

## Chapter 10

### The s- Block Element

#### Important points

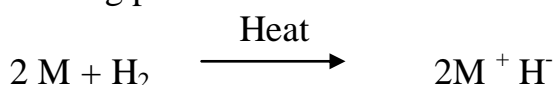
- Groups (1 & 2) belong to the s-block of the Periodic Table.
- Group 1 consists of : lithium, sodium, potassium, rubidium, caesium and francium and collectively known as the *alkali metals*.
- Group 2 include : beryllium, magnesium, calcium, strontium, barium and radium. Except Beryllium they are known as *alkaline*

#### ❖ Physical properties-

- a) **Large atomic radii:** The atomic radii of alkali metals are the largest in their respective periods. These increase as we travel down the group.
- b) **Large ionic radii:** The ionic radii increase as we move down the group due to the addition of a new energy shell with each succeeding element.
- c) **Low ionization enthalpy:** The ionization enthalpies decrease as we move down the group. The ionization enthalpies of the alkali metals are the lowest due to loosely held s- electron.
- d) **Hydration enthalpy:** It decreases with the increase in ionic radii. The hydration enthalpy of Li ion is the maximum and the hydration enthalpy of Cs ion is the minimum.
- e) **Oxidation state:** The alkali metals exhibit oxidation state of +1 in their compounds and are strongly electropositive in character. The electropositive character increases from Li to Cs.
- f) **Metallic character:** The metallic character increases down the group.
- g) **Melting point and boiling point::** The m p and b p of alkali metals are very low and decrease with increase in atomic number.
- h) **Nature of bonds formed:** These metals form ionic bonds. The ionic character increases as we down the group.
- i) **Flame colouration:** All the alkali metals impart a charactersistic colour to the flame.
- j) **Photoelectric effect:** Alkali metals (except Li) exhibits photoelectric effect.

#### ❖ Chemical features of alkali metals:

- a) **Reducing character:** As the ionization enthalpies of the alkali metals decrease down the group their reducing character or reactivity in the gaseous state increases down the group. i.e.,  $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$ .
- b) **Reaction with dihydrogen:** Alkali metals react with dry hydrogen at about 673 K to form crystalline hydrides which are ionic in nature and have high melting points.



- c) **Oxides and hydroxides:** Alkali metals when burnt in air form different compounds, for example the alkali metals on reaction with limited quantity of oxygen form normal oxides ( $\text{M}_2\text{O}$ )  $\text{M} = \text{Li}, \text{Na}, \text{K}, \text{Rb}, \text{Cs}$

- d) **Reaction with halogens:** The members of the family combine with halogen to form corresponding halides which are ionic crystalline solids.  
Reactivity of alkali metals with particular halogen increases from Li to Cs.
- e) **Reaction with water:** Alkali metals react with water and other compounds containing acidic hydrogen atoms such as hydrogen halides, acetylene etc. to liberate hydrogen gas.
- f) **Solubility in liquid ammonia:** All alkali metals dissolve in liquid ammonia giving deep blue solutions which are conducting in nature.
- g) **Reaction with sulphur and phosphorus:** Alkali metals react with sulphur and phosphorus on heating to form sulphides and phosphides respectively.

#### ❖ Diagonal relationship between Li and Al

Li resembles Mg mainly due to similarity in sizes of their atoms and ions. The main points of similarity are:

- i) Both are quite hard.
- ii) Both LiOH and  $\text{Mg}(\text{OH})_2$  are weak bases.
- iii) Carbonates of both on heating decompose to produce oxides and carbon dioxide.
- iv) Both react with nitrogen to give ionic nitrides.
- v) Nitrates of both decompose on heating to give oxides.
- vi) Both Li and Mg do not form solid bicarbonates.
- vii) Because of covalent character LiCl and  $\text{MgCl}_2$  are soluble in ethanol.
- viii) The hydroxides, bicarbonates and fluorides of both Li and Mg are sparingly soluble in water.

