***GROUP 2, THE ALKALINE EARTH METALS***

*Trends in size, first ionization energy and electronegativity*

*Trends in reaction with water*

*Trends in solubility of group 2 sulphates and hydroxides*

## PROPERTIES OF GROUP II ELEMENTS

The elements of group II are known as the **alkali earth metals**.

All of these elements are reactive metals. Magnesium and calcium are abundant in the earth's crust. The others are relatively rare.

Group II elements have the outer shell configuration  $ns^2$ .

### 1. Properties of individual atoms

a) atomic size

On descending a group, the number of shells increases. Therefore the outer shells are more shielded from the nucleus, are less closely held and move further away.

**Therefore the size of the atoms increases down group II.**

b) first ionisation energies

On descending a group, the number of shells increases. Therefore the outer shells are more shielded from the nucleus, are less closely held and are easier to remove.

**Therefore the first ionisation energies of the atoms decrease down group II.**

c) electronegativities

On descending a group, the number of shells increases. Therefore the outer shells are more shielded from the nucleus and bonded pairs of electrons are less strongly held.

**Therefore the electronegativities of the atoms decrease down group II.**

### 2. Melting points and electrical conductivity

The stronger the metallic bonding, the harder it is to separate the cations. This means that the melting point and hardness are higher if the metallic bonding is strong.

On descending a group, the size of the cations increases and the charge on the cations remains constant. The charge density thus decreases and the attraction between the cations and the delocalized electrons also decreases. The melting points and hardness therefore decrease.

**The melting points of the elements decrease down group II.**

In fact, there are discrepancies between magnesium and calcium and between strontium and barium. This is due to the change in crystal structure, which affects the distance between the ions and thus the strength of the metallic bonding. Be and Mg have hcp structures, Ca and Sr have fcc structures and Ba has a bcc structure.

Element	Size of cation/nm	T <sub>m</sub> /°C
Be (hcp)	0.027	1278
Mg (hcp)	0.072	649
Ca (fcc)	0.100	839
Sr (fcc)	0.113	769
Ba (bcc)	0.136	725

The overall trend is for **melting points** and **boiling points** to decrease down group II.

All group II elements will conduct electricity because the delocalised electrons are free to move throughout the solid.

### 3. Reaction of group II elements with water

Metals react with water to form metal hydroxides and hydrogen.



The vigour of these reactions depends on two factors:

- how easily two electrons can be removed from an atom
- how soluble the hydroxides of the metals are

Since ionisation energies decrease down a group, the electrons are more easily removed from the atoms and the reaction with water becomes increasingly vigorous on descending group II.

Since  $\text{Be(OH)}_2$  and  $\text{Mg(OH)}_2$  are insoluble in water, the hydroxide formed remains on the surface of the metal, preventing further reaction. Thus the reactions of magnesium and beryllium with water are very slow.  $\text{Ca(OH)}_2$ ,  $\text{Sr(OH)}_2$  and  $\text{Ba(OH)}_2$  are more soluble and thus the formation of the hydroxide does not prevent further reaction. So since hydroxide solubility increases down group II, so does the rate of reaction with water.

**Thus the reactivity to water increases down group II.**

Summary:

- Be does not react with water or steam.  
Mg does not react with water but reacts with steam. Since  $\text{Mg(OH)}_2$  actually decomposes at high temperatures into the oxide, MgO is the main product when magnesium reacts with steam:  
 $\text{Mg(s)} + \text{H}_2\text{O(g)} \rightarrow \text{MgO(s)} + \text{H}_2(\text{g})$
- Calcium reacts steadily with cold water. **The solution goes cloudy** as the hydroxide is not very soluble:  
 $\text{Ca(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{Ca(OH)}_2(\text{s}) + \text{H}_2(\text{g})$
- Strontium and barium react vigorously with cold water, barium even more vigorously than strontium:  
 $\text{Sr(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{Sr(OH)}_2(\text{aq}) + \text{H}_2(\text{g})$   
 $\text{Ba(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{Ba(OH)}_2(\text{aq}) + \text{H}_2(\text{g})$

## SOLUBILITY OF COMPOUNDS OF GROUP II ELEMENTS

The compounds formed by the elements of group II are predominantly ionic. The solubility of ionic compounds depends on the balance between the attraction of the oppositely charged ions to each other and the attraction of the separate ions to water.

If the attraction of the ions to each other is stronger than their attraction to water, the compounds will not be soluble. The solids will not dissolve in water and when the two ions are mixed together a precipitation reaction will take place.

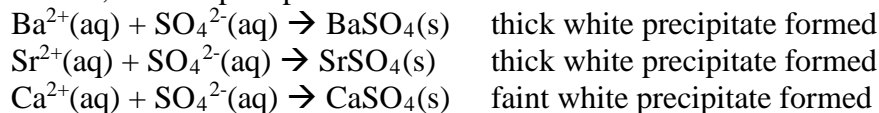
If the attraction of the ions to water is stronger than their attraction to each other, the compounds will be soluble. The solids will dissolve in water and when the two ions are mixed together no precipitation reaction will take place.

### i) trend in solubility of sulphates

**The solubility of the sulphates decreases down group II.**

**MgSO<sub>4</sub> is soluble, CaSO<sub>4</sub> is sparingly soluble and SrSO<sub>4</sub> and BaSO<sub>4</sub> are insoluble.**

If sulphuric acid or sodium sulphate is added to aqueous solutions of calcium, strontium or barium ions, a white precipitate will be formed:



If dilute sulphuric acid or sodium sulphate is added to a solution of Mg<sup>2+</sup> or Be<sup>2+</sup>, there is no reaction and no precipitate is formed.

The thick white precipitate formed when barium ions are added to sulphate ions can be used as a **test for sulphate ions**:

- add 1 cm<sup>3</sup> of aqueous barium chloride, and also 1 cm<sup>3</sup> of dilute hydrochloric acid, to 1 cm<sup>3</sup> of the unknown solution. If a thick white precipitate is formed, and that precipitate does not dissolve in dilute HCl, then sulphate or hydrogensulphate ions are present.

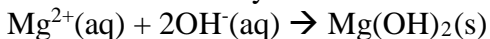
The hydrochloric acid removes any other ions such as carbonate or hydroxide ions which may affect the test.

## ii) trend in solubility of hydroxides

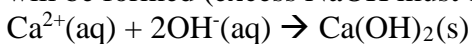
**The solubility of the group II hydroxides increases on descending the group.**

**Mg(OH)<sub>2</sub> is insoluble, Ca(OH)<sub>2</sub> is sparingly soluble and Sr(OH)<sub>2</sub> and Ba(OH)<sub>2</sub> are soluble.**

If dilute sodium hydroxide is added to a solution of Mg<sup>2+</sup> ions, a white precipitate will be formed immediately:



If dilute sodium hydroxide is added to a solution of Ca<sup>2+</sup> ions, a faint white precipitate will be formed (excess NaOH must be added).



If dilute sodium hydroxide is added to a solution of Sr<sup>2+</sup> or Ba<sup>2+</sup>, there is no reaction and no precipitate is formed.

## iii) uses of sulphates and hydroxides

**Magnesium hydroxide** is almost insoluble and is sold as a suspension in water. In this form it is known as “milk of magnesia”. It is taken to alleviate constipation.

**Calcium hydroxide** in solid form is known as “slaked lime” and is used to neutralize acidic soil.

**Barium sulphate** can be eaten as part of a “barium meal”. Barium is good at absorbing X-rays and so when the barium sulphate gets to the gut the outline of the gut can be located using X-rays. Although barium ions are very toxic, this technique is harmless because barium sulphate is completely insoluble.