

Elements :- Be, Mg, Ca, Sr, Ba, Ra

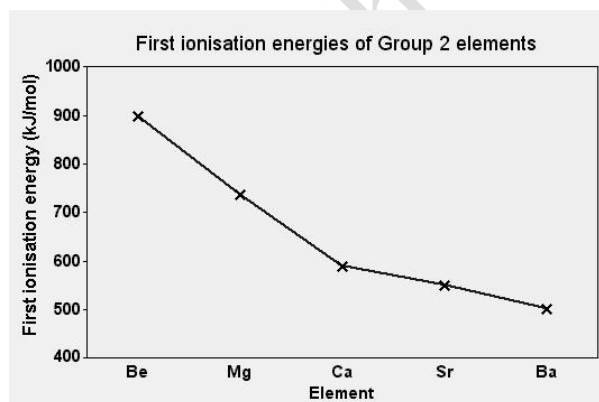
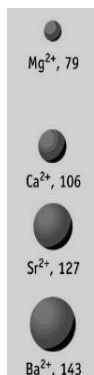
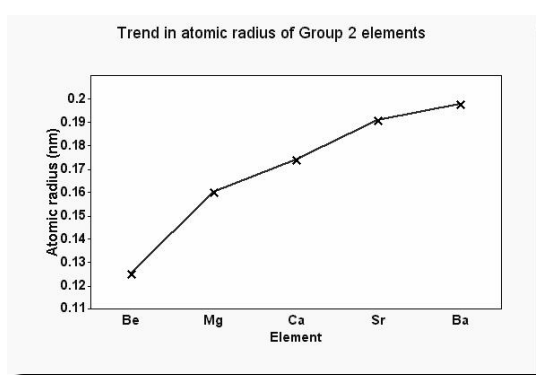
(Be and Ra are not discussed since Be shows ATYPICAL behaviour while Ra is RADIOACTIVE)

Members of group II are called **ALKALINE EARTH METALS**. They all have an oxidation number of +2 in their compounds. Their compounds are predominately ionic **NB Exception: compounds of beryllium are predominately covalent.** They are strong reducing agents and they form basic oxides.

Variation in atomic/ionic radii & ionisation energies

Atomic/ionic radii **INCREASE** on descending group II. This occurs as there are more core shells as one descends the group. This increases the distance between the nucleus and the valence shell. More core shells means a greater **shielding** effect between the nucleus and the valence electrons. This results in a weaker force of attraction between the nucleus and the valence electrons and the distance would therefore be greater. The residual attraction of the nucleus to valence electrons *after* shielding has taken place is called the *effective nuclear charge*.

Ionisation energy and atomic radii are **indirectly** related. On descending the group, the electrons are getting further away, this means the attraction gets weaker and therefore less energy is needed. Therefore ionisation energy **decreases down** the group.



Reactions of elements with oxygen, water and dilute acids

Equations of general reactions of group II metals using M to represent the metal atom

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

Reaction with cold water (exception Mg) $M(s) + H_2O(l) \rightarrow M(OH)_2(aq) + H_2(g)$

Reaction with dilute acids $M(s) + 2HCl(aq) \rightarrow MCl_2(aq) + H_2(g)$

	Reaction with oxygen	Reaction with water	Reaction with dilute acids e.g. HCl
Mg	It burns with a bright white light when either ignited by a flame or by electric current	It reacts very slowly with cold water but significantly with steam $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$	Reacts moderately producing effervescence
Ca	Surface becomes tarnished slowly at room temperature	It reacts slowly with cold water	Reacts rapidly producing effervescence
Sr	Surface becomes tarnished moderately at room temperature	It reacts rapidly with cold water	Reacts very rapidly producing effervescence
Ba	Surface becomes tarnished rapidly at room temperature	It reacts violently with cold water	Reacts violently producing effervescence

Note the change in reactivity (it increases) on descending the group for the metal and either oxygen, water or dilute acid

Thermal stability of group II carbonates/nitrates

Thermal stability of the nitrates/carbonates **increase DOWN** the group. As you descend the group, the metal ion gets **larger**.

