

### Question 10.1:

What are the common physical and chemical features of alkali metals?

Answer

Physical properties of alkali metals are as follows.

(1) They are quite soft and can be cut easily. Sodium metal can be easily cut using a knife.

(2) They are light coloured and are mostly silvery white in appearance.

(3) They have low density because of the large atomic sizes. The density increases down the group from Li to Cs. The only exception to this is K, which has lower density than Na.

(4) The metallic bonding present in alkali metals is quite weak. Therefore, they have low melting and boiling points.

(5) Alkali metals and their salts impart a characteristic colour to flames. This is because the heat from the flame excites the electron present in the outermost orbital to a high energy level. When this excited electron reverts back to the ground state, it emits excess energy as radiation that falls in the visible region.

(6) They also display photoelectric effect. When metals such as Cs and K are irradiated with light, they lose electrons.

### Chemical properties of alkali metals

Alkali metals are highly reactive due to their low ionization enthalpy. As we move down the group, the reactivity increases.

(1) They react with water to form respective oxides or hydroxides. As we move down the group, the reaction becomes more and more spontaneous.

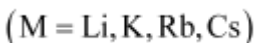
(2) They react with water to form their respective hydroxides and dihydrogens. The general reaction for the same is given as



(3) They react with dihydrogen to form metal hydrides. These hydrides are ionic solids and have high melting points.



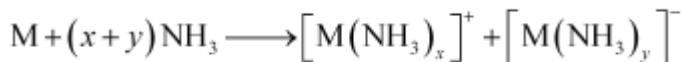
(4) Almost all alkali metals, except Li, react directly with halogens to form ionic halides.



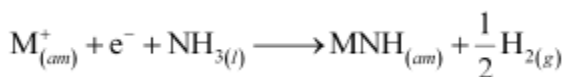
Since  $\text{Li}^+$  ion is very small in size, it can easily distort the electron cloud around the negative halide ion. Therefore, lithium halides are covalent in nature.

**(5)** They are strong reducing agents. The reducing power of alkali metals increases on moving down the group. However, lithium is an exception. It is the strongest reducing agent among the alkali metals. It is because of its high hydration energy.

**(6)** They dissolve in liquid ammonia to form deep blue coloured solutions. These solutions are conducting in nature.



The ammoniated electrons cause the blue colour of the solution. These solutions are paramagnetic and if allowed to stand for some time, then they liberate hydrogen. This results in the formation of amides.



In a highly concentrated solution, the blue colour changes to bronze and the solution becomes diamagnetic.

### Question 10.2:

Discuss the general characteristics and gradation in properties of alkaline earth metals.

Answer

General characteristics of alkaline earth metals are as follows.

**(i)** The general electronic configuration of alkaline earth metals is [noble gas]  $ns^2$ .

**(ii)** These metals lose two electrons to acquire the nearest noble gas configuration. Therefore, their oxidation state is +2.

**(iii)** These metals have atomic and ionic radii smaller than that of alkali metals. Also, when moved down the group, the effective nuclear charge decreases and this causes an increase in their atomic radii and ionic radii.

**(iv)** Since the alkaline earth metals have large size, their ionization enthalpies are found to be fairly low. However, their first ionization enthalpies are higher than the corresponding group 1 metals.

**(v)** These metals are lustrous and silvery white in appearance. They are relatively less soft as compared to alkali metals.

**(vi)** Atoms of alkaline earth metals are smaller than that of alkali metals. Also, they have two valence electrons forming stronger metallic bonds. These two factors cause alkaline earth metals to have high melting and boiling points as compared to alkali metals.

**(vii)** They are highly electropositive in nature. This is due to their low ionization enthalpies. Also, the electropositive character increases on moving down the group from Be to Ba.

**(viii)** Ca, Sr, and Ba impart characteristic colours to flames.

Ca – Brick red

Sr – Crimson red

Ba – Apple green

In Be and Mg, the electrons are too strongly bound to be excited. Hence, these do not impart any colour to the flame.

The alkaline earth metals are less reactive than alkali metals and their reactivity increases on moving down the group. Chemical properties of alkaline earth metals are as follows.

**(i)** Reaction with air and water: Be and Mg are almost inert to air and water because of the formation of oxide layer on their surface.

