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

Ans. Common physical properties of alkali metals are:

(a) Large atomic radii. The atomic radii of alkali metals are largest in their respective periods. This increase as we travel down the group from Li to Cs.

(b) The ionisation enthalpies of the alkali metals are lowest as compared to the elements in the other groups.

(c) They all show +1 oxidation state.

(d) The elements of this group are typical metals and are soft.

(e) All of them form ionic bond in their compound though the ionic character increases down the group.

(f) All alkali metals impart a characteristic colour to the flame.

Common chemical properties of alkali metals are:

(a) All alkali metals are highly reactive and have the reducing property.

(b) Alkali metal react with water to release hydrogen.

(c) All the alkali metals on exposure to atmosphere (air and moisture) get converted into oxides, hydroxides and finally to carbonates.

(d) Alkali metals react vigorously with halogens to form metal halides of the type MX.

(e) The metals and their oxides on reaction with water give a strong alkali.

(f) All alkali metals dissolve in liquid ammonia giving highly conducting deep blue solutions.

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

Ans. **Trend in physical properties**

(a) The atomic radii of alkaline earth metals are fairly large though smaller than the corresponding alkali metals and they increase down the group.

(b) The alkaline earth metals have fairly low ionisation enthalpies though greater than those of the corresponding elements of group–I and they decrease down the group.

(c) They all exhibit the oxidation state of +2 in their solid state as well as in solution.

(d) They are less electropositive than alkali metals but are fairly electropositive and metallic.

(e) Like alkali metals, alkaline earth metals predominantly form ionic bonds in their compounds but are less ionic than alkali metals.

(f) The alkaline earth metals are less reducing than alkali metals. Their reducing character increases down the group.

Chemical properties :

(a) They are less reactive than alkali metals. Be does not react with water while Mg can only react with steam. But as we go down the group, their reactivity with the water increases. They react with water to give hydroxides and hydrogen gas.

(b) Their reactivity towards air is less than alkali metals. Be and Mg are kinetically inert to oxygen but down the group their reactivity increases. They form oxides and nitrides when react in air.

(c) All alkaline earth metals combine with halogens at elevated temperature forming their halides of the type MX_2.

(d) The metals and their oxides are quite basic since they form alkali when treated with water. However, they are less alkaline than group–I metals. The basicities of their oxides increase down the group.

(e) Like alkali metals , alkaline earth metals dissolve in liquid ammonia.

3. Why are alkali metals not found in nature ?

Ans. Alkali metals are highly reactive and hence they do not occur in the free state.

4. Find the oxidation state of sodium in Na_2O_2.

Ans. Let x be the oxidation state of Na in Na_2O_2. Since Na_2O_2 contains a peroxide linkage in which O has an oxidation state of – 1, therefore, Na_2O_2 or $2x + 2(-1) = 0$ or $x = +1$.

5. Explain why sodium is less reactive than potassium?

Ans. The ionisation enthalpy ($\Delta_I H$) of potassium (496 kJ mol^{-1}) is less than that of sodium (520 kJ mol^{-1}) or more precisely the standard electrode potential (E°) of potassium (–2.925V) is more negative than that of sodium (–2.714V) and hence potassium is more reactive than sodium.

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

Ans. (i) **Ionisation enthalpy ($\Delta_I H$)**. Because of higher nuclear charge, the ΔH of alkaline earth metals are higher than those of the corresponding alkali metals.

(ii) **Basicities of oxides**. The oxides of alkali and alkaline earth metals dissolve in water to form their respective hydroxides. These hydroxides are strong bases. However, since the ionisation enthalpy of alkali metals is lower or the electropositive character of alkali metals is higher than that of the corresponding alkaline earth metal, therefore the M– OH bond in alkali metals can more easily ionised ($MOH \rightarrow M^+ + OH^-$), than in alkaline earth metals and hence alkali metal oxides are more basic than the corresponding alkaline earth metal oxides.

(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 metal hydroxides. However, the solubility of the hydroxides of both alkali and alkaline earth metals increase down the group due to larger decrease in their lattices enthalpies as compared to their hydration enthalpies.

