

## The s-block elements

The s-block elements of the Periodic Table are those in which the last electron enters the outermost s-orbital. As the s-orbital can accommodate only two electrons, two groups (1 & 2) belong to the s-block of the Periodic Table.

Group 1 of the Periodic Table consists of the elements: lithium, sodium, potassium, rubidium, caesium and francium. They are collectively known as the alkali metals.

These are so called because they form hydroxides on reaction with water which are strongly alkaline in nature.

The elements of Group 2 include beryllium, magnesium, calcium, strontium, barium and radium. These elements with the exception of beryllium are commonly known as the alkaline earth metals. These are so called because their oxides and hydroxides are alkaline in nature and these metal oxides are found in the earth's crust\*.

Among the alkali metals sodium and potassium are abundant and lithium, rubidium and caesium have much lower abundances. Francium is highly radioactive; its longest-lived isotope  $^{223}\text{Fr}$  has a half-life of only 21 minutes. Of the alkaline earth metals calcium and magnesium ranks fifth and sixth in abundance respectively in the earth's crust. Strontium and barium have much lower abundances. Beryllium is rare and radium is the rarest of all comprising only  $10^{-10}$  per cent of igneous rocks†

The general electronic configuration of s-block elements is [noble gas] $ns^1$  for alkali metals and [noble gas]  $ns^2$  for alkaline earth metals.

Lithium and beryllium, the first elements of Group 1 and Group 2 respectively exhibit some properties which are different from those of the other members of the respective group. In these anomalous properties they resemble the second element of the following group.

Thus, lithium shows similarities to magnesium and beryllium to aluminium in many of their properties. This type of diagonal similarity is commonly referred to as **diagonal relationship** in the periodic table.

The diagonal relationship is due to the similarity in ionic sizes and /or Charge/radius ratio of the elements.

Monovalent sodium and potassium ions and divalent magnesium and calcium ions are found in large proportions in biological fluids.

These ions perform important biological functions such as maintenance of ion balance and nerve impulse conduction.

### GROUP 1 ELEMENTS: ALKALI METALS

The alkali metals show regular trends in their physical and chemical properties with the increasing atomic number. The atomic, physical and chemical properties of alkali metals are discussed below.

#### Electronic Configuration

All the alkali metals have one valence electron,  $ns^1$  outside the noble gas core.

The loosely held s-electron in the outermost valence shell of these elements makes them the most electropositive metals.

They readily lose electron to give monovalent  $M^+$  ions. Hence they are never found in Free State in nature.

Element	Symbol	Electronic configuration
Lithium	Li	$1s^2 2s^1$ or [He] $2s^1$
Sodium	Na	$1s^2 2s^2 2p^6 3s^1$ or [Ne] $3s^1$
Potassium	K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ or [Ar] $4s^1$
Rubidium	Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$ or [Kr] $5s^1$
Caesium	Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2 4d^{10} 5p^6 6s^1$
		or [Xe] $6s^1$
Francium	Fr	[Rn] $7s^1$

#### Atomic and Ionic Radii

The alkali metal atoms have the largest sizes in a particular period of the periodic table. With increase in atomic number, the atom becomes larger. The monovalent ions ( $M^+$ ) are smaller than the parent atom. The atomic and ionic radii of alkali metals increase on moving down the group i.e., they increase in size while going from Li to Cs.

## Ionization Enthalpy

The ionization enthalpies of the alkali metals are considerably low and decrease down the group from Li to Cs. This is because the effect of increasing size outweighs the increasing nuclear charge, and the outermost electron is very well screened from the nuclear charge.

## Hydration Enthalpy

The hydration enthalpies of alkali metal ions decrease with increase in ionic sizes.