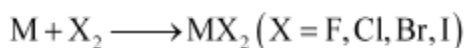
**(a)** Powdered Be burns in air to form BeO and Be<sub>3</sub>N<sub>2</sub>.

**(b)** Mg, being more electropositive, burns in air with a dazzling sparkle to form MgO and Mg<sub>3</sub>N<sub>2</sub>.

**(c)** Ca, Sr, and Ba react readily with air to form respective oxides and nitrides.

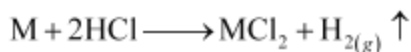
**(d)** Ca, Ba, and Sr react vigorously even with cold water.

**(ii)** Alkaline earth metals react with halogens at high temperatures to form halides.



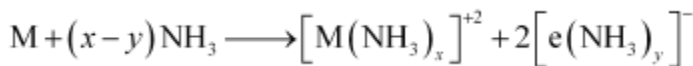
**(iii)** All the alkaline earth metals, except Be, react with hydrogen to form hydrides.

**(iv)** They react readily with acids to form salts and liberate hydrogen gas.



**(v)** They are strong reducing agents. However, their reducing power is less than that of alkali metals. As we move down the group, the reducing power increases.

**(vi)** Similar to alkali metals, the alkaline earth metals also dissolve in liquid ammonia to give deep blue coloured solutions.

**Question 10.3:**

Why are alkali metals not found in nature?

Answer

Alkali metals include lithium, sodium, potassium, rubidium, cesium, and francium. These metals have only one electron in their valence shell, which they lose easily, owing to their low ionization energies. Therefore, alkali metals are highly reactive and are not found in nature in their elemental state.

**Question 10.4:**

Find the oxidation state of sodium in  $Na_2O_2$ .

Answer

Let the oxidation state of Na be  $x$ . The oxidation state of oxygen, in case of peroxides, is  $-1$ .

Therefore,

$$2(x) + 2(-1) = 0$$

$$2x - 2 = 0$$

$$2x = 2$$

$$x = +1$$

Therefore, the oxidation state of sodium is  $+1$ .

**Question 10.5:**

Explain why is sodium less reactive than potassium?

Answer

In alkali metals, on moving down the group, the atomic size increases and the effective nuclear charge decreases. Because of these factors, the outermost electron in potassium can be lost easily as compared to sodium. Hence, potassium is more reactive than sodium.

**Question 10.6:**

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

	<b>Alkali metals</b>		<b>Alkaline earth metals</b>
<b>(i)</b>	<p>Ionization enthalpy:</p> <p>These have lowest ionization enthalpies in respective periods. This is because of their large atomic sizes. Also, they lose their only valence electron easily as they attain stable noble gas configuration after losing it.</p>	<b>(i)</b>	<p>Ionization enthalpy:</p> <p>Alkaline earth metals have smaller atomic size and higher effective nuclear charge as compared to alkali metals. This causes their first ionization enthalpies to be higher than that of alkali metals. However, their second ionization enthalpy is less than the corresponding alkali metals. This is because alkali metals, after losing one electron, acquires noble gas configuration, which is very stable.</p>
<b>(ii)</b>	<p>Basicity of oxides:</p> <p>The oxides of alkali metals are very basic in nature. This happens due to the highly electropositive nature of alkali metals, which makes these oxides highly ionic. Hence, they readily dissociate in water to give hydroxide ions.</p>	<b>(ii)</b>	<p>Basicity of oxides:</p> <p>The oxides of alkaline earth metals are quite basic but not as basic as those of alkali metals. This is because alkaline earth metals are less electropositive than alkali metals.</p>
<b>(iii)</b>	<p>Solubility of hydroxides:</p> <p>The hydroxides of alkali metals are more soluble than those of alkaline earth metals.</p>	<b>(iii)</b>	<p>Solubility of hydroxides:</p> <p>The hydroxides of alkaline earth metals are less soluble than those of alkali metals. This is due to the high lattice energies of alkaline earth metals. Their higher charge densities (as compared to alkali metals) account for higher lattice</p>

**Question 10.7:**

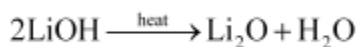
In what ways lithium shows similarities to magnesium in its chemical behaviour?

Answer

Similarities between lithium and magnesium are as follows.

