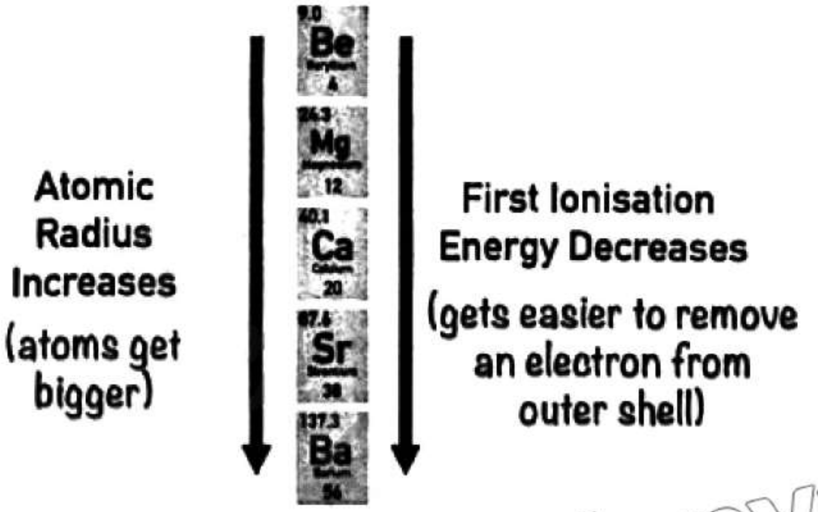
UNIT 05



GROUP 2 ELEMENTS

Student Learning Outcomes (SLOs)

- Describe the properties and trends of group 2 elements, including their electron configuration, reactivity, and common compounds such as oxides, hydroxides and carbonates.
- Explain the chemical reactivity of Group 2 elements, including their reactions with oxygen, water and acids (Be, Mg, Ca).
- Explain the reactivity of group 2 elements in terms of their electron configuration and valence electrons.
- Describe the industrial and everyday use of group 2 compounds, including their role in medicine and agriculture.
- Explain the term reactivity series and its application in predicting the outcome of chemical reactions.
- Use the term reactivity series and its application in predicting the outcome of chemical reactions.

- Explain the extraction and purification process of group II elements and their compounds.
- Understand the term thermal decomposition and its application in the analysis of group 2 compounds, especially carbonates and nitrates.
- Use the term thermal decomposition and its application in the analysis of group 2 compounds, especially carbonates and nitrates.
- Explain the trend in solubility of group II sulphates and hydroxides using the terms enthalpy of hydration and enthalpy of solution.
- Compare the properties and reactivity of group 2 elements with group 1 in the periodic table.
- Explain the term complex ion and its application in the formation of group 2 compounds.
- Explain the term basic oxide and its application in the formation of group 2 compounds.
- Describe qualitatively the trend in the thermal stability of nitrates and carbonates including the effect of ionic radius on the polarization of the large anion.
- Describe qualitatively the variation in solubility and enthalpy change of solution, ΔH sol, of hydroxides and sulphates in terms of relative magnitudes of the enthalpy change of hydration and the lattice energy.

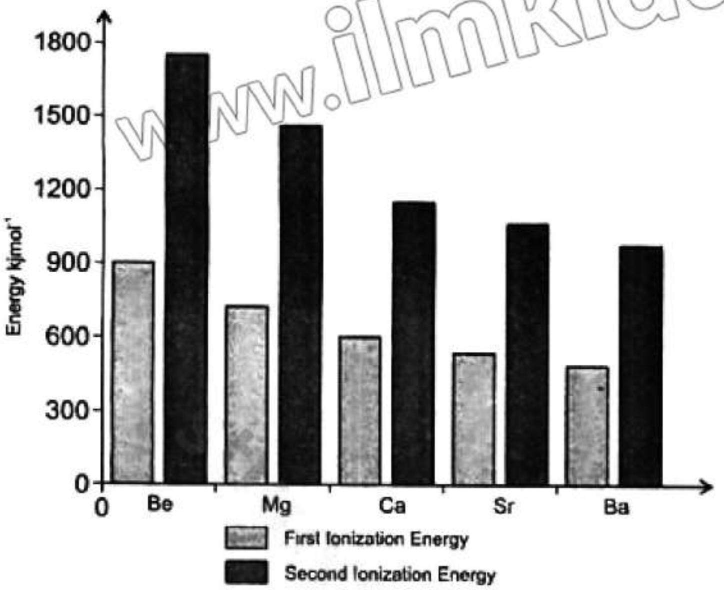
Elements from Group 2 are used in a wide range of applications. For example, Group 2 metals produce coloured flames when heated, leading to their use in flares and fireworks. Magnesium in powdered form is used in flares.