$\text{Li}^+$  has maximum degree of hydration and for this reason lithium salts are mostly hydrated, e.g.,  $\text{LiCl} \cdot 2\text{H}_2\text{O}$

## Physical Properties

1. All the alkali metals are silvery white, soft and light metals.
2. Because of the large size, these elements have low density which increases down the group from Li to Cs. However, potassium is lighter than sodium.
3. The melting and boiling points of the alkali metals are low indicating weak metallic bonding due to the presence of only a single valence electron in them.
4. The alkali metals and their salts impart characteristic colour to an oxidizing flame. This is because the heat from the flame excites the outermost orbital electron to a higher energy level. When the excited electron comes back to the ground state, there is emission of radiation in the visible region as given below:

Metal	Li	Na	K	Rb	Cs
Colour	Crimson red	Yellow	Violet	Red violet	Blue
$\lambda/\text{nm}$	670.8	589.2	766.5	780.0	455.5

Alkali metals can therefore, be detected by the respective flame tests and can be determined by flame photometry or atomic absorption spectroscopy. These elements when irradiated with light, the light energy absorbed may be sufficient to make an atom lose electron. This property makes caesium and potassium useful as electrodes in photoelectric cells.

## Chemical Properties

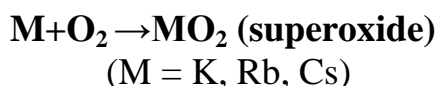
The alkali metals are highly reactive due to their large size and low ionization enthalpy. The reactivity of these metals increases down the group.

### (i) Reactivity towards air:

The alkali metals tarnish in dry air due to the formation of their oxides which in turn react with moisture to form hydroxides.

They burn vigorously in oxygen forming oxides.

Lithium forms monoxide, sodium forms peroxide, the other metals form superoxides. The superoxide  $O_2^-$  ion is stable only in the presence of large cations such as K, Rb, Cs.



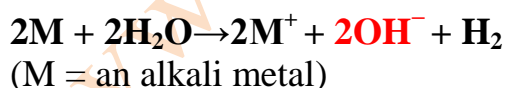
In all these oxides the oxidation state of the alkali metal is +1.

Lithium shows exceptional behaviour in reacting directly with nitrogen of air to form the nitride,  $Li_3N$  as well.

Because of their **high reactivity towards air and water**, they are normally kept in **kerosene oil**.

### (ii) Reactivity towards water:

The alkali metals react with water to form hydroxide and dihydrogen.



It may be noted that although lithium has most negative  $E^0$  value its reaction with water is less vigorous than that of sodium which has the least negative  $E^0$  value among the alkali metals. This behaviour of lithium is attributed to its small size and very high hydration energy.

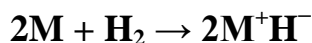
Other metals of the group react explosively with water.

They also react with proton donors such as alcohol, gaseous ammonia and alkynes.

### (iii) Reactivity towards dihydrogen:

The alkali metals react with dihydrogen at about 673K (lithium at 1073K) to form hydrides.

All the alkali metal hydrides are ionic solids with high melting points.



### (iv) Reactivity towards halogens :

The alkali metals readily react vigorously with halogens to form ionic halides,  $M^+X^-$ . However, lithium halides are somewhat covalent. It is because of the high polarisation capability of lithium ion (*The distortion of electron cloud of the anion by the cation is called polarisation*).

The  $Li^+$  ion is very small in size and has high tendency to distort electron cloud around the negative halide ion. Since anion with large size can be easily distorted, among halides, lithium iodide is the most covalent in nature.

### Uses:

1. Lithium metal is used to make useful alloys, for example with lead to make 'white metal' bearings for motor engines, with aluminium to make aircraft parts, and with magnesium to make armour plates.
2. Lithium is used in thermonuclear reactions.
3. Lithium is also used to make electrochemical cells.
4. Sodium is used to make a Na/Pb alloy needed to make  $PbEt_4$  and  $PbMe_4$ . These organolead compounds were earlier used as anti-knock additives to petrol, but nowadays vehicles use lead-free petrol.
5. Liquid sodium metal is used as a coolant in fast breeder nuclear reactors.
6. Potassium has a vital role in biological systems.
7. Potassium chloride is used as a fertilizer.
8. Potassium hydroxide is used in the manufacture of soft soap.
9. Potassium hydroxide is also used as an excellent absorbent of carbon dioxide.
10. Caesium is used in devising photoelectric cells.