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

Ans. Lithium resembles magnesium mainly due to same charge/radius ratio or polarising power. The main points of similarity are:

(a) Both LiOH and $Mg(OH)_2$ are weak bases.

(b) Both form ionic nitrides when heated in atmosphere of nitrogen, Li_3N and Mg_3N_2.

(c) The hydroxides and carbonates of both of them decompose on heating.

$2LiOH \rightarrow Li_2O + H_2O$

$Mg(OH)_2 \rightarrow MgO + H_2O$

$Li_2CO_3 \rightarrow Li_2O + CO_2$

$MgCO_3 \rightarrow MgO + CO_2$

(d) Both Li and Mg do not form solid bicarbonates.

(e) Li and Mg do not form peroxides and super oxides.

(f) Both Li and Mg nitrates decompose on heating producing NO_2.

$4LiNO_3 \rightarrow 2Li_2O + 4NO_2 + O_2$

$2Mg(NO_3)_2 \rightarrow 2MgO + 4NO_2 + O_2$

(g) The hydroxides, carbonates and fluorides of lithium and magnesium are sparingly soluble in water.

(h) LiCl and $MgCl_2$ are highly soluble in ethanol.

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

Ans. Alkali and alkaline earth metals are themselves strong reducing agents. The reducing agents better than them are not available. Therefore, these metals cannot be obtained by reduction of their oxides or chlorides.

9. Why are potassium and caesium, rather than lithium used in photoelectric cells?

Ans. Potassium and caesium have much lower ionisation enthalpy than that of lithium. As a result, these metals on exposure to light, easily emit electrons but lithium does not. Therefore, K and Cs rather than Li are used in photoelectric cells.

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

Ans. The dilute solutions of alkali metals in liquid ammonia exhibit dark blue colour because ammoniated electrons absorb energy corresponding to the red region of the visible light.

$M + (x + y)\,NH_3 \rightarrow [M(NH_3)_x]^+ + \underbrace{e^-(NH_3)_y}_{\text{Ammoniated electrons}}$

However, if the concentration increases above 3 M, the colour changes to copper-bronze and the solution acquires metallic lustre due to the formation of metal ion clusters.

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

Ans. Because of the small size, the ionisation enthalpies of Be and Mg are much higher than those of other alkaline earth metals. Therefore, they need large amount of energy for excitation of electrons to higher energy levels. Since such a large amount of energy is not available in Bunsen flame, therefore, these metals do not impart any colour to the flame.

12. Discuss the various reactions that occur in the Solvay ammonia process.

Ans. In Solvay ammonia process, CO_2 is passed through brine, (i.e., a concentrated solution of NaCl) saturated with ammonia where sodium bicarbonate being sparingly soluble gets precipitated.

$NaCl + NH_3 + CO_2 + H_2O \rightarrow NaHCO_3 + NH_4Cl$ ...(i)

Sodium bicarbonate on heating gives sodium carbonate.

$2\,NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$ ...(ii)

CO_2 needed for the reaction is prepared by heating calcium carbonate and the quick lime, CaO is dissolved in water to form slaked lime, $Ca(OH)_2$

$CaCO_3 \rightarrow CaO + CO_2$,

$CaO + H_2O \rightarrow Ca(OH)_2$ ...(iii)

NH_3 needed for the purpose is prepared by heating NH_4Cl obtained in Eq (i) with $Ca(OH)_2$ obtained in Eq (iii) $2NH_4Cl + Ca(OH)_2 \rightarrow 2NH_3 + CaCl_2 + 2H_2O$

Therefore, the only by product of the reaction is calcium chloride, $CaCl_2$.

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

Ans. Potassium carbonate cannot be prepared by Solvay process because potassium bicarbonate being more soluble than sodium bicarbonate does not get precipitated when CO_2 is passed through a concentrated solution of KCl saturated with ammonia. $KCl + CO_2 + NH_3 + H_2O \rightarrow KHCO_3 + NH_4Cl$

14. Why is Li_2CO_3 decomposes at lower temperature whereas Na_2CO_3 at higher temperature?

Ans. Li_2CO_3 is a salt of a weak acid (H_2CO_3) with a weak base (LiOH). Since the weak base cannot attract CO_2 strongly, therefore, Li_2CO_3 decomposes at lower temperature. On the other hand, NaOH is a much stronger base than LiOH and hence can attract CO_2 more strongly. Therefore, Na_2CO_3 is much more stable than Li_2CO_3 and hence decomposes at much higher temperature than Li_2CO_3.