#### ❖ Biological importance of Na and K

- i) Sodium ions participate in the transmission of nerve signals.
- ii) Sodium ions also regulate flow of water across the cell membranes and in transport of sugars and amino acids into the cells.
- iii) Potassium ions are the most abundant cations within cell fluids, where they activate many enzymes, participate in oxidation of glucose to produce ATP.
- iv) Potassium ions in combination with sodium ions are responsible for transmission of nerve signals.
- v) The functional features of nerve cells depend upon the sodium potassium ion gradient that is established in the cell.

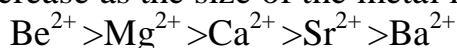
#### ❖ Group 2 elements: Alkaline earth metals

(a) **Atomic radii :** The atomic radii of alkaline earth metals are fairly large though smaller than the corresponding alkali metals and they increase down the group. This is because on moving down the group, atomic radii increase primarily due to the addition of an extra shell of electrons in each succeeding element.

(b) **Ionic radii:** the atoms of these elements form divalent ions which show the same trend of increase in their size down the group.

(c) **Ionization enthalpy:** The alkaline earth metals have fairly low Ionization enthalpies though greater than those of the corresponding elements of group 1 and this value decreases down the group.

**(d) Hydration enthalpy:** the Hydration enthalpies of alkaline earth metal ion decrease as the size of the metal ion increases down the Group



**(e) Oxidation State:** All the members of the family exhibit +2 oxidation state in their compounds and the form divalent cations ( $\text{M}^{2+}$ )

**(f) Electro negativity :** The electro negativity values of alkaline earth metals are quite close to those of alkali metals, though slightly more.

**(g) Metallic Character :** Alkaline earth metals have stronger metallic bonds as compared to the alkali metals present in the same period.

**(h) Melting and boiling point :** The melting and Boiling points of these metals are higher than those of alkali metals present in the same period.

**(i) Colouration to the flame :** With the exception of beryllium and magnesium, the rest of the elements impart characters in colour to the same flame. For example,

Be	Mg	Ca	Sr	Ba	Ra
-	-	Brick Red	Crimson	Grassy Green	Crimson

**J) Complex formation:** Generally the members do not form complexes. However, smaller ions ( Be & Mg Ions) form complexes with the electron donor species

**k) Formation of organo-metallic compounds:** Both beryllium and magnesium form a number of organo-metallic compounds containing M-C bond with certain organic compounds. For example, magnesium reacts with alkyl halide in the presence of dry ether to give Grignard reagent.

**l) Reducing character:** Alkaline earth metals are weak reducing agent than the corresponding alkali metals which have lower ionization enthalpies and comparatively bigger atomic sizes.

**m) Reaction with oxygen:** With the exception of Ba and Ra which form peroxides ( $\text{MO}_2$ ) rest of the metals form normal oxides ( $\text{MO}$ ) on heating with excess of oxygen.

**n) Reaction with halogens:** The members of the family combine directly with halogen at appropriate temperature to form corresponding halides.

**o) Reaction with water:** The members of this group are less reactive towards water as compared to the corresponding alkali metals because these are less electropositive in nature.

**p) Reaction with hydrogen:** The members except Be combine with hydrogen directly upon heating to form metal hydrides.

**Uses of some important compounds:-**

**(i) Caustic soda:**

It is used: in soap, paper, textile, petroleum industry

**ii) Sodium carbonate**

It is used:

- in glass and soap industry
- in paper making and textile manufacturing
- in paint and dye stuffs
- in metal refining

- e) in production of sodium compounds such as borax, caustic soda, sodium phosphate etc.

### iii) Quick lime:

It is used:

- in the preparation of cement, glass and calcium carbide.
- In the purification of sugar
- In softening of hard water
- As a flux in the extraction of metal

### iv) Lime stone: It is used

- as building material
- in the manufacture of quick lime
- in Solvay process to prepare  $\text{Na}_2\text{CO}_3$  as it is a source of  $\text{CO}_2$
- in metallurgy for the extraction of iron
- in toothpaste and certain cosmetics

### v) Cement: It is an important building material. It is used in concrete and reinforced concrete, in plastering and in the construction of bridges, dams and buildings.