The elements in Group 2 of the Periodic Table are referred to as alkaline earth metals because their oxides and hydroxides are water-soluble (alkaline) in nature and oxides are found in the earth's crust. As they are in Group 2, these elements have two electrons in their outermost shell. These two outer electrons occupy s subshell. Their general electronic configuration is ns^2. Electronic configurations of the first three elements in Group 2 are:

Beryllium (Be)	$1s^2 2s^2$
Magnesium (Mg)	$1s^2 2s^2 2p^6 3s^2$
Calcium (Ca)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

5.1 General Trends of Physical Properties and Chemical Reactivity of Group 2 Elements

All Group 2 metals can form ionic compounds in which they donate two outermost electrons to become an ion with +2 charge. So, they act as reducing agents as they get oxidised. When going down the group, the metals become more reactive. This can be explained by looking at the Group 2 ionisation energies in the figure 5.1.



Decreasing order of 1st ionization energy of group 2 elements is

Be > Mg > Ca > Sr > Ba

Fig 5.1: Trend of first and second ionization energies of group 2 elements

5.1.1 Trends in ionization energy and chemical reactivity

Ionization energy is the minimum amount of energy required to remove one mole of electron from the isolated gaseous atom or ion. This graph shows that both the first and second ionization energies decrease going down the group. The graph above shows that going down the group, it becomes easier to remove the outer two electrons of the metals. This is due to the increased shielding effect and a larger distance between the outermost electrons and nucleus. As a result of this, the elements become more reactive going down the group as it gets easier for the atoms to lose two electrons and become 2+ ions. This trend is shown by looking at the reactions of the Group 2 metals:

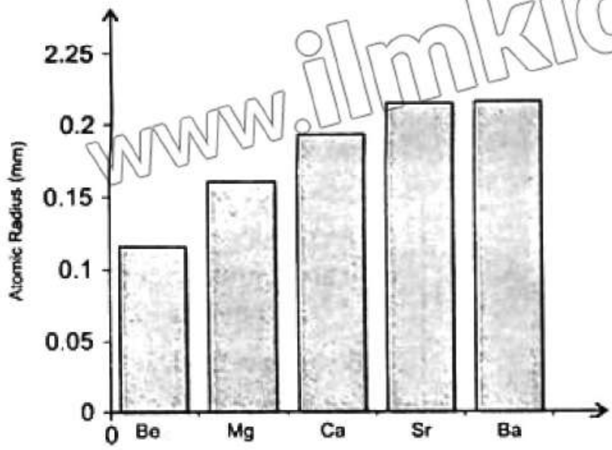
With dilute hydrochloric acid: bubbles of hydrogen gas are given off much faster indicating that the reactions become more vigorous. For example:

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

With oxygen: the metals get more reactive with oxygen down the group (Ba is so reactive that it must be stored in oil to prevent it from reacting with oxygen in air.

5.1.2 Trends in atomic radius

Look at the atomic radii of the Group 2 elements, shown in Table 5.1. The atoms of Group 2 elements get larger going down the group. This is because of the increase in extra shell and shielding effect, attraction between the nucleus and valence electrons decreases. Hence, size of atom increases. Figure 5.2 represents the atomic radius of group 2 elements.