15. Compare the solubility and thermal stability of the following compounds of the alkali metals with those of alkaline earth metals (a) nitrates (b) carbonates (c) sulphates.

Ans. **Solubility:**

(a) **Alkali metals:** Nitrates, carbonates and sulphates of alkali metals are soluble in water. Their, solubility, however, increases as we move down the group since the lattice energies decrease more rapidly than the hydration energies.

(b) **Alkaline earth metals:** Nitrates of all alkaline earth metals are soluble in water. Their solubility, however, decreases as we move down the group because their hydration energies decrease more rapidly than the lattice energies. The size of CO_3^{2-} and SO_4^{2-} anions is much larger than the cations, therefore, within a particular group lattice energies remain almost constant. Since the hydration energies decrease in own the group, therefore, the solubility of alkaline earth carbonates' and sulphates decrease down the group. However, the hydration energy of Be^{2+} and Mg^{2+} ions overcome the lattice energy factor and therefore, $BeSO_4$ and $MgSO_4$ are readily soluble in water while the solubility of other sulphates decreases down the group from $CaSO_4$ to $BaSO_4$.

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, NO_2 and O_2.

$2M(NO_3)_2 \rightarrow 2\ MO + 4\ NO_2 + O_2$ (M = Be, Mg, Ca, Sr or Ba) The nitrates of Na, K, Rb and Cs decompose to form metal nitrites and O_2.

$2MNO_3 \rightarrow 2MNO_2 + O_2$ (M = Na, K, Rb, Cs)

However, due to diagonal relationship between Li and Mg, lithium nitrate decomposes like $Mg(NO_3)_2$ to form metal oxide, NO_2 and O_2.

$4LiNO_3 \rightarrow 2Li_2O + 4NO_2 + O_2$

(b) **Carbonates**. Carbonates of alkaline earth metals decompose on heating to form metal oxide and CO_2.

$MCO_3 \rightarrow MO + CO_2$ (M = Be, Mg, Ca, Sr, Ba)

Further as the electropositive character of the metal increases down the group, the stability of these metal carbonates increase and hence the temperature of their decomposition increase as shown below:

$BeCO_3$	$MgCO_3$	$CaCO_3$	$SrCO_3$	$BaCO_3$
< 373 K	813 K	1173 K	1563 K	1633 K

Due to diagonal relationship between Li and Mg, Li_2CO_3 decomposes in the same way as $MgCO_3$.

$Li_2CO_3 \rightarrow Li_2O + CO_2$

All other alkali metal carbonates are stable and do not decompose even at high temperatures.

(c) **Sulphates**. Sulphates of alkaline earth metals decompose on heating giving the oxides and SO_3.

$MSO_4 \rightarrow MO + SO_3$

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. For example,

Compound:	$BeSO_4$	$MgSO_4$	$CaSO_4$	$SrSO_4$
Temperature of: decomposition	773 K	1168 K	1422 K	1647 K

Among alkali metals due to diagonal relationship, Li_2SO_4 decomposes like $MgSO_4$ to form the corresponding metal oxide and SO_3.

$$Li_2SO_4 \xrightarrow{\Delta} Li_2O + SO_3$$

$$MgSO_4 \xrightarrow{\Delta} MgO + SO_3$$

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

16. Starting with sodium chloride how would you prepare

(i) sodium metal (ii) sodium hydroxide

(iii) sodium peroxide and (iv) sodium carbonate.

Ans. (i) Sodium metal is manufactured by electrolysis of a fused mixture of NaCl (40%) and $CaCl_2$ (60%) in Down's cell at 873 K using iron cathode and graphite anode. Na is produced at the anode while Cl_2 is evolved at the cathode.