### vi) Plaster of paris: It is used

- in making moulds for pottery and ceramics etc.
- in surgical bandages for setting broken bones of the body
- for making statues, models, decorative materials and black board chalk.

### ❖ Biological importance of Ca and Mg

- Magnesium ions are concentrated in animal cells and Calcium ions are concentrated in body fluids, outside the cell.
- All enzymes that utilize ATP in phosphate transfer require magnesium ion as cofactor.
- In green plants magnesium is present in chlorophyll.
- Calcium and magnesium ions are also essential for the transmission of impulses along nerve fibres.
- Calcium ions are important in blood clotting and are required to trigger the contraction of muscles.
- Calcium ions also regulate the beating of the heart.

### One mark questions:

1. Why are halides of beryllium polymeric?

Ans:- the halides of Be are electron deficient as their octets are incomplete. Therefore, to complete their octets, the halides polymerize.

2. Name the groups which constitute s-block elements.

Ans:- group-1 and 2

3. Arrange the alkaline earth metal carbonates in the decreasing order of thermal stability.

Ans:-  $\text{BaCO}_3 > \text{SrCO}_3 > \text{CaCO}_3 > \text{MgCO}_3 > \text{BeCO}_3$

4. Write the general electronic configuration of s-block elements.

Ans:- [Noble gas]  $ns^{1-2}$

5. What is the chemical formula of Plaster of Paris?

Ans:-  $\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$

6. Name the compound which can be obtained by Solvay's process.

Ans:- Sodium carbonate

7. How does the basic character of hydroxides of alkali metals vary down the group?

Ans:- Increases down the group

8. Which out of  $\text{MgSO}_4$  or  $\text{BaSO}_4$  is more soluble in water?

Ans:-  $\text{MgSO}_4$

9. Name radioactive elements of group 1 and 2.

Ans:- Francium and Radium.

10. Which elements of alkaline earth metals family do not give characteristic flame colouration?

Ans:- Be and Mg

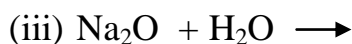
### Two marks questions

1. Among the alkali metals which has

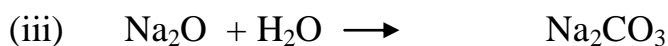
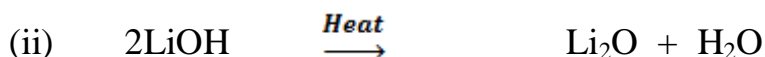
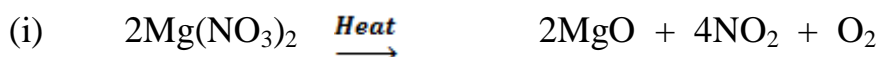
- (i) Highest melting point
- (ii) Most electropositive character
- (iii) Lowest size of ion
- (iv) Strongest reducing character.

Ans:- (i) Li (ii) Cs (iii) Li (iv) Li

2. Complete the following reactions:



Ans:-



3. Name the chief factors responsible for anomalous behaviour of lithium.

Ans:- the anomalous behaviour of lithium is because of its:

- (i) Small size of atom and ion,
- (ii) High ionization enthalpy, and
- (iii) Absence of d-orbitals in its Valence shell.

4. Which out of Li and Na has greater value for the following properties:

- (i) Hydration enthalpy
- (ii) Stability of hydride
- (iii) Stability of carbonate
- (iv) Basic character of hydroxide

Ans:- (i) Li (ii) Li (iii) Na (iv) Na

### 5. Why are alkali metals not found in nature?

**Ans.** Alkali metals are highly reactive in nature due to low ionization enthalpy and strong electropositive character. They do not occur in free state and are always combined with other elements. As a result alkali metals are not generally found in nature.

### 6. Why are lithium salts commonly hydrated and those of the other alkali ions usually anhydrous?

**Ans.** In the lithium salt, the  $\text{Li}^+$  ion due to very small size gets readily hydrated on coming in contact with moisture (water). Therefore, lithium salts are commonly hydrated. But the other alkali metal ions are comparatively big in size. They have therefore, lesser tendency to get hydrated. These salts are usually anhydrous.