Increasing order of atomic radius of group 2 elements is

Be < Mg < Ca < Sr < Ba

Fig 5.2: Trend of Atomic radius of group 2 elements

5.1.3 Trends in melting point

The melting point of the elements decreases going down the group as the outer electrons get further away from the nucleus. This means that the attraction between the nucleus and the delocalised electrons decreases hence strength of metallic bond decreases causing a decrease in melting point. Figure 5.3 represents the melting points of group 2 elements.

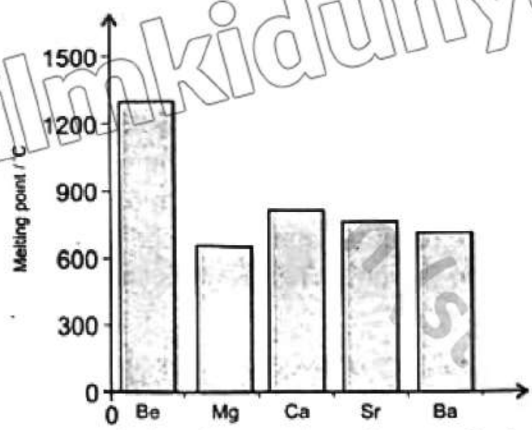


Fig 5.3: Trend of melting point of group 2 elements

Following table 5.1 represents the general trends in physical properties, such as atomic radius, melting point and density.

Table 5.1: Physical properties of alkaline earth metals

Group 2 Element	Metallic radius/nm	Atomic number	Melting point/°C	Density/gcm^{-3}
beryllium (Be)	0.122	4	1280	1.85
magnesium (Mg)	0.160	12	650	1.74
calcium (Ca)	0.197	20	838	1.55
strontium (Sr)	0.215	38	768	2.66
barium (Ba)	0.217	56	714	3.56

5.2 Trends in Chemical Reactivity with Oxygen, Water and Dilute Acids

(a) Reaction with Oxygen/ Formation of Basic oxide

The Group 2 metals burn in air, and more rapidly in oxygen, forming white solid oxides. Strontium and Barium also form peroxides (MO_2). For example, magnesium ribbon burns with a bright white flame once ignited in a Bunsen flame.

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

Beryllium and magnesium oxides are insoluble in water while the solubility of remaining oxides increases down the group, as they form soluble hydroxides.

All oxides of group 2 elements are basic in character except BeO which is amphoteric in nature as it reacts both with acid and base. Basic character increases on moving down the group.

$$BeO_{(s)} + 2HCl_{(aq)} \rightarrow BeCl_{2(aq)} + H_2O_{(l)}$$

$$BeO_{(s)} + 2NaOH_{(aq)} \rightarrow Na_2BeO_{2(aq)} + H_2O_{(l)}$$

The Group 2 metals get more reactive with oxygen going down the group. The larger atoms lose their outer two electrons more readily than the smaller atoms in the group. The greater reactivity of barium metal is illustrated by the fact that it must be stored under oil to keep it out of contact with air.

BeO is covalent in character while remaining oxides are basic in nature. This is due to the small size of Be^{2+} ion. Polarizing power of Be^{2+} is greater because of which oxide ion can easily be polarised. Hence electrons are shared by the nuclei of both atoms. That is why BeO is covalent in character.

Flame test

Some of the Group 2 metals burn with characteristic flame colours. It is the 2+ ions formed in the reaction that cause the colours. We can test for calcium, strontium and barium in compounds using flame tests. A nichrome wire, cleaned with concentrated hydrochloric acid, is dipped into a sample of the salt to be tested and heated in a non-luminous Bunsen flame:

- calcium compounds give a brick-red colour.
- strontium compounds give a scarlet/red colour
- barium compounds give an apple-green colour

(b) Reaction with water

Beryllium does not react with water because of the protective layer of BeO on its surface.

Magnesium does not react with cold water. However, it burns in steam to forms magnesium oxide and hydrogen.

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

Calcium reacts more readily than magnesium, with water.

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

This reaction forms a cloudy white suspension of slightly soluble calcium hydroxide. The calcium hydroxide dissolves make the solution weakly alkaline. The hydrogen gas is given off at a steady rate. Going down the group, reactivity of group II elements with water increases as hydrogen gas is released more and more rapidly.