### WHY LITHIUM SHOWS ANOMALOUS PROPERTIES

#### REASON FOR ANOMALOUS BEHAVIOUR

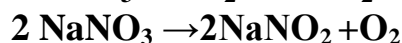
The anomalous behavior of lithium is due to the:

- (i) exceptionally small size of its atom and ion, and
- (ii) high polarizing power (i.e., charge/ radius ratio)

As a result, there is increased covalent character of lithium compounds which is responsible for their solubility in organic solvents.

### ANOMALOUS PROPERTIES OF LITHIUM

- (i) Lithium is much harder. Its m.p. and b.p. are higher than the other alkali metals.
- (ii) Lithium is least reactive but the strongest reducing agent among all the alkali metals
- (iii) On combustion in air it forms mainly monoxide,  $\text{Li}_2\text{O}$  and the nitride,  $\text{Li}_3\text{N}$  unlike other alkali metals.
- (iv)  $\text{LiCl}$  is deliquescent and crystallises as a hydrate,  $\text{LiCl}\cdot 2\text{H}_2\text{O}$  whereas other alkali metal chlorides do not form hydrates.
- (v) Lithium hydrogencarbonate is not obtained in the solid form while all other elements form solid hydrogencarbonates.
- (vi) Lithium unlike other alkali metals forms no ethynide on reaction with ethyne.
- (vii) Lithium nitrate when heated gives lithium oxide,  $\text{Li}_2\text{O}$ , whereas other alkali metal nitrates decompose to give the corresponding nitrite.



- (vii)  $\text{LiF}$  and  $\text{Li}_2\text{O}$  are comparatively much less soluble in water than the corresponding compounds of other alkali metals.

### Points of Similarities between Lithium and Magnesium /Diagonal Relationship

The similarity between lithium and magnesium is particularly striking and arises because of their similar sizes

Atomic radii,  $\text{Li} = 152 \text{ pm}$ ,  $\text{Mg} = 160 \text{ pm}$ ; ionic radii :  $\text{Li}^+ = 76 \text{ pm}$ ,  
 $\text{Mg}^{2+} = 72 \text{ pm}$ .

The main points of similarity are:

- (i) Both lithium and magnesium are harder and lighter than other elements in the respective groups.
- (ii) Lithium and magnesium react slowly with water. Their oxides and hydroxides are much less soluble and their hydroxides decompose on heating. Both form a nitride,  $\text{Li}_3\text{N}$  and  $\text{Mg}_3\text{N}_2$ , by direct combination with nitrogen.
- (iii) The oxides,  $\text{Li}_2\text{O}$  and  $\text{MgO}$  do not combine with excess oxygen to give any superoxide.
- (iv) The carbonates of lithium and magnesium decompose easily on heating to form the oxides and  $\text{CO}_2$ . Solid hydrogen carbonates are not formed by lithium and magnesium.

(v) Both LiCl and MgCl<sub>2</sub> are soluble in ethanol.

(vi) Both LiCl and MgCl<sub>2</sub> are deliquescent and crystallize from aqueous solution as hydrates, LiCl·2H<sub>2</sub>O and MgCl<sub>2</sub>·8H<sub>2</sub>O.

### SOME IMPORTANT COMPOUNDS OF SODIUM

Sodium carbonate, Sodium hydroxide, Sodium chloride and Sodium bicarbonate

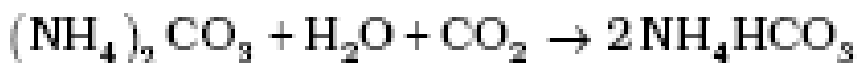
#### **Sodium Carbonate (Washing Soda), Na<sub>2</sub>CO<sub>3</sub>·10H<sub>2</sub>O**

Sodium carbonate is generally prepared by Solvay Process.