### 7. Beryllium and magnesium do not give colour to flame whereas other alkaline earth metals do so why?

**Ans:** Beryllium and magnesium atoms in comparison to other alkaline earth metals are comparatively smaller and their ionisation enthalpies are very high. Hence, the energy of the flame is not sufficient to excite their electrons to higher energy levels. These elements, therefore, do not give any colour in Bunsen flame.

### 7. Why are alkali metals soft and have low melting points?

**Ans:** Alkali metals have only one valence electron per metal atom. As a result, the binding energy of alkali metal ions in the close-packed metal lattices are weak. Therefore, these are soft and have low melting point.

### 8. Which out of the following and why can be used to store an alkali metal? $\text{H}_2\text{O}$ , $\text{C}_2\text{H}_5\text{OH}$ and Benzene

**Ans:-** Benzene can be used to store an alkali metal because other substance react with alkali metal as:



### 9. Why are alkali metals not found free in nature?

**Ans:-** alkali metals are highly reactive and therefore, are not found free in nature, they are present in the combined state in the form of halides, oxides, silicates, nitrates, etc.

### Three marks questions

#### 1. When an alkali metal dissolves in liquid ammonia the solution can acquire different colours. Explain the reasons for this type of colour change.

**Ans.** The dissolution of the metal in liquid ammonia is accompanied by their formation of ammoniated electrons that give rise to dark colour. This is because ammoniated electrons absorb energy corresponding to the red region of the visible light. However, if the concentration increases above 3M, the colour changes to copper-bronze and the solution acquires metallic luster due to the formation of metal ion clusters.





## 2. In what ways lithium shows similarities to magnesium in its chemical behaviour?

**Ans.** Li resembles Mg mainly due to similarity in sizes of their atoms and ions. The main points of similarity are:

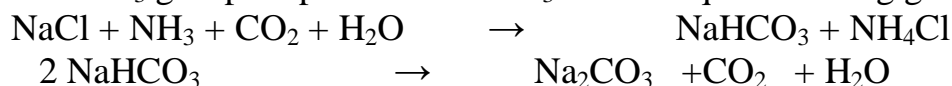
Both are quite hard.

- 1 Both LiOH and Mg(OH)<sub>2</sub> are weak bases.
- 2 Carbonates of both on heating decompose to produce oxides and carbondioxide.
- 3 Both react with nitrogen to give ionic nitrides.

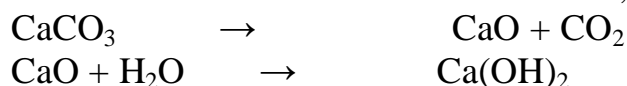
## 3. Discuss the various reactions that occur in the Solvay process.

**Ans.** In Solvay ammonia process.

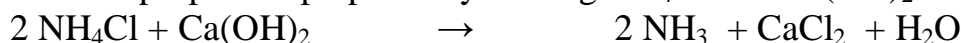
When carbon dioxide is passed through a concentrated solution of brine saturated with NH<sub>3</sub>, NaHCO<sub>3</sub> gets precipitated. NaHCO<sub>3</sub> on subsequent heating gives Na<sub>2</sub>CO<sub>3</sub>.



CO<sub>2</sub> needed for the reaction is prepared by heating calcium carbonate and the quick lime, CaO is dissolved in water to form slaked lime, Ca(OH)<sub>2</sub>

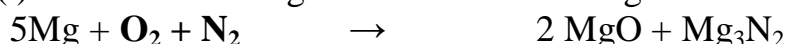


NH<sub>3</sub> needed for the purpose is prepared by heating NH<sub>4</sub>Cl and Ca(OH)<sub>2</sub>



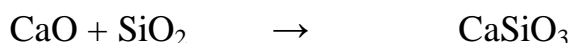
## 4. What happen when (i) magnesium is burnt in air (ii) quick lime is heated with silica (iii) chlorine reacts with slaked lime (iv) calcium nitrate is heated?

