NCERT Solutions for Class-XI Chemistry

Chapter-10
NCERT Chemistry Class 11

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

1. Physical properties of alkali metals:

(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

$$2M + 2H_2O \longrightarrow 2M^+ + 2OH^{\ominus} + H_2$$

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

$$2M + H_2 \longrightarrow 2M^+H^-$$

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

$$2M + Cl_2 \longrightarrow 2MCl$$

$$(M = Li, K, Rb, Cs)$$

Since 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.

$$M + (x + y)NH_3 \longrightarrow [M(NH_3)_x]^+ + [M(NH_3)_y]^-$$

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.

$$M^+_{(cm)} + e^- + NH_{3(l)} \longrightarrow MNH_{(cm)} + \frac{1}{2}H_{2(g)}$$

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

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

2. General characteristics:

(i) (Noble gas) ns^2 is the electronic configuration of alkaline earth metal.

(ii) To occupy the nearest inert gas configuration, these metals lose two of their electrons; and so its oxidation state is +2.

(iii) The ionic radii and atomic radii is smaller than alkali metals. When they moved down towards the group, there is an increase in ionic radii and atomic radii due to decrease in effective nuclear charge.

(iv) The ionization enthalpy is low because the alkaline earth metals are larger in size. The first ionization enthalpy is higher than metals of group 1.

(v) They appear in lustrous and silvery white. They are soft as alkali metals.

(vi) Factors that cause alkaline earth metals to contain high boiling point and melting point:

(*) Atoms of alkali metals are larger than that of alkaline earth metals.

(*) The strong metallic bonds are formed by two valence electrons.

(vii) Ca- brick red, Sr- crimson red, Ba-apple green results in colours to flames. Electrons are bounded strongly to get excited in Be and Mg. Therefore, they do not expose any colours to the flame.

The alkali metals are more reactive than alkaline earth metals. Chemical properties:

(i) Reaction with air and water: Due to the formation of oxide layer on their surface, beryllium and Mg are most inert to water and air.

(a) Beo and Be3 N2 is formed when powdered Be is burnt in air.

(b) For the formation of MgO and Mg3N2, Mg is burnt in the air with dazzling sparkle. Since Mg is more electropositive.

(c) The formation of respective nitrides and oxides is by instant reaction of Sr, Ca, and Ba with air.

(d) Ca, Sr, and Ba can able to react vigorously even with water which is cold.

(ii) when they react with halogens, halides are formed at high temperature.

$$M + X_2 \longrightarrow MX_2 \left(X = F, Cl, Br, l\right)$$

(iii) Except Be, all the alkaline earth metals react with hydrogen to form hydrides.

(iv) alkaline earth metals instantly react with acids to form salts with the liberation of hydrogen gas.

$M + 2HCl \longrightarrow MCl_2 + H_{2(X)}$

(e) Reducing Nature: Alkaline earth metals are strong reducing agents like alkali metals, but the reducing power is less when compared to alkali metals. In general, the reducing character increases from top to bottom.

(f) Solutions in liquid ammonia: Alkaline earth metals dissolve in liquid ammonia to give deep blue black solutions like alkali metals.

3. Why are alkali metals not found in nature?

3. 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.

4. Find the oxidation state of sodium in Na_2O_2.

4. Let the oxidation state of Na be y.

In case of peroxides, the oxidation state of oxygen is -1.

Therefore,

2(y) + 2(-1) = 0

2y – 2 = 0

2y = 2

y = +1

Therefore, the oxidation state of Na is +1.

5. Explain why is sodium less reactive than potassium?

5. 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.

6. Compare the alkali metals and alkaline earth metals with respect to

(i) ionization enthalpy

(ii) basicity of oxides and

(iii)solubility of hydroxides.

6.

Sr. No	Alkaline earth metals	Alkali Metals
(a)	**Solubility of hydroxide:** They are less soluble compared to alkali metals as it has high lattice energy and is having higher charge densities account for higher lattice.	**Solubility of hydroxide:** They are more soluble compared to alkaline earth metals.
(b)	**Ionization Enthalpy:** They have smaller atomic size and higher effective	**Ionization Enthalpy:** They have large atomic size compared to

	nuclear charge compared to alkali metals, which causes their 1st ionization enthalpy higher than that in alkali metals, but the 2nd ionization enthalpy is less than that of alkali metals	alkaline earth metals, so they are having less 1st ionization enthalpy, so they lose valance electrons very easily.
(c)	**Basicity of oxides:** Their oxides are quite basic but less as compared to those of alkali metals as they are less electro positive than alkali metals.	**Basicity of oxides:** Their oxides are basic in nature as they are highly electropositive, which makes their oxides highly ionic.