General reaction of all group 2 metals except Be and Mg with water follows the following general equation.

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

(c) Reaction with Dilute Acids

The Group 2 metals undergo redox reaction with dilute acids to form salts and hydrogen gas. The reactions become more vigorous and exothermic down the group. The reaction of all group 2 metals with dilute HCl follows the following general equation.

$$M_{(s)} + 2HCl_{(aq)} \rightarrow MCl_{2(aq)} + H_{2(g)}$$

(d) Reaction of group 2 oxides with acid

Group 2 oxides are basic oxides, these oxides are used in making useful compounds by their reaction with acid.

$$MgO_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_2O_{(l)}$$

$$CaO_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)}$$

Magnesium sulfate is used for short-term relief of constipation. It is also used as a soaking solution to relieve minor sprains, bruises, muscle aches or discomfort and joint stiffness. Calcium chloride is used as a solidifying agent in paint production, coagulant in the manufacture of rubber.

Concept Assessment Exercise 5.1

1. What is the general trend in the melting points going down Group 2?
2. Explain why the atoms in Group 2, as in any other group, get larger with increasing atomic number.
3. Write a balanced chemical equation, including state symbols, for the reaction of:
   i. strontium with oxygen
   ii. strontium oxide with hydrochloric acid.
4. Barium reacts vigorously when added to water.
   i. Write a balanced chemical equation, including state symbols, for the reaction of barium with water.
   ii. State two observations that could be made during this reaction.
   iii. Suggest the approximate pH of the resulting solution.
   iv. Will the reaction be more or less vigorous than the reaction of barium with water? Explain your answer.
5. Describe what you would see when magnesium reacts with cold water and steam. Also write an equation for the reaction with steam.

5.3 Comparison of Reactivity of Group 1 and Group 2 Element

Reactivity of an element depends upon the ionization energy. Smaller the value of ionization energy, greater is the reactivity. Alkaline earth metals have higher ionization energy as compared to alkali metals. It is because of their smaller size and greater attraction of electrons towards nucleus. Group II elements are less reactive than Group I elements in the same period because of low ionization energy of group I elements as compared to group II.

Alkali metals react vigorously with water. The reaction is so vigorous that the evolved hydrogen catches fire. While alkaline earth metal reacts slowly with water.

$$Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(l)}$$

$$Ca_{(s)} + 2H_2O \rightarrow Ca(OH)_{2(aq)} + H_{2(l)}$$

Alkali metals are good reducing agents compared to alkaline earth metals because they possess smaller value of I.E.

5.4 Uses of Group 2 Compounds

Use of Limestone in industry

Limestone is made up mainly of calcium carbonate. There are many types of limestone, which provide useful rocks for building. They can be shaped into blocks that can be stuck to each other using mortar. Previously, this mortar was made using lime and sand.

Now it is more usual to use cement and sand, although the cement is made from lime. Marble is another form of calcium carbonate used as a building material, for example in making of expensive tiles. However, most calcium carbonate is used to make cement. The first stage in the manufacture of cement is the roasting of limestone in a lime kiln. At high temperatures in the kiln, calcium carbonate decomposes to form calcium oxide also called lime or quick lime.

$$CaCO_{3(s)} \xrightarrow{\text{Heat}} \underset{\text{lime}}{CaO_{(s)}} + CO_{2(g)}$$

The calcium oxide made in the lime kiln goes on to be roasted with clay to make cement. Cement can be mixed with sand and small pieces of rock to make concrete, the most widely used building material in the world. Its tensile strength can be improved by letting the concrete set with iron rods running through it.

Use of slaked lime in agriculture

Calcium hydroxide (slaked lime) is used in agriculture to increase the pH of acidic soils. Calcium hydroxide is basic, so it will react with and neutralise acid, raising the pH of the soil.

$$Ca(OH)_{2(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}_{(aq)} + 2H_2O_{(l)}$$

Use of Group 2 compounds in medical

Barium sulphate ($BaSO_4$) is used in medicine because it absorbs X-rays strongly and is used to diagnose disorders of the intestines and stomach. As it is insoluble, and not absorbed into the blood stream from the gut.