**(i)** Both Li and Mg react slowly with cold water.

**(ii)** The oxides of both Li and Mg are much less soluble in water and their hydroxides decompose at high temperature.

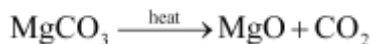
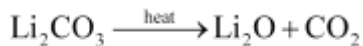


**(iii)** Both Li and Mg react with  $\text{N}_2$  to form nitrides.



**(iv)** Neither Li nor Mg form peroxides or superoxides.

**(v)** The carbonates of both are covalent in nature. Also, these decompose on heating.



**(vi)** Li and Mg do not form solid bicarbonates.

**(vii)** Both LiCl and  $\text{MgCl}_2$  are soluble in ethanol owing to their covalent nature.

**(viii)** Both LiCl and  $\text{MgCl}_2$  are deliquescent in nature. They crystallize from aqueous solutions as hydrates, for example,  $\text{LiCl} \cdot 2\text{H}_2\text{O}$  and  $\text{MgCl}_2 \cdot 8\text{H}_2\text{O}$ .

**Question 10.8:**

Explain why alkali and alkaline earth metals cannot be obtained by chemical reduction methods?

Answer

In the process of chemical reduction, oxides of metals are reduced using a stronger reducing agent. Alkali metals and alkaline earth metals are among the strongest

reducing agents and the reducing agents that are stronger than them are not available. Therefore, they cannot be obtained by chemical reduction of their oxides.

**Question 10.9:**

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

Answer

All the three, lithium, potassium, and cesium, are alkali metals. Still, K and Cs are used in the photoelectric cell and not Li.

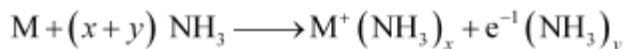
This is because as compared to Cs and K, Li is smaller in size and therefore, requires high energy to lose an electron. While on the other hand, K and Cs have low ionization energy. Hence, they can easily lose electrons. This property of K and Cs is utilized in photoelectric cells.

**Question 10.10:**

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

Answer

When an alkali metal is dissolved in liquid ammonia, it results in the formation of a deep blue coloured solution.



The ammoniated electrons absorb energy corresponding to red region of visible light. Therefore, the transmitted light is blue in colour.

At a higher concentration (3 M), clusters of metal ions are formed. This causes the solution to attain a copper-bronze colour and a characteristic metallic lustre.

**Question 10.11:**

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

Answer

When an alkaline earth metal is heated, the valence electrons get excited to a higher energy level. When this excited electron comes back to its lower energy level, it radiates energy, which belongs to the visible region. Hence, the colour is observed. In Be and Mg,

the electrons are strongly bound. The energy required to excite these electrons is very high. Therefore, when the electron reverts back to its original position, the energy released does not fall in the visible region. Hence, no colour in the flame is seen.

**Question 10.12:**

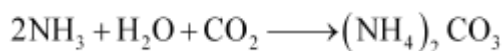
Discuss the various reactions that occur in the Solvay process.

Answer

Solvay process is used to prepare sodium carbonate.

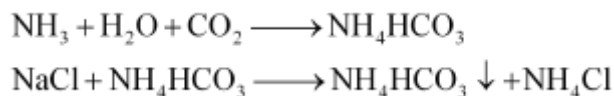
When carbon dioxide gas is bubbled through a brine solution saturated with ammonia, sodium hydrogen carbonate is formed. This sodium hydrogen carbonate is then converted to sodium carbonate.

**Step 1:** Brine solution is saturated with ammonia.



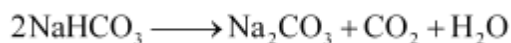
This ammoniated brine is filtered to remove any impurity.

**Step 2:** Carbon dioxide is reacted with this ammoniated brine to result in the formation of insoluble sodium hydrogen carbonate.



**Step 3:** The solution containing crystals of  $\text{NaHCO}_3$  is filtered to obtain  $\text{NaHCO}_3$ .

**Step 4:**  $\text{NaHCO}_3$  is heated strongly to convert it into  $\text{Na}_2\text{CO}_3$ .



**Step 5:** To recover ammonia, the filtrate (after removing  $\text{NaHCO}_3$ ) is mixed with  $\text{Ca}(\text{OH})_2$  and heated.