At cathode : Na^+ (melt) + $e^- \rightarrow$ Na (*l*)

At anode : 2 Cl^- (melt) $\rightarrow Cl_2$ (g) + $2e^-$

(ii) Sodium hydroxide is manufactured by electrolysis of aqueous NaCl (brine) in Castner–Kellner cell using mercury cathode and carbon anode. Sodium metal which is discharged at the cathode combines with mercury to form sodium amalgam. Cl_2 gas is evolved at the anode.

Cathode: $Na^+ + e^- \rightarrow \underset{\text{Sodium}}{Na}$

$2Na + xHg \rightarrow \underset{\text{Sodium amalgam}}{Na_2Hg_x}$,

Anode: $2Cl^- \rightarrow Cl_2 + 2e^-$

The amalgam thus obtained is treated with water to form sodium hydroxide and hydrogen gas.

$Na_2Hg_x + 2\ H_2O \rightarrow 2\ NaOH + xHg + H_2$

(iii) Sodium peroxide is obtained by heating sodium in excess of air. The initially formed sodium oxide reacts with more O_2 to form Na_2O_2.

$4\ Na + O_2 \rightarrow 2\ Na_2O$ $2\ Na_2O + O_2 \rightarrow 2\ Na_2O_2$

(iv) Sodium carbonate is obtained by Solvay ammonia process.

When CO_2 is passed through a concentrated solution of brine saturated with NH_3, $NaHCO_3$ gets precipitated. $NaHCO_3$ on subsequent heating gives Na_2CO_3.

$NaCl + NH_3 + CO_2 + H_2O \rightarrow NaHCO_3 + NH_4Cl$

$$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + CO_2 + H_2O$$

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?

Ans. (i) $2\,Mg\,(s) + O_2\,(g) \rightarrow 2\,MgO\,(s)$

(ii) $\underset{\text{Quick lime}}{CaO(s)} + \underset{\text{Silica}}{SiO_2(s)} \xrightarrow{\Delta} \underset{\text{Calcium silicate}}{CaSiO_3(s)}$

(iii) It reacts with Cl_2 to form calcium hypochlorite, $Ca(OCl)_2$—a constituent of bleaching powder

$2Ca(OH)_2 + 2Cl_2 \rightarrow CaCl_2 + CaOCl_2 + 2H_2O$

Slaked lime Bleaching powder

(iv) $2Ca(NO_3)_2(s) \rightarrow 2CaO(s) + 4NO_2(g) + O_2\,(g)$

18. Describe two important uses of the following:

(i) caustic soda, (ii) sodium carbonate (iii) quicklime.

Ans. 1. Caustic soda,

(i) It is used in the manufacture of soap, paper, artificial silk, etc. and in petroleum refining and purification of bauxite,

(ii) It is used in the textile industries for mercerizing cotton fabrics.

2. Sodium carbonate,

(i) It is used in water softening, laundry and cleaning.

(ii) It is used in the manufacture of glass, soap, borax, etc. and in paper, paints and textile industries.

3. Quick lime,

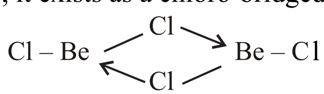
(i) It is used in the manufacture of sodium carbonate from caustic soda,

(ii) It is employed in the purification of sugar and in the manufacture of dyestuffs.

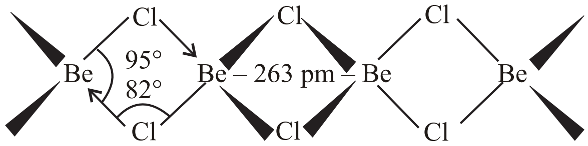
19. Draw the structure of

(i) $BeCl_2$ (vapour) (ii) $BeCl_2$ (solid).

Ans. In the vapour state, it exists as a chloro-bridged dimer.



In the solid state, $BeCl_2$ has polymeric structure with chloro bridges.



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.

Ans. Due to larger size of Na and K as compared to that of Mg and Ca, the lattice energies of hydroxides and carbonates of sodium and potassium are much lower than those of the hydroxides and carbonates of magnesium and calcium. As a result, the hydroxides of Na and K are easily soluble in water while the corresponding salts of magnesium and calcium are sparingly soluble in water.