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

7. Similarities between lithium and magnesium are as follows:

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

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

$$2LiOH \xrightarrow{heat} Li_2O + H_2O$$

$$Mg(OH)_2 \xrightarrow{heat} MgO + H_2O$$

(3) Both Li and Mg react with N_2 to form nitrides.

$$6Li + N_2 \xrightarrow{heat} 2Li_3N$$

$$3Mg + N_2 \xrightarrow{heat} Mg_3N_2$$

(4) Neither Li nor Mg form peroxides or superoxides.

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

$$Li_2CO_3 \xrightarrow{heat} Li_2O + CO_2$$

$$MgCO_3 \xrightarrow{heat} MgO + CO_2$$

(6) Li and Mg do not form solid bicarbonates.

(7) Both LiCl and MgCl2 are soluble in ethanol owing to their covalent nature.

(8) Both LiCl and MgCl2 are deliquescent in nature. They crystallize from aqueous solutions as hydrates, for example,

$$LiCl \cdot 2H_2O \text{ and } MgCl_2 \cdot 8H_2O$$

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

8. By using a stronger reducing agent, the oxides of metals get reduced by the process called chemical reduction. Alkaline earth metals and alkali metals are strong among the reducing agents. No stronger reducing agent is available than them. Therefore alkaline earth metals and alkali cannot be obtained by chemical reduction of their oxides.

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

9. 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.

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

10. When the alkali metal is dissolved in liquid ammonia, a deep blue coloured solution is formed.

$$M + (x+y)NH_3 \longrightarrow M^{+}(NH_3)_x + e^{-1}(NH_3)_y$$

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

Clusters of metal ions are formed at higher concentration (3M) which causes the solution to attain a copper-bronze colour and a metallic lustre.

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

11. 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.

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

12. The process of preparing sodium carbonate is called Solvay process. Sodium hydrogen carbonate is formed when carbon dioxide gas is bubbled through a brine solution saturated with ammonia. The obtained sodium hydrogen carbonate is then converted into sodium carbonate.

(i) Brine solution is saturated with ammonia.

$$2NH_3 + H_2O + CO_2 \longrightarrow (NH_4)_2CO_3$$

This ammoniated brine is filtered for purity removal.

(ii) When carbondioixide reacts with ammoniated brine, it results in the formation of insoluble sodium hydrogen carbonate.

$(NH_4)_2CO_3 + H_2O + CO_2 \rightarrow 2NH_4HCO_3$

$NH_4HCO_3 + NaCl \rightarrow NaHCO_3\downarrow + NH_4Cl$

(iii) $NaHCO_3$ is obtained by the solution which contains crystals of $NaHCO_3$ is filtered.

(iv) $NaHCO_3$ is heated strongly to convert it into $Na2CO_3$.

$2\ NaHCO_3 \rightarrow Na2CO_3 + CO_2 + H_2O$

(v) The carbon dioxide required for the reaction can be obtained by heating limestone. CaO dissolves in water to form calcium hydroxide which is then transferred to the ammonia recovery tower.

$CaCO_3 \rightarrow CaO + CO_2$
$CaO + H_2O \rightarrow Ca(OH)_2$
(vi) Ammonia is recovered when the filterate which is removed after NaHCO3 is mixed with Ca(OH)2 and heated.

$$Ca(OH)_2 + 2NH_4Cl \longrightarrow NH_3 + 2H_2O + CaCl_2$$

The overall reaction taking place in Solvay process is

$$2NaCl + CaCO_3 \longrightarrow Na_2CO_3 + CaCl_2$$

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

13. Solvay 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.

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

14. The electropositive character increases while moving down in the group of alkali metal which results in an increase in stability of alkali carbonates. Generally, lithium carbonate is not stable when it reacts to heat because lithium carbonate is covalent. Due to the smaller size of lithium ion it polarizes large carbonate ion which results in the formation of stable lithium oxide.

$$Li_2CO_3 \longrightarrow Li_2O + CO_2$$

This is why sodium carbonate decomposes at high temperature and lithium carbonate decomposes at low temperature.