Group 2 hydroxides and carbonates are bases and are often used as antacids to treat acid indigestion (an excess of hydrochloric acid in the stomach) by neutralisation. Gaviscon and Rennies use calcium carbonate and magnesium carbonate, while Milk of Magnesia is a suspension of magnesium hydroxide in water.

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

$$Mg(OH)_{2(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + 2H_2O_{(l)}$$

5.4 Thermal Decomposition of Group 2 Carbonates and Nitrates

Thermal decomposition is the breakdown of a compound into two or more different substances by the application of heat.

The carbonates and nitrates of group 2 elements decompose, when heated. The carbonates break down to form the metal oxide and give off carbon dioxide gas. For example:

$$MgCO_{3(s)} \xrightarrow{\text{Heat}} MgO_{(s)} + CO_{2(g)}$$

The temperature at which thermal decomposition takes place increases going down Group 2 carbonate.

The Group 2 nitrates also undergo thermal decomposition. For example:

$$2Ca(NO_3)_{2(s)} \xrightarrow{\text{Heat}} 2CaO_{(s)} + 4NO_{2(g)} + O_{2(g)}$$

A brown gas is observed when a Group 2 nitrate is heated. This is toxic nitrogen dioxide, NO_2 (nitrogen dioxide). Like carbonates, a higher temperature is needed to thermally decompose the nitrates as group 2 is descended.

5.4.1 Trend in thermal stabilities

Thermal stability of group 2 carbonates or nitrates increases down the group. The relative ease of thermal decomposition is shown by the values of the enthalpy changes of reaction. The more positive the enthalpy change, the more energy will be needed to decompose the metal carbonates.

Table 5.2: Enthalpy change of reaction values for the decomposition of some Group 2 carbonates.

Group 2 carbonate	Decomposition temperature (°C)	Enthalpy change of reaction ($KJmol^{-1}$)
Magnesium carbonate	540	+117
Calcium carbonate	900	+176
Strontium carbonate	1280	+238
Barium carbonate	1360	+268

We can explain this trend using ideas about ion polarisation:

The carbonate ion has a relatively large ionic radius, so it is easily polarised by a small highly charged cation. Size of Group 2 cations increases down the group:

$$Mg^{2+} < Ca^{2+} < Sr^{2+} < Ba^{2+}$$

The smaller the ionic radius of the cation, the better it is at polarising the carbonate ion. A small highly charged cation such as Mg^{2+} can attract electrons and distort a larger carbonate anion to such an extent that the bond formed has a considerable amount of covalent character. This is shown in figure 5.4.

Fig 5.4 Magnesium ions are better polarisers of carbonate ions than calcium ions.

So, the degree of polarisation of the carbonate ion by the group 2 cation follows the order

$$Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$$

the greater the polarisation of the carbonate ion, the easier it is to weaken a carbon-oxygen bond in the carbonate and form carbon dioxide and the metal oxide on heating. The order of stability of carbonates is

$$BaCO_3 > SrCO_3 > CaCO_3 > MgCO_3$$

A similar pattern is observed with the thermal decomposition of Group 2 nitrates. The order of stability with respect to the products is in the order:

$$Ba(NO_3)_2 > Sr(NO_3)_2 > Ca(NO_3)_2 > Mg(NO_3)_2$$

Concept Assessment Exercise 5.2

1. Which one of the three compounds listed will decompose at the lowest temperature?
   i. Calcium carbonate, strontium carbonate, barium carbonate
   ii. Barium nitrate, calcium nitrate, magnesium nitrate
2. Write a balanced chemical equation, including state symbols, for the thermal decomposition of:
   i. Strontium carbonate
   ii. Barium nitrate.

5.5 Trends in Solubility of the Group 2 Hydroxides and Sulphate

(a) Solubility of hydroxides of group II elements

The solubility of hydroxides increases as we move down the group, with barium hydroxide being highly soluble in water. The order of solubility is

$$Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$$

Table 5.3 shows the solubility in water of some Group 2 hydroxides. We can explain this variation in solubility in terms of the relative values of enthalpy change of hydration and the corresponding lattice energy.