**Ans.** (i) A mixture of magnesium oxide and magnesium nitride is formed

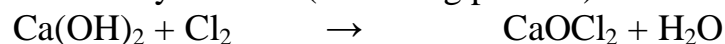


**In air**

(ii) Calcium silicate is formed.



(iii) Calcium oxychloride (bleaching powder) is formed



(iv) Nitrogen dioxide is evolved.



## 5. Describe the importance of the following (i) limestone (ii) cement (iii) plaster of paris.

**Ans. i) Lime stone:** It is used

- f) as building material
- g) in the manufacture of quick lime
- h) in Solvay process to prepare Na<sub>2</sub>CO<sub>3</sub> as it is a source of CO<sub>2</sub>
- i) in metallurgy for the extraction of iron
- j) in toothpaste and certain cosmetics

**ii) Cement:** It is an important building material. It is used in concrete and reinforced concrete, in plastering and in the construction of bridges, dams and buildings.

iii) **Plaster of paris:** It is used

- d) in making moulds for pottery and ceramics etc.
- e) in surgical bandages for setting broken bones of the body
- f) for making statues, models, decorative materials and black board chalk.

**6. What happens when:**

a) **Sodium metal is dropped in water?**

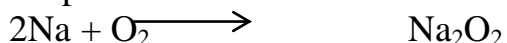
b) **Sodium metal is heated in free supply of air?**

c) **Sodium peroxide dissolves in water?**

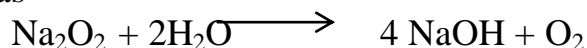
**Ans.** a) Sodium metal catches fire and hydrogen gas is evolved



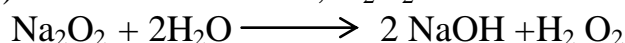
b) Sodium peroxide is formed



c) (i) Sodium peroxide reacts with water at ordinary temperature to liberate oxygen gas



ii) With ice cold water,  $\text{H}_2\text{O}_2$  is formed



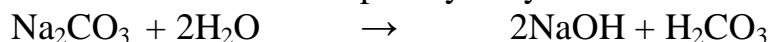
7. **State as to why**

a) **a solution of  $\text{Na}_2\text{CO}_3$  is alkaline?**

b) **alkali metals are prepared by electrolysis of their fused chlorides?**

c) **sodium is found to be more useful than potassium?**

**Ans.** (a) Sodium carbonate being a salt of strong base ( $\text{NaOH}$ ) and weak acid ( $\text{H}_2\text{CO}_3$ ) forms alkaline solution upon hydrolysis



(b) Since the discharge potential of alkali metals is much higher than that of hydrogen, therefore, when the aqueous solution of any alkali metal chloride is subjected to electrolysis,  $\text{H}_2$  instead of the alkali metal is produced at the cathode. Therefore, to prepare alkali metals, electrolysis of their fused chlorides is carried out.

(c) Sodium is relatively more abundant than potassium. At the same time, it is also less reactive and its reactions with other substances can be better controlled.

**8. Why are potassium and cesium, rather than lithium used in photoelectric cells?**

**Ans.** The ionization enthalpy of lithium is quite high. The photons of light are not in a position to eject electrons from the surface of lithium metal. Therefore photoelectric effect is not noticed. However, both potassium and cesium have comparatively low ionization enthalpies. This means that the electrons can quite easily be ejected from the surface of these metals when photons of certain minimum frequency (threshold frequency) strike against their surface

**9. Why is  $\text{Li}_2\text{CO}_3$  decomposed at a lower temperature whereas  $\text{Na}_2\text{CO}_3$  at higher temperature?**

**Ans.**  $\text{Li}^+$  ion is very small in size. It is stabilized more by smaller anions such as oxide ion rather than large anions such as carbonate. Therefore  $\text{Li}_2\text{CO}_3$  decomposes into  $\text{Li}_2\text{O}$  on mild heating. On the other hand,  $\text{Na}^+$  ion is larger in size. It is stabilized



more by carbonate ion than oxide ion. Hence,  $\text{Na}_2\text{CO}_3$  does not undergo thermal decomposition easily.