15. Compare the solubility and thermal stability of the following compounds of the alkali metals with those of the alkaline earth metals.
(i) Nitrates
(ii) Carbonates
(iii)Sulphates.

15. (i) Nitrates

Thermal stability

Nitrates of alkali metals, except $LiNO_3$, decompose on strong heating to form nitrites.

$$2KNO_{3(s)} \longrightarrow 2KNO_{2(s)} + O_{2(g)}$$

$LiNO_3$, on decomposition, gives oxide.

$$2LiNO_{3(s)} \xrightarrow{\Delta} Li_2O_{(s)} + 2NO_{2(g)} + O_{2(g)}$$

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

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

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.

$$Na_2CO_3 \xrightarrow{\Delta} \text{No effect}$$

$$Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2$$

$$MgCO_3 \xrightarrow{\Delta} MgO + CO_2$$

Solubility

Carbonates of alkali metals are soluble in water with the exception of Li2CO3. 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.

$BeSO_4$	Fairly soluble
$MgSO_4$	Soluble
$CaSO_4$	Sparingly soluble
$SrSO_4$	Insoluble
$BaSO_4$	Insoluble

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

16. Starting with sodium chloride how would you proceed to prepare

(1) sodium metal

(2) sodium hydroxide

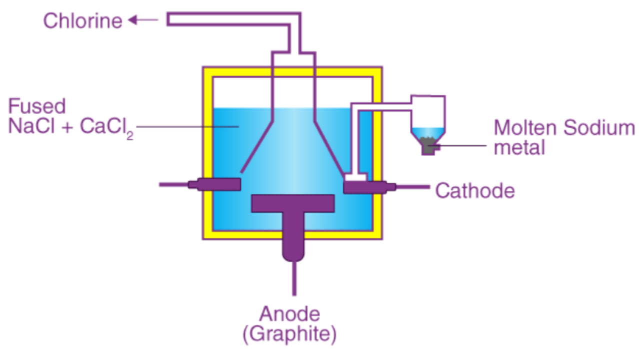
(3) sodium peroxide

(4) sodium carbonate?

16. (i) Sodium metal

Sodium chloride can be converted into sodium by Downs process.

It can be achieved by electrolysis of fused $CaCl_2$ (60 %) and NaCl (40%) at 1123 K in a special apparatus (Downs cell).



A graphite block is the anode while steel is made the cathode. Metallic Ca and Na are formed at the cathode. Molten Na is supported by dipping in kerosene.

$NaCl \xrightarrow{Electrolysis} Na^+ + Cl^-$

(Molten) At Cathode: $Na^+ + e^- \rightarrow Na$

At Anode: $Cl^- \rightarrow Cl + e^-$

$Cl + Cl \rightarrow Cl_2$

(ii) Sodium hydroxide

By electrolyzing a solution of sodium chloride, we can get Sodium hydroxide. This process is commonly known as Castner-Kellner process.

The process is carried out using a mercury cathode and a carbon anode.

Sodium metal, deposited at cathode forms an Amalgam by combining with Mercury.

Cathode: $Na^+ + e^- \xrightarrow{Hg} Na - Amalgam$

$2Na\text{-}Hg + 2H_2O \rightarrow 2NaOH + H_2 + Hg$

Anode: $Cl^- \rightarrow \frac{1}{2} Cl2 + e^-$

(iii) Sodium peroxide

After Na metal is gotten from Downs process, it is heated on Aluminium trays in presence of air (without CO2) to form Sodium peroxide.

$2Na + O_{2(air)} \rightarrow Na_2O_2$

(iv) Sodium carbonate

Sodium hydrogen carbonate is obtained as a precipitate by reacting sodium chloride with ammonium hydrogen carbonate. The resultant crystals can be heated to obtain Sodium Carbonate.

$2NH_3 + H_2O + CO_2 \rightarrow (NH_4)2CO_3$

$(NH_4)_2CO_3 + H_2O + CO_2 \rightarrow 2NH_4HCO_3$

$2NH_4HCO_3 + NaCl \rightarrow NH_4Cl + NaHCO_3$

The resultant crystals can be heated to obtain Sodium Carbonate.