This has two effects: (1) the charge density decreases and (2) the polarising power of the metal ion decreases.

What is charge density? The charge of an ion within a given area. It is related to ionic radius by the following formula:

$$\text{Charge density} = \text{charge} / \text{ionic radius}$$

What is polarising power? Remember, a cation would attract an oppositely charged species; like electrons. In an ionic compound a cation with a high charge density would cause the electron cloud of the anion to be attracted towards it. **The smaller the ionic radius, the higher the charge density, the greater the polarising power.**

By attracting the electron cloud of the anion **BACK** towards the cation, the ionic bond now possesses covalent character (*as electrons begin to reside in the middle of the nuclei of opposing atoms like covalent bonding*) and it is not as strong as a pure ionic bond and it can be **more easily** disrupted by the addition of heat.

Since the Mg^{2+} ion has the **smallest** ionic radius, it would have the **highest** charge density and would therefore cause the **greatest** disruption of the anion's electron cloud. Therefore magnesium compounds i.e. nitrates/carbonates would be the **LEAST** thermally stable. The Ba^{2+} ion being the **largest** ion, would have the **lowest** charge density and therefore would disrupt the electron cloud the **least**, and its compounds would be the **MOST** thermally stable.

NB The smaller the cation, the stronger the ionic bond with the small oxide ion would be (**because both the ions are small i.e. the metal ion and oxide ion**) and the more stable the compound.

Alternative explanation for thermal stability

The greater the difference in stability between the oxide and the carbonate/nitrate, the more readily the thermal decomposition would occur in carbonates or nitrates. Use the table below to help illustrate the point.

Pleas note that Group II carbonates decompose to the oxide and carbon dioxide gas AND Group II nitrates decompose to the oxide, nitrogen dioxide gas and oxygen gas

In all cases, decomposition occurs because of the greater stability of the resulting oxide

Group II carbonate/nitrate	Relative size of metal ion and anion in carbonate/nitrate	Relative size of metal ion and anion in oxide	Difference in stability of metal carbonate/nitrate and metal oxide	Rate of decomposition
$\text{MgCO}_3 / \text{Mg(NO}_3)_2$	Small & large	Small & small	LARGE	HIGH
$\text{CaCO}_3 / \text{Ca(NO}_3)_2$	Moderate & large	Moderate & small	MODERATE	MODERATE
$\text{BaCO}_3 / \text{Ba(NO}_3)_2$	Large & large	Large & small	SMALL	LOW

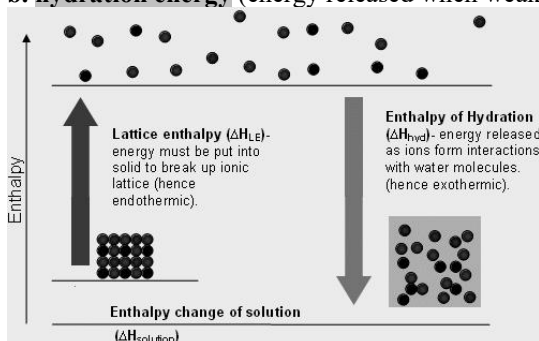
Trend of the solubilities of the group II sulphates

The sulphates become LESS SOLUBLE on descending the group

Two factors must be considered

a. **lattice energy** (energy required to cause the dissociation of the ionic compound into its respective ions)

b. **hydration energy** (energy released when weak bonds are formed from dative bonds between the water molecules and the ions)



Lattice energy **absorbs** energy decreases on descending the group as the **larger** the cation, the **weaker** the ionic bond.

Hydration energy **RELEASES** energy. When ions become hydrated, the smaller the ion, the **LARGER** the hydration energy (given the same charge) and vice versa. This is based on its charge density, the higher the charge density, the more water molecules are attracted to it and the more bonds are formed between the water molecules and the ion and the more energy released. The reverse is true as the ion gets larger. Therefore hydration energy decreases on descending the group.