In this process, advantage is taken of the low solubility of sodium hydrogencarbonate whereby it gets precipitated in the reaction of sodium chloride with ammonium hydrogencarbonate.

The latter is prepared by passing CO<sub>2</sub> to a concentrated solution of sodium chloride saturated with ammonia, where ammonium carbonates followed by ammonium hydrogencarbonate are formed.

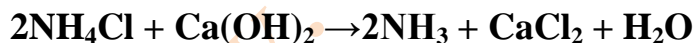
The equations for the complete process may be written as :



Sodium hydrogencarbonate crystal separates. These are heated to give sodium carbonate.



In this process NH<sub>3</sub> is recovered when the solution containing NH<sub>4</sub>Cl is treated with Ca(OH)<sub>2</sub>. Calcium chloride is obtained as a by-product.



It may be mentioned here that Solvay process cannot be extended to the manufacture of potassium carbonate because potassium hydrogencarbonate is too soluble to be precipitated by the addition of ammonium hydrogencarbonate to a saturated solution of potassium chloride.

#### **Properties :**

1. Sodium carbonate is a white crystalline solid which exists as a decahydrate, **Na<sub>2</sub>CO<sub>3</sub>·10H<sub>2</sub>O**.

2. This is also called **washing soda**. It is readily soluble in water. On heating, the decahydrate loses its water of crystallization to form monohydrate. Above 373K, the monohydrate becomes completely anhydrous and changes to a white powder called soda ash.



3. Carbonate part of sodium carbonate gets hydrolysed by water to form an alkaline solution.



#### Uses:

- (i) It is used in water softening, laundering and cleaning.
- (ii) It is used in the manufacture of glass, soap, borax and caustic soda.
- (iii) It is used in paper, paints and textile industries.
- (iv) It is an important laboratory reagent both in qualitative and quantitative analysis.

#### Sodium Chloride, NaCl

The most abundant source of sodium chloride is sea water which contains 2.7 to 2.9% by mass of the salt. In tropical countries like India, common salt is generally obtained by evaporation of sea water.

Approximately 50 lakh tons of salt are produced annually in India by solar evaporation. Crude sodium chloride, generally obtained by crystallization of brine solution, contains sodium sulphate, calcium sulphate, calcium chloride and magnesium chloride as impurities.

Calcium chloride,  $\text{CaCl}_2$ , and magnesium chloride,  $\text{MgCl}_2$  are impurities because they are deliquescent (absorb moisture easily from the atmosphere). To obtain pure sodium chloride, the crude salt is dissolved in minimum amount of water and filtered to remove insoluble impurities. The solution is then saturated with hydrogen chloride gas. Crystals of pure sodium chloride separate out. Calcium and magnesium chloride, being more soluble than sodium chloride, remain in solution.

#### Properties :

1. Sodium chloride melts at 1081K.
2. It has a solubility of 36.0 g in 100 g of water at 273 K.
3. The solubility does not increase appreciably with increase in temperature.

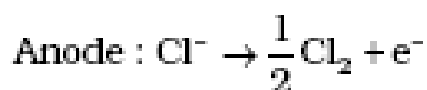
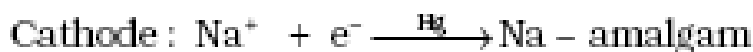


**Uses :**

- (i) It is used as a common salt or table salt for domestic purpose.
- (ii) It is used for the preparation of  $\text{Na}_2\text{O}_2$ ,  $\text{NaOH}$  and  $\text{Na}_2\text{CO}_3$ .