The overall reaction taking place in Solvay process is



**Question 10.13:**

Potassium carbonate cannot be prepared by Solvay process. Why?

Answer



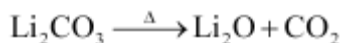
Solvol process cannot be used to prepare potassium carbonate. This is because unlike sodium bicarbonate, potassium bicarbonate is fairly soluble in water and does not precipitate out.

**Question 10.14:**

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

Answer

As we move down the alkali metal group, the electropositive character increases. This causes an increase in the stability of alkali carbonates. However, lithium carbonate is not so stable to heat. This is because lithium carbonate is covalent. Lithium ion, being very small in size, polarizes a large carbonate ion, leading to the formation of more stable lithium oxide.



Therefore, lithium carbonate decomposes at a low temperature while a stable sodium carbonate decomposes at a high temperature.

**Question 10.15:**

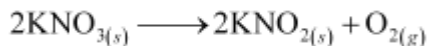
Compare the solubility and thermal stability of the following compounds of the alkali metals with those of the alkaline earth metals. (a) Nitrates (b) Carbonates (c) Sulphates.

Answer

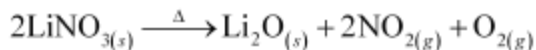
(i) Nitrates

**Thermal stability**

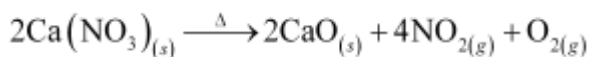
Nitrates of alkali metals, except  $\text{LiNO}_3$ , decompose on strong heating to form nitrites.



$\text{LiNO}_3$ , on decomposition, gives oxide.



Similar to lithium nitrate, alkaline earth metal nitrates also decompose to give oxides.



As we move down group 1 and group 2, the thermal stability of nitrate increases.

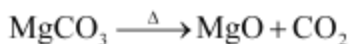
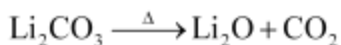
## Solubility

Nitrates of both group 1 and group 2 metals are soluble in water.

### (ii) Carbonates

#### Thermal stability

The carbonates of alkali metals are stable towards heat. However, carbonate of lithium, when heated, decomposes to form lithium oxide. The carbonates of alkaline earth metals also decompose on heating to form oxide and carbon dioxide.



## Solubility

Carbonates of alkali metals are soluble in water with the exception of  $\text{Li}_2\text{CO}_3$ . Also, the solubility increases as we move down the group.

Carbonates of alkaline earth metals are insoluble in water.

### (iii) Sulphates

#### Thermal stability

Sulphates of both group 1 and group 2 metals are stable towards heat.

## Solubility

Sulphates of alkali metals are soluble in water. However, sulphates of alkaline earth metals show varied trends.

$\text{BeSO}_4$  Fairly soluble

$\text{MgSO}_4$  Soluble

$\text{CaSO}_4$  Sparingly soluble

$\text{SrSO}_4$  Insoluble

$\text{BaSO}_4$  Insoluble

In other words, while moving down the alkaline earth metals, the solubility of their sulphates decreases.

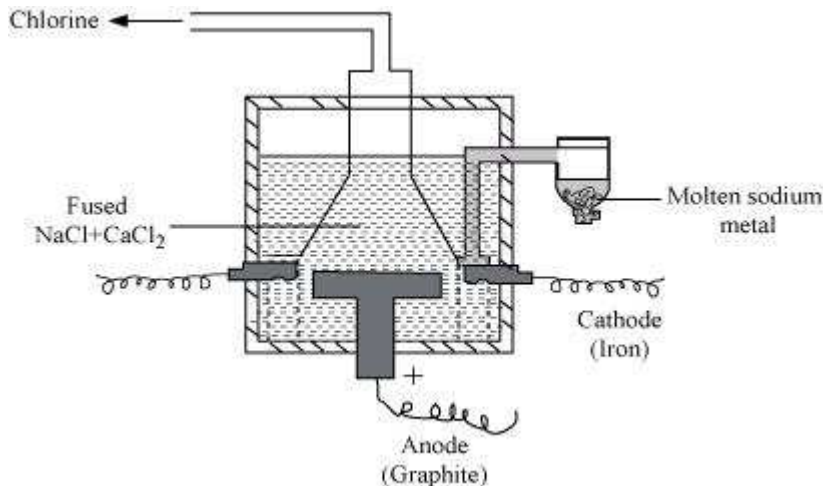
### Question 10.16:

Starting with sodium chloride how would you proceed to prepare (i) sodium metal (ii) sodium hydroxide (iii) sodium peroxide (iv) sodium carbonate?