Lattice energy is **NOT influenced to a great extent by the changing size of the cation IF THE ANION IS LARGE**, thus lattice energy of group II sulphate decreases at a slow rate.

Hydration energy is **greatly** influenced by changing size of the cation and would then decrease at a much faster rate than lattice energy. Since hydration energy decreases **FASTER** than lattice energy, the tendency for the compound to dissolve **DECREASES** as well.

If the hydration energy released does not compensate for the lattice energy input, the compound WILL NOT DISSOLVE.

The trend for group II hydroxides is reversed as the anion is SMALL and thus lattice energy decreases FASTER than hydration energy, causing the hydroxide to MORE soluble on descending the group.

A good point to note :- since the Be^{2+} ion is so small, it would have a VERY high charge density, this distorts the anion's electron cloud so much so that its compounds are predominately COVALENT in nature.

Uses of group II compounds

<i>Element/compound</i>	<i>Uses</i>
<i>Magnesium oxide</i>	It has a melting point of 2852 °C, therefore it is a suitable material to use as a refractory lining in furnaces.
<i>Magnesium hydroxide (milk of magnesia)</i>	Component used in laxatives
<i>Magnesium sulphate</i>	Hydrated magnesium sulphates is Epsom salts, which can be used as a laxative or to sooth aching muscles.
<i>Calcium hydroxide (slaked lime)</i>	Used in agriculture to reduce the acidity of soils. It can also be used to make mortar by mixing it with water and sand (it hardens by the loss of water and the slow reaction with atmospheric carbon dioxide to form calcium carbonate) $\text{Ca(OH)}_2 + \text{CO}_2 \rightarrow \text{CaCO}_3 + \text{H}_2\text{O}$
<i>Calcium carbonate (limestone)</i>	Used as a component in eggshells. Also used in the extraction of iron from its ores to remove impurities. Used to manufacture cement (clay and limestone are heated together in a furnace). Also used to form calcium hydrogensulphate which is used in the paper-making industry. Also used as a building material.
<i>Calcium sulphate</i>	Can also be used to make cement by heating it with coke and then clay. Also used to make plaster for buildings, in a purer form, it is known as Plaster of Paris.
Barium sulphate and barium carbonate	Barium sulphate used in the X-ray imaging of digestive system and barium carbonate is used as a rat poison

END OF GROUP II ELEMENTS

Worksheet

1.

- (a) The labels are accidentally removed from the bottles of four Group II metal sulphates (Mg, Ca, Sr, Ba). Outline a plan by which a chemist, who does not have access to flame test rods, could use physical properties to determine the identity of these sulphates. [6]
- (b) The reaction of magnesium in air when heated was used by photographers in days gone by.
- (i) Describe the environmental conditions under which the photographer would have decided to use this reaction.

[1 mark]

- (ii) Suggest ONE disadvantage of using the reaction described above.

[1 mark]

- (iii) Write the chemical equation to illustrate the reaction described above.

[2 marks]

Total 10 marks

2.

The following observations are made about elements in a group in the periodic table.

Observation 1:

A crystalline metal nitrate melts on gentle heating and decomposes at about 200 °C to produce the metal oxide residue and a gaseous mixture that includes a reddish brown gas.

Observation 2:

A second metal nitrate, that is produced from a metal within the same group, is anhydrous and requires temperatures over 800 °C to produce the metal oxide and the gaseous mixture.

[All the nitrates of elements in this group are decomposed to produce the oxide as the solid residue.]

- (i) Suggest an explanation for Observations 1 and 2.

[2 marks]

- (ii) Write a balanced equation for the reaction stated in EITHER Observation 1 OR Observation 2.

[2 marks]

3. a) Describe the behaviour of the metals barium and magnesium with the following

- being exposed to air on a watch glass [2]
- dropped in a beaker of water [2]

b) Using the electronic structure of the two metals magnesium and barium, explain the difference in their behaviour as mentioned in part a) above. [4]