**Sodium Hydroxide (Caustic Soda), NaOH**

Sodium hydroxide is generally prepared commercially by the electrolysis of sodium chloride in Castner-Kellner cell. A brine solution is electrolysed using a mercury cathode and a carbon anode. Sodium metal discharged at the cathode combines with mercury to form sodium amalgam. Chlorine gas is evolved at the anode.



The amalgam is treated with water to give sodium hydroxide and hydrogen gas.

**Properties :**

1. Sodium hydroxide is a white, translucent solid.
2. It melts at 591 K.
3. It is readily soluble in water to give a strong alkaline solution.
4. Crystals of sodium hydroxide are deliquescent.
5. The sodium hydroxide solution at the surface reacts with the  $\text{CO}_2$  in the atmosphere to form  $\text{Na}_2\text{CO}_3$ .

**Uses:**

1. It is used in the manufacture of soap, paper, artificial silk and a number of chemicals,
2. in petroleum refining,
3. in the purification of bauxite,
4. in the textile industries for mercerizing cotton fabrics,
5. for the preparation of pure fats and oils, and
6. as a laboratory reagent.

**Sodium Hydrogencarbonate (BakingSoda), NaHCO<sub>3</sub>**

Sodium hydrogencarbonate is known as baking soda because it decomposes on heating to generate bubbles of carbon dioxide (leaving holes in cakes or pastries and making them light and fluffy).

Sodium hydrogencarbonate is made by saturating a solution of sodium carbonate with carbon dioxide. The white crystalline powder of sodium hydrogencarbonate, being less soluble, gets separated out.



Sodium hydrogencarbonate is a mild antiseptic for skin infections.

It is used in fire extinguishers.

### **BIOLOGICAL IMPORTANCE OF SODIUM AND POTASSIUM**

1. A typical 70 kg man contains about 90 g of Na and 170 g of K compared with only 5 g of iron and 0.06 g of copper.
2. Sodium ions are found primarily on the outside of cells, being located in blood plasma and in the interstitial fluid which surrounds the cells.
3. These ions participate in the transmission of nerve signals, in regulating the flow of water across cell membranes and in the transport of sugars and amino acids into cells.
4. Sodium and potassium, although so similar chemically, differ quantitatively in their ability to penetrate cell membranes, in their transport mechanisms and in their efficiency to activate enzymes.
5. Thus, potassium ions are the most abundant cations within cell fluids, where they activate many enzymes, participate in the oxidation of glucose to produce ATP and, with sodium, are responsible for the transmission of nerve signals.
6. There is a very considerable variation in the concentration of sodium and potassium ions found on the opposite sides of cell membranes.
7. As a typical example, in blood plasma, sodium is present to the extent of  $143 \text{ mmolL}^{-1}$ , whereas the potassium level is only  $5 \text{ mmolL}^{-1}$  within the red blood cells.
8. These concentrations change to  $10 \text{ mmolL}^{-1}$  ( $\text{Na}^+$ ) and  $105 \text{ mmolL}^{-1}$  ( $\text{K}^+$ ). These ionic gradients demonstrate that a discriminatory mechanism, called the sodium-potassium pump, operates across the cell membranes which consumes more than one-third of the ATP used by a resting animal and about 15 kg per 24 hour in a resting human.

(NOTE:-  $\text{mmolL}^{-1} = \text{mili moles per litre}$ )

### **GROUP 2 ELEMENTS : ALKALINE EARTH METALS**

The group 2 elements comprise beryllium, magnesium, calcium, strontium, barium and radium. They follow alkali metals in the periodic table. These (except beryllium) are known as alkaline earth metals. The first element beryllium differs from the rest of the members and shows diagonal relationship to aluminium.

## Electronic Configuration

These elements have two electrons in the s-orbital of the valence shell. Their general electronic configuration may be represented as

**[noble gas] ns<sup>2</sup>.**

Like alkali metals, the compounds of these elements are also predominantly ionic.