### 10. Explain why can alkali and alkaline earth metals not be obtained by chemical reduction methods?

**Ans.** The metals belonging to both these families are very strong reducing agents. It is therefore not possible to reduce their oxides by reacting with common reducing agents like carbon (coke), zinc etc. These are normally isolated by carrying out the electrolysis of the salts of these metals in the molten state.

### Five marks questions:

1. Compare the solubility and thermal stability of the following compounds of the alkali metals with those of the alkaline earth metals. (a) Nitrates (b) Carbonates (c) Sulphates.

**Ans. Solubility:**

In case of alkali metals: Nitrates, carbonates and sulphates of alkali metals are soluble in water. In alkali metals lattice energies decrease more rapidly than the hydration energies, therefore their solubility increases down the group.

In case of alkaline earth metals: Nitrates of all alkaline earth metals are soluble in water but their solubility decreases down the group because their hydration energies decrease more rapidly than their lattice energies.

Since the size of  $\text{CO}_3^{2-}$  and  $\text{SO}_4^{2-}$  anions is much larger than the cations, therefore lattice energies remain almost constant within a particular group. Since, the hydration energies decrease as we move down the group, therefore the solubility of alkaline earth metal carbonates and sulphates decrease down the group. However, the hydration energy of  $\text{Be}^{2+}$  and  $\text{Mg}^{2+}$  ions overcome the lattice energy factor and therefore  $\text{BeSO}_4$  and  $\text{MgSO}_4$  are readily soluble in water while the solubility of other sulphates decreases down the group from  $\text{CaSO}_4$  to  $\text{BaSO}_4$ .

**Thermal Stability:**

- a) Nitrates: Nitrates of both alkali and alkaline earth metals decompose on heating.

All alkaline earth metal nitrates decompose to form metal oxide,  $\text{NO}_2$  and  $\text{O}_2$ .

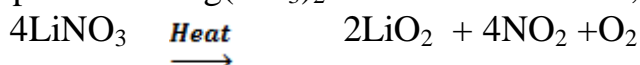


M = Be, Mg, Ca, Sr, or Ba

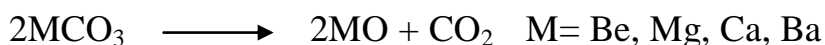
The nitrates of Na, K, Rb and Cs decompose to form metal nitrites and  $\text{O}_2$ .



However, due to diagonal relationship between Li and Mg, lithium nitrate decomposes like  $\text{Mg}(\text{NO}_3)_2$  to form metal oxide,  $\text{NO}_2$  and  $\text{O}_2$ .

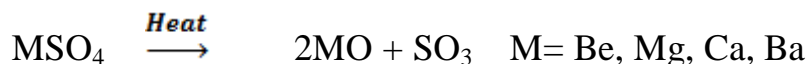


- b) Carbonates: Carbonates of alkaline earth metals decompose on heating to form metal oxide and carbon dioxide.



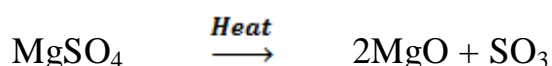
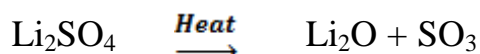
Further as the electropositive character of the metal increases down the group the stability of these metal carbonates increases or the temperature of their decomposition increases.

c) Sulphates: Sulphates of alkaline earth metals decompose on heating to form metal oxide and  $\text{SO}_3$ .



The temperature of decomposition of these sulphates increases as the electropositive character of the metal or the basicity of the metal hydroxide increases down the group.

Among the alkali metals due to diagonal relationship,  $\text{Li}_2\text{SO}_4$  decomposes like  $\text{MgSO}_4$  to form the corresponding metal oxide and  $\text{SO}_3$ .



Other alkali metals are stable to heat and do not decompose easily.