Table 5.3 Solubilities in water of some group 2 hydroxides.

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

(b) Change in hydration enthalpy down the group

Hydration enthalpy is directly proportional to the charge on ion and inversely proportional to the size of ion. Smaller ions (with the same charge) have greater enthalpy changes of hydration. So, the enthalpy change of hydration decreases (gets lessexothermic) in the order.

$$Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$$

this decrease is relatively small down the group and it depends entirely on the increase in the size of the cation, as the anion is unchanged (it is the hydroxide ion in every case).

(c) Change in lattice energy down the group

Lattice energy is greater if the ions (with the same charge) forming the lattice are small so the lattice energy decreases in the order.

$$Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$$

The lattice energy is also inversely proportional to the sum of the radii of the anion and cation. So, the decrease in lattice energy is relatively large down the group and it is determined more by the size of the large cation ion than the size of the hydroxide.

(d) Difference in enthalpy change of solution of Group 2 hydroxide

Substances that have a very high solubility in water are likely to have enthalpy of solution with a high negative (exothermic) value. As a rough guide, the higher the negative value of enthalpy of solution the more soluble is the salt.

Conclusion:

The lattice energy of the hydroxides decreases by relatively higher values down the group. The enthalpy changes of hydration decreases (gets less exothermic) by relatively lower values down the group. The value of enthalpy of solution gets more exothermic down the group. so, the solubility of the Group 2 hydroxides increases down the group. Compound is likely to be soluble in water only if enthalpy of solution is negative or has a small positive value.

(e) Solubility of sulfates of group II elements

The solubility of sulphates decreases as we move down the group, with barium sulphate being insoluble in water. Table 5.4 shows the solubility in water of some Group 2 sulfates. The solubility decreases as the radius of the metal ion increases.

$$MgSO_4 > CaSO_4 > SrSO_4 > BaSO_4$$

Table 5.4 Solubilities in water of some group 2 sulphates.

Group 2 Sulphates	Solubility at 298 k (mol/dm^3)
$MgSO_4$	1.83
$CaSO_4$	4.66×10^{-2}
$SrSO_4$	7.11×10^{-4}
$BaSO_4$	9.43×10^{-6}

When we move down the group, both the lattice dissociation enthalpy and hydration enthalpy decrease. The hydration enthalpy decreases more than the lattice dissociation enthalpy. So, the enthalpy of the solution becomes more endothermic.

Concept Assessment Exercise 5.3

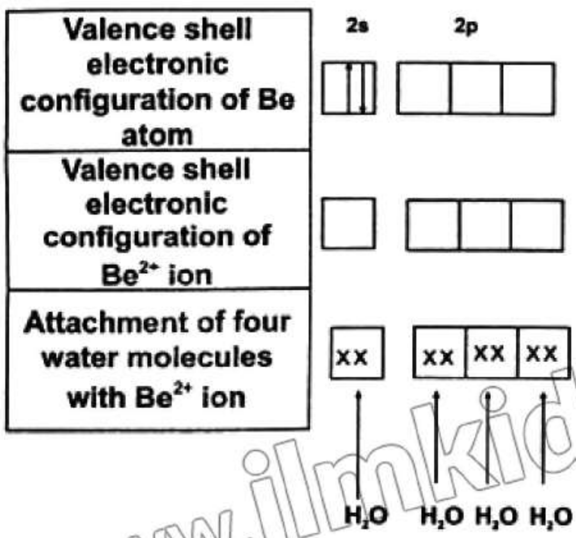
Explain why magnesium sulphates is more soluble than barium sulfate by referring to the relative values of the lattice energies and enthalpy changes of hydration.

5.6 Complexes of The Alkaline Earth Metals

Group 2 elements almost exclusively form ionic compounds containing the M^{2+} ion, They are more reactive towards group 5 elements, and they have a greater tendency to form complexes with Lewis bases than do the alkali metals.