Element	Symbol	Atomic number	Electronic configuration
Beryllium	Be	4	[He]2s <sup>2</sup>
Magnesium	Mg	12	[Ne]3s <sup>2</sup>
Calcium	Ca	20	[Ar]4s <sup>2</sup>
Strontium	Sr	38	[Kr]5s <sup>2</sup>
Barium	Ba	56	[Xe]6s <sup>2</sup>
Radium	Ra	88	[Rn]7s <sup>2</sup>

## Atomic and Ionic Radii

The atomic and ionic radii of the alkaline earth metals are smaller than those of the corresponding alkali metals in the same periods. This is due to the increased nuclear charge in these elements.

Within the group, the atomic and ionic radii increase with increase in atomic number.

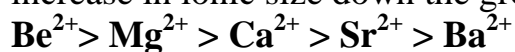
## Ionization Enthalpies

The alkaline earth metals have low ionization enthalpies due to fairly large size of the atoms. Since the atomic size increases down the group, their ionization enthalpy decreases.

The first ionisation enthalpies of the alkaline earth metals are higher than those of the corresponding Group 1 metals. This is due to their small size as compared to the corresponding alkali metals. It is interesting to note that the second ionisation enthalpies of the alkaline earth metals are smaller than those of the corresponding alkali metals.

## Hydration Enthalpies

Like alkali metal ions, the hydration enthalpies of alkaline earth metal ions decrease with increase in ionic size down the group.



The hydration enthalpies of alkaline earth metal ions are larger than those of alkali metal ions. Thus, compounds of alkaline earth metals are more extensively hydrated than those of alkali metals, e.g.,  $\text{MgCl}_2$  and  $\text{CaCl}_2$  exist as  $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$  and  $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$  while  $\text{NaCl}$  and  $\text{KCl}$  do not form such hydrates.

## Physical Properties

- (i) The alkaline earth metals, in general, are silvery white, lustrous and relatively soft but harder than the alkali metals.
- (ii) Beryllium and magnesium appear to be somewhat greyish.
- (iii) The melting and boiling points of these metals are higher than the corresponding alkali metals due to smaller sizes. The trend is, however, not systematic.
- (iv) Because of the low ionization enthalpies, they are strongly electropositive in nature.
- (v) The electropositive character increases down the group from Be to Ba.
- (vi) Calcium, strontium and barium impart characteristic brick red, crimson and apple green colours respectively to the flame.\*

\*-(In flame the electrons are excited to higher energy levels and when they drop back to the ground state, energy is emitted in the form of visible light.)

- (vii) The electrons in beryllium and magnesium are too strongly bound to get excited by flame. Hence, these elements do not impart any colour to the flame.
- (viii) The flame test for Ca, Sr and Ba is helpful in their detection in qualitative analysis and estimation by flame photometry.
- (ix) The alkaline earth metals like those of alkali metals have high electrical and thermal conductivities which are typical characteristics of metals.

## Chemical Properties

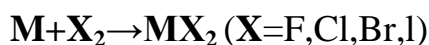
The alkaline earth metals are less reactive than the alkali metals. The reactivity of these elements increases on going down the group.

**(i) Reactivity towards air and water:**

- Beryllium and magnesium are kinetically inert to oxygen and water because of the formation of an oxide film on their surface.
- However, powdered beryllium burns brilliantly on ignition in air to give BeO and Be<sub>3</sub>N<sub>2</sub>.
- Magnesium is more electropositive and burns with dazzling brilliance in air to give MgO and Mg<sub>3</sub>N<sub>2</sub>.
- Calcium, strontium and barium are readily attacked by air to form the oxide and nitride.
- They also react with water with increasing vigour even in cold to form hydroxides.