Answer

(a) Sodium can be extracted from sodium chloride by Downs process.

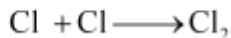
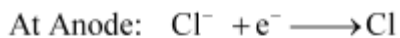
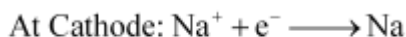
This process involves the electrolysis of fused NaCl (40%) and CaCl<sub>2</sub> (60 %) at a temperature of 1123 K in Downs cell.



Steel is the cathode and a block of graphite acts as the anode. Metallic Na and Ca are formed at cathode. Molten sodium is taken out of the cell and collected over kerosene.

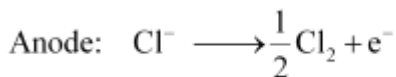
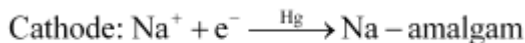


Molten



**(ii)** Sodium hydroxide can be prepared by the electrolysis of sodium chloride. This is called Castner-Kellner process. In this process, the brine solution is electrolysed using a carbon anode and a mercury cathode.

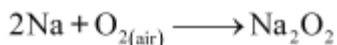
The sodium metal, which is discharged at cathode, combines with mercury to form an amalgam.



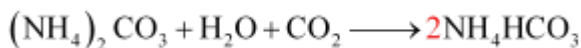
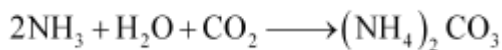
**(iii)** Sodium peroxide

First, NaCl is electrolysed to result in the formation of Na metal (Downs process).

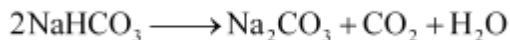
This sodium metal is then heated on aluminium trays in air (free of CO<sub>2</sub>) to form its peroxide.



**(iv)** Sodium carbonate is prepared by Solvay process. Sodium hydrogen carbonate is precipitated in a reaction of sodium chloride and ammonium hydrogen carbonate.



These sodium hydrogen carbonate crystals are heated to give sodium carbonate.



### Question 10.17:

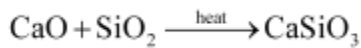
What happens when (i) magnesium is burnt in air (ii) quick lime is heated with silica (iii) chlorine reacts with slaked lime (iv) calcium nitrate is heated ?

Answer

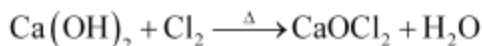
**(i)** Magnesium burns in air with a dazzling light to form MgO and  $\text{Mg}_3\text{N}_2$ .



**(ii)** Quick lime (CaO) combines with silica ( $\text{SiO}_2$ ) to form slag.

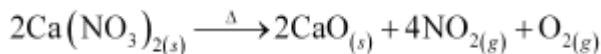


**(iii)** When chloride is added to slaked lime, it gives bleaching powder.



Bleaching  
powder

**(iv)** Calcium nitrate, on heating, decomposes to give calcium oxide.



### Question 10.18:

Describe two important uses of each of the following: (i) caustic soda (ii) sodium carbonate (iii) quicklime.

Answer

**(i)** Uses of caustic soda

**(a)** It is used in soap industry.

**(b)** It is used as a reagent in laboratory.

**(ii)** Uses of sodium carbonate

**(a)** It is generally used in glass and soap industry.

**(b)** It is used as a water softener.

**(iii)** Uses of quick lime

**(a)** It is used as a starting material for obtaining slaked lime.

**(b)** It is used in the manufacture of glass and cement.

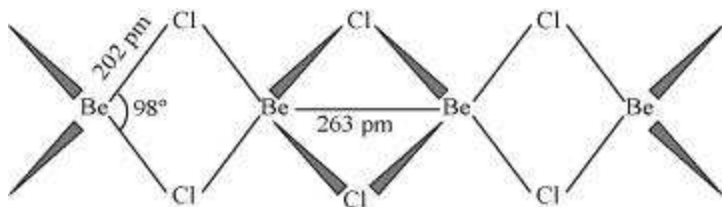
**Question 10.19:**

Draw the structure of (i)  $\text{BeCl}_2$  (vapour) (ii)  $\text{BeCl}_2$  (solid).