Because of their higher positive charge (+2) and smaller ionic radii, the alkaline earth metals have a much greater tendency to form complexes with Lewis bases than do the alkali metals. This tendency is most important for the lightest cation (Be^{2+}) and decreases rapidly with the increasing radius of the metal ion.

The chemistry of Be^{2+} is dominated by its behaviour as a Lewis acid, forming complexes with Lewis bases that produce an octet of electrons around beryllium. For example, Be^{2+} salts dissolve in water to form acidic solutions that contain the tetrahedral $[Be(H_2O)_4]^{2+}$ ion. Because of its high charge-to-radius ratio, the Be^{2+} ion polarizes coordinated water molecules, thereby increasing their acidity.



$$[Be(H_2O)_4]^{2+} \longrightarrow [Be(H_2O)_3(OH)]^{1+} + H^+$$

Similarly, in the presence of a strong base, beryllium and its salts form the tetrahedral hydroxo complex $[Be(OH)_4]^{2-}$. Hence beryllium oxide is amphoteric. Beryllium also forms a very stable tetrahedral fluoride complex $[BeF_4]^{2-}$, $[BeF_3]^{1-}$. This is all because of small size and high charge density of Be^{2+} ion. Recall that beryllium halides behave like Lewis acids by forming adducts with Lewis bases.

The heavier alkaline earth metals also form few complexes, but usually with a coordination number of 6 or higher. Complex formation is most important for the smaller cations (Mg^{2+} and Ca^{2+}). Thus aqueous solutions of Mg^{2+} contain the octahedral $[Mg(H_2O)_6]^{2+}$ ion. In this complex Mg can extend their coordination number to six by using one 3s, three 3p and two 3d orbitals which are present in its outer most shell.

5.7 Extraction and Purification Process of Group 2 Elements and their Compounds

Group 2 elements, also known as alkaline earth metals, include beryllium, magnesium, calcium, strontium, barium, and radium.The extraction and purification processes of Group 2 elements and their compounds involve several steps tailored to the specific properties of each element.

1. Extraction:

a. Ore Preparation: Group 2 elements are typically found in minerals such as carbonates, sulfates, oxides, and silicates. The ore containing the desired metal is first mined and then crushed to obtain a fine powder in order to extract the elemental metal.

b. Roasting: Group 2 Carbonates are unstable towards heat. They thermally decompose to form metal oxide and carbon dioxide. For example, to extract magnesium from its carbonate first it is roasted to form magnesium oxide and CO_2.

$$MgCO_{3(s)} \xrightarrow{\text{Heat}} MgO_{(s)} + CO_{2(g)}$$

c. Reduction: Depending on the ore, reduction is often carried out to extract the metal. For example, magnesium can be extracted by the electrolysis of molten magnesium oxide, while calcium and strontium can be obtained by reducing their halides with sodium or magnesium.

2. Electrolysis (Purification):

Pure samples of most of the alkaline earth metals can be obtained by electrolysis of the chlorides or oxides. In some cases, electrolysis is used to further purify the extracted metal. Beryllium was first obtained by the reduction of its chloride; radium chloride, which is radioactive, was obtained through a series of reactions and separations.

Examples

Here is an overview of the extraction and purification methods for some key Group 2 elements:

1. Beryllium:

Extraction: Beryllium is usually extracted from beryl ore ($Be_3Al_2(SiO_3)_6$) through a chemical process involving acid digestion and solvent extraction.

Purification: It is purified by converting beryllium hydroxide into beryllium *fluoride*, which is then reduced by magnesium to obtain high-purity beryllium metal.

2. Magnesium:

Extraction: Magnesium is commonly produced by the electrolysis of molten magnesium chloride (from sea water or brines).

Purification: The magnesium obtained is further purified by fractional distillation or vacuum distillation processes.