**(ii) Reactivity towards the halogens:**

All the alkaline earth metals combine with halogen at elevated temperatures forming their halides



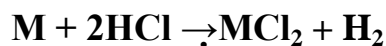
Thermal decomposition of (NH<sub>4</sub>)<sub>2</sub>BeF<sub>4</sub> is the best route for the preparation of BeF<sub>2</sub>, and BeCl<sub>2</sub> is conveniently made from the oxide.

**(iii) Reactivity towards hydrogen:**

All the elements except beryllium combine with hydrogen upon heating to form their hydrides, MH<sub>2</sub>. BeH<sub>2</sub>, however, can be prepared by the reaction of BeCl<sub>2</sub> with LiAlH<sub>4</sub>.

**(iv) Reactivity towards acids:**

The alkaline earth metals readily react with acids liberating dihydrogen.



**Uses:**

- (i) Beryllium is used in the manufacture of alloys.
- (ii) Copper-beryllium alloys are used in the preparation of high strength springs.
- (iii) Metallic beryllium is used for making windows of X-ray tubes.
- (iv) Magnesium-aluminium alloys being light in mass are used in air-craft construction.
- (v) Magnesium (powder and ribbon) is used in flash powders and bulbs, incendiary bombs and signals.
- (vi) A suspension of magnesium hydroxide in water (called milk of magnesia) is used as antacid in medicine.
- (vii) Magnesium carbonate is an ingredient of toothpaste.
- (viii) Calcium is used in the extraction of metals from oxides which are difficult to reduce with carbon.
- (ix) Calcium and barium metals, owing to their reactivity with oxygen and nitrogen at elevated temperatures, have often been used to remove air from vacuum tubes.
- (x) Radium salts are used in radiotherapy, for example, in the treatment of cancer.

### ANOMALOUS BEHAVIOUR OF BERYLLIUM

Beryllium, the first member of the Group 2 metals, shows anomalous behaviour as compared to magnesium and rest of the members. Further, it shows diagonal relationship to aluminium

- (i) Beryllium has exceptionally small atomic and ionic sizes and thus does not compare well with other members of the group. Because of high ionisation enthalpy and small size it forms compounds which are largely covalent and get easily hydrolysed.
- (ii) Beryllium does not exhibit coordination number more than four as in its valence shell there are only four orbitals. The remaining members of the group can have a coordination number of six by making use of d-orbitals.
- (iii) The oxide and hydroxide of beryllium, unlike the hydroxides of other elements in the group, are amphoteric in nature.

### Diagonal Relationship between Beryllium and Aluminium

The ionic radius of  $\text{Be}^{2+}$  is estimated to be 31 pm; the charge/radius ratio is nearly the same as that of the  $\text{Al}^{3+}$  ion. Hence beryllium resembles aluminium in some ways.

### Similarities between Beryllium and Aluminium

- (i) Like aluminium, beryllium is not readily attacked by acids because of the presence of an oxide film on the surface of the metal.
- (ii) Beryllium hydroxide dissolves in excess of alkali to give a beryllate ion,  $[\text{Be}(\text{OH})_4]^{2-}$  just as aluminium hydroxide gives aluminate ion,  $[\text{Al}(\text{OH})_4]^-$ .
- (iii) The chlorides of both beryllium and aluminium have  $\text{Cl}^-$  bridged chloride structure in vapour phase. Both the chlorides are soluble in organic solvents and are strong Lewis acids. They are used as Friedel Craft catalysts.
- (iv) Beryllium and aluminium ions have strong tendency to form complexes,  $\text{BeF}_4^{2-}$ ,  $\text{AlF}_6^{3-}$ .

### SOME IMPORTANT COMPOUNDS OF CALCIUM

Important compounds of calcium are calcium oxide, calcium hydroxide, calcium sulphate, calcium carbonate and cement. These are industrially important compounds..

### Calcium Oxide or Quick Lime, CaO

#### Preparation

It is prepared on a commercial scale by heating limestone ( $\text{CaCO}_3$ ) in a rotary kiln at 1070-1270 K.