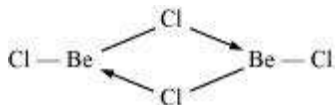
Answer

**(a)** Structure of  $\text{BeCl}_2$  (solid)

$\text{BeCl}_2$  exists as a polymer in condensed (solid) phase.



In the vapour state,  $\text{BeCl}_2$  exists as a monomer with a linear structure.



**Question 10.20:**

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.

Answer

The atomic size of sodium and potassium is larger than that of magnesium and calcium. Thus, the lattice energies of carbonates and hydroxides formed by calcium and magnesium are much more than those of sodium and potassium. Hence, carbonates and hydroxides of sodium and potassium dissolve readily in water whereas those of calcium and magnesium are only sparingly soluble.

### Question 10.21:

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

Answer

**(i)** Chemically, limestone is  $\text{CaCO}_3$ .

#### **Importance of limestone**

(a) It is used in the preparation of lime and cement.

(b) It is used as a flux during the smelting of iron ores.

**(ii)** Chemically, cement is a mixture of calcium silicate and calcium aluminate.

#### **Importance of cement**

(a) It is used in plastering and in construction of bridges.

(b) It is used in concrete.

**(iii)** Chemically, plaster of Paris is  $2\text{CaSO}_4 \cdot \text{H}_2\text{O}$ .

#### **Importance of plaster of Paris**

(a) It is used in surgical bandages.

(b) It is also used for making casts and moulds.

### Question 10.22:

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

Answer

Lithium is the smallest in size among the alkali metals. Hence,  $\text{Li}^+$  ion can polarize water molecules more easily than other alkali metals. As a result, water molecules get attached to lithium salts as water of crystallization. Hence, lithium salts such as trihydrated lithium chloride ( $\text{LiCl} \cdot 3\text{H}_2\text{O}$ ) are commonly hydrated. As the size of the ions increases, their polarizing power decreases. Hence, other alkali metal ions usually form anhydrous salts.

### Question 10.23:

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

Answer

LiF is insoluble in water. On the contrary, LiCl is soluble not only in water, but also in acetone. This is mainly because of the greater ionic character of LiF as compared to LiCl. The solubility of a compound in water depends on the balance between lattice energy and hydration energy. Since fluoride ion is much smaller in size than chloride ion, the lattice energy of LiF is greater than that of LiCl. Also, there is not much difference between the hydration energies of fluoride ion and chloride ion. Thus, the net energy change during the dissolution of LiCl in water is more exothermic than that during the dissolution of LiF in water. Hence, low lattice energy and greater covalent character are the factors making LiCl soluble not only in water, but also in acetone.

#### **Question 10.24:**

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

Answer

Importance of sodium, potassium, magnesium, and calcium in biological fluids:

#### **(i) Sodium (Na):**

Sodium ions are found primarily in the blood plasma. They are also found in the interstitial fluids surrounding the cells.

- (a) Sodium ions help in the transmission of nerve signals.
- (b) They help in regulating the flow of water across the cell membranes.
- (c) They also help in transporting sugars and amino acids into the cells.

#### **(ii) Potassium (K):**

Potassium ions are found in the highest quantity within the cell fluids.

- (a) K ions help in activating many enzymes.
- (b) They also participate in oxidising glucose to produce ATP.
- (c) They also help in transmitting nerve signals.

#### **(iii) Magnesium (Mg) and calcium (Ca):**

Magnesium and calcium are referred to as macro-minerals. This term indicates their higher abundance in the human body system.

- (a) Mg helps in relaxing nerves and muscles.
- (b) Mg helps in building and strengthening bones.
- (c) Mg maintains normal blood circulation in the human body system.
- (d) Ca helps in the coagulation of blood

(e) Ca also helps in maintaining homeostasis.

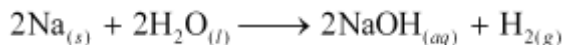
### Question 10.25:

What happens when

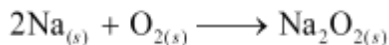
- (i) sodium metal is dropped in water ?
- (ii) sodium metal is heated in free supply of air ?
- (iii) sodium peroxide dissolves in water ?