Concept Assessment Exercise 5.4

1. Why beryllium form complex compounds?
2. Starting from calcium carbonate how would you extract pure calcium metal.

KEY POINTS

- The Group 2 elements magnesium to barium are typical metals with high melting points and they are good conductors of heat and electricity.
- Progressing down Group 2 from magnesium to barium, the atomic radius increases. This is due to the addition of an extra shell of electrons for each element as the group is descended.
- The Group 2 elements magnesium to barium react with water to produce hydrogen gas and the metal hydroxide, which may be only slightly soluble.
- The Group 2 elements magnesium to barium burn in air to form white solid oxides. These oxides form hydroxides with water. The hydroxides get more soluble in water going down the group so their solutions can become more alkaline.
- The sulfates of Group 2 elements get less soluble in water going down the group.
- Reactivity of the elements with oxygen or water increases down Group 2 as the first and second ionisation energies decrease.
- The Group 2 carbonates and nitrates get more resistant to thermal decomposition descending the group.
- Many of the compounds of Group 2 elements have important uses. Limestone, which contains mainly calcium carbonate, is used as a building material and is used to make cement, which is a component in the mixtures that make concrete and mortar. Slaked lime (calcium hydroxide) is used to neutralise acids in acidic soil.
- Pure samples of most of the alkaline earth metals can be obtained by electrolysis of the chlorides or oxides.

References for additional information.

- Inorganic Chemistry by R.L Madan and G.D Tuli
- Disha Expert- Disha NCERT Xtract errorless objective Chemistry
- Cambridge International A LevelChemistry by Lawrie Ryan and Roger Norris
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EXERCISE

1. Multiple Choice Questions (MCQs)

i. The oxidation state shown by alkaline earth metals is

a) +2 b) -2 c) +1,+2 d) -1, -2

ii. Which one of the following is the most soluble in water?

a) $Mg(OH)_2$ b) $Ca(OH)_2$ c) $Sr(OH)_2$ d) $Ba(OH)_2$

iii. Which of the following alkaline earth metal hydroxides is amphoteric in character

a) $Be(OH)_2$ b) $Sr(OH)_2$ c) $Ca(OH)_2$ d) $Ba(OH)_2$

iv. Of the metals Be, Mg, Ca and Sr of group 2 A. In the periodic table the least ionic chloride would be formed by

a) Be b) Ca c) Mg d) Sr

v. The order of solubility of sulphates of alkaline earth metals in water is

a) Be > Mg > Ca > Sr > Ba b) Mg > Be > Ba > Ca > Sr
c) Be > Ca > Mg > Ba > Sr d) Mg > Ca > Ba > Be > Sr

vi. The solubilities of carbonates decrease down the magnesium group due to a decrease in

a) hydration energies of cations b) inter-ionic attraction
c) entropy of solution formation d) lattice energies of solids

vii. In which of the following the hydration energy is higher than the lattice energy?

a) $MgSO_4$ b) $SrSO_4$ c) $RaSO_4$ d) $BaSO_4$

viii. Which of the following alkaline earth metal sulphates has hydration enthalpy higher than the lattice enthalpy?

a) $CaSO_4$ b) $BaSO_4$ c) $BeSO_4$ d) $SrSO_4$

ix. Which gas is released when $CaCO_3$ reacts with dilute HCl?

a) H_2 b) O_2 c) CO_2 d) Cl_2

x. Brick red is characteristics flame colour of

a) Beryllium b) Magnesium c) Calcium d) Barium

2. Short Answer Questions

i. Describe and explain the trend observed in the thermal stabilities of the carbonates of Group II.

ii. Describe the use of Group II elements or its compounds in agriculture.

iii. Write an equation to represent the thermal decomposition of calcium nitrate,

iv. Why are the elements of Group 2 called alkaline earth metals?

v. How do group 1 metals differ from Group 2 metals?

vi. Explain with the help of an equation the amphoteric nature of beryllium oxide.

vii. Explain qualitatively the variation in solubility of the hydroxides of the elements in Group II down the Group from magnesium to barium.

3. Long Answer Questions

i. Explain the reactivity of group II elements with water and dilute HCl with chemical reactions.

ii. Analyse the trends in the thermal stability of nitrates and carbonates of alkaline earth metals.