$2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$

17. What happens when

(1) magnesium is burnt in air

(2) quick lime is heated with silica

(3) chlorine reacts with slaked lime

(4) calcium nitrate is heated?

17. (1) Magnesium burns in air with a dazzling light to form MgO and Mg_3N_2.

$$2Mg + O_2 \xrightarrow{Burning} 2MgO$$

$$3Mg + N_2 \xrightarrow{Burning} Mg_3N_2$$

(2) Quick lime (CaO) combines with silica (SiO_2) to form slag.

$$CaO + SiO_2 \xrightarrow{heat} CaSiO_3$$

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

$$Ca(OH)_2 + Cl_2 \xrightarrow{\Delta} \underset{\text{Bleaching powder}}{CaOCl_2} + H_2O$$

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

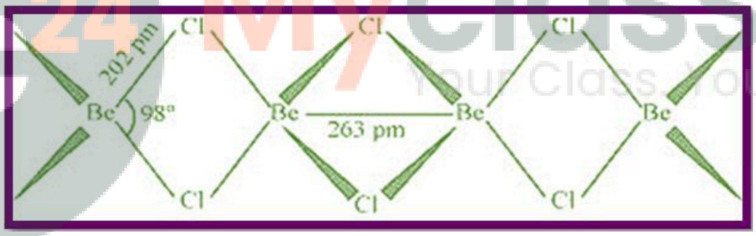
$$2Ca(NO_3)_{2(s)} \xrightarrow{\Delta} 2CaO_{(s)} + 4NO_{2(g)} + O_{2(g)}$$

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

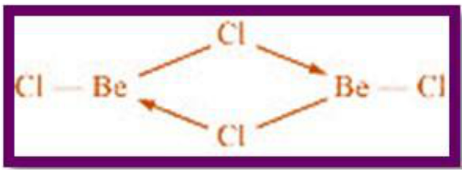
18. (i) Caustic soda
(a) Heavily used in soap industries.
(b) Common reagent in laboratories.
(ii) Sodium carbonate
(a) Finds uses in both soap and glass industries.
(b) Also finds use as a water softener.
(iii) Quick lime
(a)Finds use as a primary material for manufacturing slaked lime.
(b) It helps in the manufacture of cement and glass.

19. Draw the structure
(i) $BeCl_2$ (Solid)
(ii) $BeCl_2$ (Vapour).

19. (i) Structure of $BeCl_2$ (solid)
$BeCl_2$ exists as a polymer in condensed (solid) phase.



(ii) In the vapour state, $BeCl_2$ exists as a monomer with a linear structure.



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.

20. Since the atomic sizes of magnesium and calcium are smaller than that of sodium and potassium, calcium and magnesium form carbonates and hydroxides with higher lattice energies. Thus, they are only sparingly soluble whereas those of potassium and sodium are readily soluble due to low lattice energies.

21. Describe the importance of the following:
(i) limestone

(ii) cement
(iii)plaster of paris

21. (i) Chemically, limestone is $CaCO_3$

Importance of limestone

(a) It is used in the presentation 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 $2CaSO_4$, H_2O.

Importance of plaster of Paris

(a) It is used in surgical bandages.
(b) It is also used for making casts and moulds.

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

22. Since Lithium has the smallest size among all the alkali metals, it can easily polarized water molecules. Thus, smaller the size of the ion, greater is its ability to polarize water molecules.
Hence, trihydrated Lithium Chloride and other Lithium salts can be easily polarized. Due to this reason, other alkali metal ions can only form anhydrous salts.

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

23. 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.

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

24. **Sodium (Na):**
They are found in our blood plasma and the interstitial fluids around the cells. They help in
(a) Transmission of nerve signals.
(b) They regulate the flow of water across the membranes of the neighboring cells.
(c) Transport sugars and amino acids from and to cells.

Potassium (K):

They are found mostly in the cell fluids in greater quantities.

They help in

(a) Activating enzymes.

(b) Oxidising glucose to form ATP.

(c) Transmitting nerve signals.

Magnesium (Mg) and calcium (Ca):

They are also called as macro-minerals named so because of their abundance in our body. Mg helps in (a) Relaxing nerves and muscles.

(b) Building and strengthening bones.

(c) Maintaining blood circulation in our body.

Ca helps in

(a) coagulation of blood

(b) Maintaining homeostasis.