21. Describe the importance of the following:

(i) limestone (ii) cement

(iii) Plaster of Paris.

Ans. (i) **Limestone**. Specially precipitated $CaCO_3$ is extensively used in the manufacture of high quality paper. It is also used as an antacid, mild abrasive in toothpaste, a constituent of chewing gum and as a filler in 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 extensively used in the building industry as well as in plasters. It is used in dentistry, in ornamental work and for taking the casts of statues and busts. It is also used for immobilising the affected part of organ where there is bone fracture or sprain.

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

Ans. Because of smallest size among alkali metals, Li^+ can polarise water molecules more easily than the other alkali metal ions and hence get attached to lithium salts as water of crystallisation. For example, lithium chloride crystallises as $LiCl.2\,H_2O$.

23. Why LiF is almost insoluble in water whereas LiCl is soluble not only in water but also in acetone?

Ans. Difference in lattice energy and hydration energy of LiCl is higher, i.e., –31 kJmol^{-1} [– 876 – (– 845)] than that of LiF, i.e., –14 kJ mol^{-1} [–1019 – (– 1005)] and hence LiF is sparingly soluble in water while LiCl is soluble. In nut shell, we can say that LiF is almost insoluble in water because of much higher lattice energy (– 1005 kJ mol^{-1}) than that of LiCl (–845 kJ mol^{-1}). Furthermore, Li^+ ion can polarise bigger Cl^- ion more easily than the smaller F^- ion. As a result, according to Fajan rules, LiCl has more covalent character than LiF and hence is soluble in organic solvents like acetone

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

Ans. Biological importance of sodium, potassium, magnesium and calcium in biological fluids:

(a) **Sodium and potassium.** Na^+ ions are primarily found outside the cells in the blood plasma and other interstitial fluids , K^+ ions are present inside the cell. These ions help in transmission of nerve signals, in regulating the flow of water across the cell membranes and in the transport of sugars and amino acids into the cells. As K^+ ions are the most abundant cations within the cell fluids, they activate many enzymes and participate in the oxidation of glucose to ATP. The sodium–potassium pump operates across the cell membranes which consume more than one-third of the ATP used by resting animal.

(b) **Magnesium and calcium.** Mg^{2+} ions are concentrated in animal cells and Ca^{2+} ions are concentrated in the body fluids outside the cell. Both Mg^{2+} and Ca^{2+} catalyse a number of enzymatic reactions. Energy is stored in form of phosphate linkages in ATP. The formation of these linkages, i.e., storage of energy is catalysed by Mg^{2+} and Ca^{2+} ions. Hydrolysis of phosphate linkages is accompanied by release of energy which is catalysed by Ca^{2+} ions.

Mg^{2+} ions are present in chlorophyll —a green colour pigments in plants which absorbs light and is essential for photosynthesis.

Ca^{2+} ions are present in bones and teeth. They are also important in blood clotting and are required to trigger the contraction of muscles and to maintain regular beating of heart.

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?

Ans. (i) H_2 gas is evolved which catches fire due to the exothermicity of the reaction.

$2Na(s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq) + H_2\ (g)$

(ii) Na_2O_2 along with a small amount of Na_2O is formed.

$$2Na(s) + O_2 \longrightarrow \underset{(minor)}{Na_2O(s)} \quad Na_2O(s) + \tfrac{1}{2}O_2(s) \longrightarrow \underset{(major)}{Na_2O_2(s)}$$

(iii) H_2O_2 is formed.

$Na_2O_2(s) + 2\ H_2O(l) \rightarrow 2NaOH\ (aq) + H_2O_2\ (l)$

26. Comment on each of the following observations:

(a) The mobility of alkali metal ions in aqueous solution are $Li^+ < Na^+ < K^+ < Rb^+$.

(b) Lithium is the only alkali metal which forms nitride directly.

(c) E° for $M^{2+}\ (aq) + 2e^- \rightarrow M\ (s)$ (where M = Ca, Sr or Ba) is nearly constant.