Answer

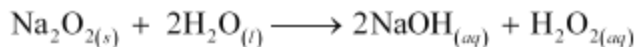
**(i)** When Na metal is dropped in water, it reacts violently to form sodium hydroxide and hydrogen gas. The chemical equation involved in the reaction is:



**(ii)** On being heated in air, sodium reacts vigorously with oxygen to form sodium peroxide. The chemical equation involved in the reaction is:



**(iii)** When sodium peroxide is dissolved in water, it is readily hydrolysed to form sodium hydroxide and water. The chemical equation involved in the reaction is:



### Question 10.26:

Comment on each of the following observations:

- (a) The mobilities of the alkali metal ions in aqueous solution are  $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$
- (b) Lithium is the only alkali metal to form a nitride directly.
- (c)  $E^\circ$  for  $\text{M}^{2+}(\text{aq}) + 2e^- \longrightarrow \text{M}(\text{s})$  (where M = Ca, Sr or Ba) is nearly constant.

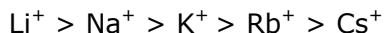
Answer

**(a)** On moving down the alkali group, the ionic and atomic sizes of the metals increase. The given alkali metal ions can be arranged in the increasing order of their ionic sizes as:  $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$

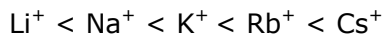
Smaller the size of an ion, the more highly is it hydrated. Since  $\text{Li}^+$  is the smallest, it gets heavily hydrated in an aqueous solution. On the other hand,  $\text{Cs}^+$  is the largest and



so it is the least hydrated. The given alkali metal ions can be arranged in the decreasing order of their hydrations as:



Greater the mass of a hydrated ion, the lower is its ionic mobility. Therefore, hydrated  $\text{Li}^+$  is the least mobile and hydrated  $\text{Cs}^+$  is the most mobile. Thus, the given alkali metal ions can be arranged in the increasing order of their mobilities as:



**(b)** Unlike the other elements of group 1, Li reacts directly with nitrogen to form lithium nitride. This is because  $\text{Li}^+$  is very small in size and so its size is the most compatible with the  $\text{N}^{3-}$  ion. Hence, the lattice energy released is very high. This energy also overcomes the high amount of energy required for the formation of the  $\text{N}^{3-}$  ion.

**(c)** Electrode potential ( $E^\circ$ ) of any  $\text{M}^{2+}/\text{M}$  electrode depends upon three factors:

- (i) Ionisation enthalpy
- (ii) Enthalpy of hydration
- (iii) Enthalpy of vaporisation

The combined effect of these factors is approximately the same for Ca, Sr, and Ba. Hence, their electrode potentials are nearly constant.

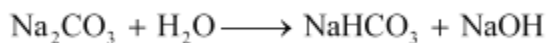
### Question 10.27:

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 ?

Answer

**(a)** When sodium carbonate is added to water, it hydrolyses to give sodium bicarbonate and sodium hydroxide (a strong base). As a result, the solution becomes alkaline.



**(b)** It is not possible to prepare alkali metals by the chemical reduction of their oxides as they themselves are very strong reducing agents. They cannot be prepared by displacement reactions either (wherein one element is displaced by another). This is because these elements are highly electropositive. Neither can electrolysis of aqueous solutions be used to extract these elements. This is because the liberated metals react with water.

Hence, to overcome these difficulties, alkali metals are usually prepared by the electrolysis of their fused chlorides.

**(c)** Blood plasma and the interstitial fluids surrounding the cells are the regions where sodium ions are primarily found. Potassium ions are located within the cell fluids. Sodium ions are involved in the transmission of nerve signals, in regulating the flow of water across the cell membranes, and in transporting sugars and amino acids into the cells. Hence, sodium is found to be more useful than potassium.

### Question 10.28:

Write balanced equations for reactions between

(a)  $\text{Na}_2\text{O}_2$  and water

(b)  $\text{KO}_2$  and water

(c)  $\text{Na}_2\text{O}$  and  $\text{CO}_2$

Answer

**(a)** The balanced chemical equation for the reaction between  $\text{Na}_2\text{O}_2$  and water is:



**(b)** The balanced chemical equation for the reaction between  $\text{KO}_2$  and water is:



**(c)** The balanced chemical equation for the reaction between  $\text{Na}_2\text{O}$  and  $\text{CO}_2$  is:



### Question 10.29:

How would you explain the following observations?