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?

25. (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:

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

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

$$2Na_{(s)} + O_{2(s)} \longrightarrow Na_2O_{2(s)}$$

(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:

$$Na_2O_{2(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_2O_{2(aq)}$$

26. Comment on each of the following observations:

(a) The mobilities of the alkali metal ions in aqueous solution are

$Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$

(b) Lithium is the only alkali metal to form a nitride directly.

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

26. (a) The ionic and atomic sizes of the metals tend to increase while going down the alkali group.

The increasing order of the ionic sizes of the alkali metal ions is as shown below:

$Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$

Smaller the size of an ion, greater is its ability to get hydrated. Li^+ ion gets heavily hydrated since it is the smallest in size whereas Cs+ has the largest size and is the least hydrated ion. The alkali metal ions when arranged in the decreasing order of their hydrations is as shown below:

$Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

Higher the mass of a hydrated ion, the lesser is its ionic mobility. Thus, hydrated Li^+ is the least mobile ion whereas hydrated Cs^+ is the most mobile ion.
The ionic mobility of the alkali metal ions are in the following order:
$Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$
(b) The only metal that can form a nitride directly is Lithium because Li^+ has a smaller size and is easily compatible with the N^{3-} ion. Thus, the lattice energy released is very high which is enough to overcome the amount of energy needed to form N^{3-} ion.
(c) Electrode potential (E°) of any M^{2+}/M electrode is decided by three factors:
(i) Enthalpy of hydration
(ii) Enthalpy of vaporization
(iii) Ionisation enthalpy
The cumulative effect of these factors on Ba, Sr, and Ca is almost the same.
As a result, their electrode potentials are also same.

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?

27. (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.

$$Na_2CO_3 + H_2O \longrightarrow NaHCO_3 + NaOH$$

(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.

28. Write balanced equations for reactions between
(a) Na_2O_2 and water
(b) KO_2 and water
(c) Na2O and CO_2

28. (a) $Na_2O_2 + 2H_2O \rightarrow 2NaOH + H_2O_2$
(b) $2KO_{2(s)} + 2H_2O_{(l)} \rightarrow 2KOH^{(aq)} + H_2O_{2(aq)} + O_{2(aq)}$
(c) $Na_2O_{(s)} + CO_{2(g)} \rightarrow Na_2CO_3$

29. How would you explain the following observations?

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

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

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

29. (i) BeO is almost insoluble in water and $BeSO^4$ is soluble in water. Be^{2+} is a small cation with a high polarising power and O^{2-} is a small anion. The size compatibility of Be^{2+} and O^{2-} is high. Therefore, the lattice energy released during their formation is also very high. When BeO is dissolved in water, the hydration energy of its ions is not sufficient to overcome the high lattice energy. Therefore, BeO is insoluble in water. On the other hand, SO_4^{2-} ion is a large anion. Hence, Be^{2+} can easily polarize SO_4^{2-} ions, making $BeSO_4$ unstable. Thus, the lattice energy of $BeSO_4$ is not very high and so it is soluble in water.

(ii) BaO is soluble in water, but $BaSO_4$ is not. Ba_{2+} is a large cation and O^{2-} is a small anion. The size compatibility of Ba^{2+} and O^{2-} is not high. As a result, 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, BaO is soluble in water. In $BaSO_4$, Ba^{2+} and SO_4^{2-} are both large-sized. The lattice energy released is high. Hence, it is not soluble in water.

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

30. Which of the alkali metal is having least melting point?

(a) Na (b) K

(c) Rb (d) Cs

30. (d) Cs

Cs has the least melting point of the given alkali metals since it has the largest size. Due to a larger size,

the binding capability of Cs is limited and the lattice energy released during the formation of its

compounds is less and can be easily broken.

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

(a) Li (b) Na

(c) K (d) Cs

31. 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 $LiCl.2H_2O$. The other alkali metals are larger than Li and have weaker charge densities. Hence, they usually do not form hydrated salts.

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

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

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

32. (d) $BaCO_3$

Thermal stability is directly proportional to the size of the cation i.e., larger the size of the atom, greater is its thermal stability. The biggest cation among the given compounds is Ba.

Thus, $BaCO_3$ will be the most thermal carbonate among the given compounds followed by $SrCO_3$, $CaCO_3$ and $MgCO_3$.