Ans. (a) Smaller the size of the ion, more it is hydrated and hence greater is the mass of the hydrated ion and hence smaller is its ionic mobility. Since the extent of hydration decreases in the order: $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

Therefore, ionic mobility increases in the same order: $Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$

(b) Because of the diagonal relationship of Li and Mg, lithium like magnesium forms a nitride while other alkali metals do not.

$6Li(s) + N_2(g) \rightarrow 2Li_3N(s)$

(c) E° of any M^{2+}/M electrode depends upon three factors:

(i) enthalpy of vaporisation,

(ii) ionisation enthalpy

(iii) enthalpy of hydration.

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

27. State as to why

(a) A solution of Na_2CO_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) Na_2CO_3 is a salt of a weak acid, carbonic acid (H_2CO_3) and a strong base, sodium hydroxide (NaOH) therefore, it undergoes hydrolysis to produce strong base NaOH and hence its aqueous solution is alkaline in nature.

$$Na_2CO_3(s) + H_2O(l) \longrightarrow \underset{\text{Strong base}}{2NaOH(aq)} + \underset{\text{Weak acid}}{H_2CO_3(aq)}$$

(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. 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 ions are found primarily in the blood plasma and in the interstitial fluid which surrounds the cells while potassium ions are present within the cell fluids. Sodium ions help 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 the cells. Thus, sodium is found to be more useful than potassium.

28. Write balanced equations for the reactions between

(a) Na_2O_2 and water (b) KO_2 and water

(c) Na_2O and CO_2

Ans. (a) $Na_2O_2(s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq) + H_2O_2\ (aq)$

(b) $2KO_2(s) + 2H_2O\ (l) \rightarrow 2\ KOH\ (aq) + H_2O_2\ (aq) + O_2\ (g)$

(c) $Na_2O + CO_2 \rightarrow Na_2CO_3$

29. How would you explain ?

(i) BeO is insoluble but $BeSO_4$ is soluble in water,

(ii) BaO is soluble but $BaSO_4$ is insoluble in water

(iii) LiI is more soluble than KI in ethanol.

Ans. (i) Because of smaller size, higher ionisation enthalpy and higher electro negativity, BeO is essentially covalent and hence is insoluble in water. In contrast, $BeSO_4$ is ionic. Further because of small size of Be^{2+} ion, the hydration energy of $BeSO_4$ is much higher than its lattice energy and hence $BeSO_4$ is highly soluble in water.

(ii) Both BaO and $BaSO_4$ are ionic compounds. However, the size of O^{2-} ion is much smaller than that of the SO_4^{-2} ion. Since a bigger anion stabilises a bigger cation more than a smaller anion stabilises a bigger cation, therefore, the lattice energy of BaO is much smaller than that of $BaSO_4$ and hence BaO is soluble while $BaSO_4$ is insoluble in water.

(iii) Li^+ is much smaller than K^+ ion. Therefore, according to Fajan rule, Li^+ ion can polarise bigger I^- ion to a greater extent than K^+ ion. As a result, LiI is more covalent than KI and hence is more soluble in organic solvents like ethanol.

30. Which of the alkali metals has least melting point ?

(a) Na (b) K

(c) Rb (d) Cs.

Ans. As the size of the metal increases, the strength of metallic bonding decreases and hence its melting point decreases. Since the size of Cs is the biggest, therefore, its melting point is the lowest. Thus, option (d) is correct.

31. Which of the following alkali metals gives hydrated salts ?

(a) Li (b) Na

(c) K (d) Cs

Ans. Among alkali metal ions, Li^+ is the smallest. Therefore, it has the highest charge density and hence attracts the water molecules more strongly than any other alkali metal cation. Thus, option (a) is correct.

32. Thermally the most stable alkaline earth metal carbonate is

(a) $MgCO_3$ (b) $CaCO_3$

(c) $SrCO_3$ (d) $BaCO_3$

Ans. As the electropositive character of the metal increases or the basicities of their hydroxides increases down the group, their thermal stability increases. Thus, $BaCO_3$ is the most stable and hence option (d) is Correct.