(i)  $\text{BeO}$  is almost insoluble but  $\text{BeSO}_4$  is soluble in water,

(ii)  $\text{BaO}$  is soluble but  $\text{BaSO}_4$  is insoluble in water,

(iii)  $\text{LiI}$  is more soluble than  $\text{KI}$  in ethanol.

Answer

**(i)**  $\text{BeO}$  is almost insoluble in water and  $\text{BeSO}_4$  is soluble in water.  $\text{Be}^{2+}$  is a small cation with a high polarising power and  $\text{O}^{2-}$  is a small anion. The size compatibility of  $\text{Be}^{2+}$  and  $\text{O}^{2-}$  is high. Therefore, the lattice energy released during their formation is also very high. When  $\text{BeO}$  is dissolved in water, the hydration energy of its ions is not sufficient to overcome the high lattice energy. Therefore,  $\text{BeO}$  is insoluble in water. On the other

hand,  $\text{SO}_4^{2-}$  ion is a large anion. Hence,  $\text{Be}^{2+}$  can easily polarise  $\text{SO}_4^{2-}$  ions, making  $\text{BeSO}_4$  unstable. Thus, the lattice energy of  $\text{BeSO}_4$  is not very high and so it is soluble in water.

**(ii)**  $\text{BaO}$  is soluble in water, but  $\text{BaSO}_4$  is not.  $\text{Ba}^{2+}$  is a large cation and  $\text{O}^{2-}$  is a small anion. The size compatibility of  $\text{Ba}^{2+}$  and  $\text{O}^{2-}$  is not high. As a result,  $\text{BaO}$  is unstable. The lattice energy released during its formation is also not very large. It can easily be overcome by the hydration energy of the ions. Therefore,  $\text{BaO}$  is soluble in water. In  $\text{BaSO}_4$ ,  $\text{Ba}^{2+}$  and  $\text{SO}_4^{2-}$  are both large-sized. The lattice energy released is high. Hence, it is not soluble in water.

**(iii)**  $\text{LiI}$  is more soluble than  $\text{KI}$  in ethanol. As a result of its small size, the lithium ion has a higher polarising power than the potassium ion. It polarises the electron cloud of the iodide ion to a much greater extent than the potassium ion. This causes a greater covalent character in  $\text{LiI}$  than in  $\text{KI}$ . Hence,  $\text{LiI}$  is more soluble in ethanol.

### Question 10.30:

Which of the alkali metal is having least melting point?

(a) Na (b) K (c) Rb (d) Cs

Answer

Atomic size increases as we move down the alkali group. As a result, the binding energies of their atoms in the crystal lattice decrease. Also, the strength of metallic bonds decreases on moving down a group in the periodic table. This causes a decrease in the melting point. Among the given metals, Cs is the largest and has the least melting point.

### Question 10.31:

Which one of the following alkali metals gives hydrated salts?

(a) Li (b) Na (c) K (d) Cs

Answer

Smaller the size of an ion, the more highly is it hydrated. Among the given alkali metals, Li is the smallest in size. Also, it has the highest charge density and highest polarising power. Hence, it attracts water molecules more strongly than the other alkali metals. As a result, it forms hydrated salts such as

$\text{LiCl} \cdot 2 \text{H}_2\text{O}$ . The other alkali metals are larger than Li and have weaker charge densities. Hence, they usually do not form hydrated salts.

**Question 10.32:**

Which one of the alkaline earth metal carbonates is thermally the most stable?

(a)  $\text{MgCO}_3$  (b)  $\text{CaCO}_3$  (c)  $\text{SrCO}_3$  (d)  $\text{BaCO}_3$

Answer

Thermal stability increases with the increase in the size of the cation present in the carbonate. The increasing order of the cationic size of the given alkaline earth metals is

$\text{Mg} < \text{Ca} < \text{Sr} < \text{Ba}$

Hence, the increasing order of the thermal stability of the given alkaline earth metal carbonates is

$\text{MgCO}_3 < \text{CaCO}_3 < \text{SrCO}_3 < \text{BaCO}_3$