2. Compare the alkali metals and alkaline earth metals with respect to (i) ionization enthalpy (ii) basicity of oxides and (iii) solubility of hydroxides..

Ans.

- (i) Ionization enthalpy (I E): I E of alkaline earth metals are higher than those of alkali metals of group 1. This is because the atoms of alkaline earth metals have smaller size (due to higher nuclear charge) as compared to the alkali metals.
- (ii) Basicity of oxides: The oxides of alkali and alkaline earth metals dissolve in water to form their respective hydroxides. These hydroxides are strong bases. The hydroxides of alkaline earth metals are less basic than of alkali metals of the corresponding periods. This is due to their (i) high ionization enthalpy (ii) small ionic size and (iii) dipositive charge on the ions.

As a result M-O bond in these hydroxides is relatively stronger than that of corresponding alkali metals and therefore does not break.

- (iii) Solubility of hydroxides: Because of smaller size and higher ionic charge, the lattice enthalpies of alkaline earth metals are much higher than those of alkali metals and hence the solubility of alkali metal hydroxides is much higher than that of alkaline earth metals. However the solubility of the hydroxides of both alkali and alkaline earth metals increase down the group due to large decrease in their lattice enthalpies as compared to their hydration enthalpies.

3. Explain the significance of sodium, potassium, magnesium and calcium in biological fluids.

Ans. Significance of sodium and potassium:

- (i) Sodium ions participate in the transmission of nerve signals.
- (ii) Sodium ions also regulate flow of water across the cell membranes and in transport of sugars and amino acids into the cells.
- (iii) Potassium ions are the most abundant cations within cell fluids, where they activate many enzymes, participate in oxidation of glucose to produce ATP.

- (iv) Potassium ions in combination with sodium ions are responsible for transmission of nerve signals.
- (v) The functional features of nerve cells depend upon the sodium potassium ion gradient that is established in the cell.

Significance of Magnesium and Calcium:

1. Magnesium ions are concentrated in animal cells and Calcium ions are concentrated in body fluids, outside the cell.
2. All enzymes that utilize ATP in phosphate transfer require magnesium ion as cofactor.
3. In green plants magnesium is present in chlorophyll.
4. Calcium and magnesium ions are also essential for the transmission of impulses along nerve fibres.
5. Calcium ions are important in blood clotting and are required to trigger the contraction of muscles.
6. Calcium ions also regulate the beating of the heart.

### HOTS QUESTIONS

1. Potassium carbonate cannot be prepared by Solvay process. Why?

Ans. This is due to the reason that potassium bicarbonate ( $\text{KHCO}_3$ ) formed as an intermediate (when  $\text{CO}_2$  gas is passed through ammoniated solution of potassium chloride) is highly soluble in water and cannot be separated by filtration.

2. The hydroxides and carbonates of sodium and potassium are easily soluble in water while the corresponding salts of magnesium and calcium are sparingly soluble in water. Explain.

Ans. All the compounds are crystalline solids and their solubility in water is guided by both lattice enthalpy and hydration enthalpy. In case of sodium and potassium compounds, the magnitude of lattice enthalpy is quite small as compared to hydration enthalpy since the cationic sizes are large. Therefore, the compounds of sodium and potassium that are mentioned, readily dissolve in water. However, in case of corresponding magnesium and calcium compounds, the cations have smaller sizes and more magnitude of positive charge. This means that their lattice enthalpies are more as compared to the compounds of sodium and potassium. Therefore, the hydroxides and carbonates of these metals are only sparingly soluble in water.

3. Why is  $\text{LiF}$  almost insoluble in water whereas  $\text{LiCl}$  soluble not only in water but also in acetone?

Ans. The low solubility of  $\text{LiF}$  in water is due to its very high lattice enthalpy ( $\text{F}^-$  ion is very small in size). On the other hand, in lithium chloride ( $\text{LiCl}$ ) the lattice enthalpy is comparatively very small. This means that the magnitude of hydration enthalpy is quite large. Therefore lithium chloride dissolves in water. It is also soluble in acetone due to dipolar attraction. (Acetone is polar in nature)