The carbon dioxide is removed as soon as it is produced to enable the reaction to proceed to completion.

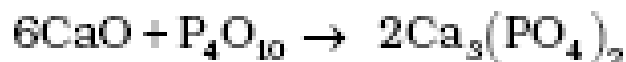
#### Properties

1. Calcium oxide is a white amorphous solid.
2. It has a melting point of 2870 K.
3. On exposure to atmosphere, it absorbs moisture and carbon dioxide.



The addition of limited amount of water breaks the lump of lime.

This process is called **slaking of lime**. Quick lime slaked with soda gives solid **sodalime**. Being a basic oxide, it combines with acidic oxides at high temperature.



#### Uses:

- (i) It is an important primary material for manufacturing cement and is the cheapest form of alkali.
- (ii) It is used in the manufacture of sodium carbonate from caustic soda.
- (iii) It is employed in the purification of sugar and in the manufacture of dye stuffs.

#### Calcium Hydroxide (Slaked lime), $\text{Ca}(\text{OH})_2$

Calcium hydroxide is prepared by adding water to quick lime,  $\text{CaO}$ .

It is a white amorphous powder. It is sparingly soluble in water. The aqueous solution is known as lime water and a suspension of slaked lime in water is known as milk of lime.

When carbon dioxide is passed through lime water it turns milky due to the formation of calcium carbonate.



On passing excess of carbon dioxide, the precipitate dissolves to form calcium hydrogencarbonate.



Milk of lime reacts with chlorine to form hypochlorite, a constituent of bleaching powder.





**Uses:**

- (i) It is used in the preparation of mortar, a building material.
- (ii) It is used in white wash due to its disinfectant nature.
- (iii) It is used in glass making, in tanning industry, for the preparation of bleaching powder and for purification of sugar.

**Calcium Carbonate, CaCO<sub>3</sub>**

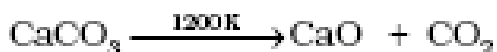
Calcium carbonate occurs in nature in several forms like limestone, chalk, marble etc. It can be prepared by passing carbon dioxide through slaked lime or by the addition of sodium carbonate to calcium chloride.



Excess of carbon dioxide should be avoided since this leads to the formation of water soluble calcium hydrogencarbonate.

Calcium carbonate is a white fluffy powder.

It is almost insoluble in water. When heated to 1200 K, it decomposes to evolve carbon dioxide.



It reacts with dilute acid to liberate carbon dioxide.

**Uses:**

- (i) It is used as a building material in the form of marble and in the manufacture of quick lime.
- (ii) Calcium carbonate along with magnesium carbonate is used as a flux in the extraction of metals such as iron.
- (iii) Specially precipitated  $\text{CaCO}_3$  is extensively used in the manufacture of high quality paper.
- (iv) It is also used as an antacid, mild abrasive in tooth paste, a constituent of chewing gum, and a filler in cosmetics.

**BIOLOGICAL IMPORTANCE OF MAGNESIUM AND CALCIUM**

- (i) An adult body contains about 25 g of Mg and 1200 g of Ca compared with only 5 g of iron and 0.06 g of copper. The daily requirement in the human body has been estimated to be 200 – 300 mg.
- (ii) All enzymes that utilise ATP in phosphate transfer require magnesium as the cofactor.
- (iii) The main pigment for the absorption of light in plants is chlorophyll which contains magnesium.
- (iv) About 99 % of body calcium is present in bones and teeth.
- (v) It also plays important roles in neuromuscular function, interneuronal transmission, cell membrane integrity and blood coagulation.
- (vi) The calcium concentration in plasma is regulated at about  $100 \text{ mgL}^{-1}$ . It is maintained by two hormones: calcitonin and parathyroid hormone.
- (vii) Bone is not an inert and unchanging substance but is continuously being solubilised and redeposited to the extent of 400 mg per day in man? All this calcium passes through the plasma.

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