# The *s*-Block Elements

## OBJECTIVE TYPE QUESTIONS

### Multiple Choice Questions (MCQs)

**1.** Which of the following alkali metals gives hydrated salts?
(a) Li (b) Na (c) K (d) Cs

**2.** Which of the following order is correct for the thermal stability of alkali metal carbonates?
(a) $Li_2CO_3 < K_2CO_3 < Na_2CO_3 < Cs_2CO_3 < Rb_2CO_3$
(b) $Cs_2CO_3 < Rb_2CO_3 < K_2CO_3 < Na_2CO_3 < Li_2CO_3$
(c) $Na_2CO_3 < K_2CO_3 < Cs_2CO_3 < Li_2CO_3 < Rb_2CO_3$
(d) $Li_2CO_3 < Na_2CO_3 < K_2CO_3 < Rb_2CO_3 < Cs_2CO_3$

**3.** The reducing power of a metal depends on various factors. Suggest the factor which makes Li, the strongest reducing agent in aqueous solution.
(a) Sublimation enthalpy
(b) Ionisation enthalpy
(c) Hydration enthalpy
(d) Electron- gain enthalpy

**4.** Strongest reducing agent in the aqueous solution is
(a) Na (b) Rb (c) Ca (d) Li

**5.** Which of the following statements is false regarding alkali metals?
(a) Alkali metals are soft and can be cut with the help of knife.
(b) Alkali metals do not occur in free state in nature.
(c) Alkali metals are highly electropositive elements.
(d) Alkali metal hydrides are covalent in character.

**6.** Which of the following does not illustrate the anomalous properties of Li?
(a) The m.pt. and b.pt. of Li are comparatively high.
(b) Li forms a nitride $Li_3N$ unlike group 1 metals.
(c) Li is much softer than the other group 1 metals.
(d) $Li^+$ ion and its compounds are more heavily hydrated than those of the rest of the group.

**7.** The solubility of metal halides depends on their nature, lattice enthalpy and hydration enthalpy of the individual ions. Amongst fluorides of alkali metals, the lowest solubility of LiF in water is due to
(a) ionic nature of lithium fluoride
(b) high lattice enthalpy
(c) high hydration enthalpy for lithium ion
(d) low ionisation enthalpy of lithium atom.

**8.** Which of the following does not show diagonal relationship between beryllium and aluminium?
(a) Both BeO and $Al_2O_3$ are amphoteric in nature.
(b) Both beryllium and aluminium form polymeric covalent hydrides.
(c) Both beryllium and aluminium form nitrides with nitrogen which evolve $NH_3$ with water.
(d) Both metal carbonates are highly stable.

**9.** The formula of nitre is
(a) $KNO_3$ (b) $NaNO_2$ (c) $BaCl_2$ (d) $Na_2CO_3$

**10.** Amphoteric hydroxides react with both alkalies and acids. Which of the following group 2 metal hydroxides is soluble in sodium hydroxide?
(a) $Be(OH)_2$ (b) $Mg(OH)_2$
(c) $Ca(OH)_2$ (d) $Ba(OH)_2$

**11.** Match the elements given in Column I with the properties mentioned in Column II.

| Column I | Column II |
|---|---|
| (i) Li | (A) Insoluble sulphate |
| (ii) Na | (B) Strongest monoacidic base |
| (iii) Ca | (C) Most negative $E^\circ$ value among alkali metals |
| (iv) Ba | (D) Insoluble oxalate |
| | (E) $6s^2$ outer electronic configuration |

(a) (i) → (C, D), (ii) → (B), (iii) → (D), (iv) → (A, E)
(b) (i) → (C), (ii) → (A), (iii) → (B), (iv) → (D, E)
(c) (i) → (A), (ii) → (D), (iii) → (E), (iv) → (B, C)
(d) (i) → (B, C), (ii) → (A), (iii) → (E), (iv) → (D)

**12.** If $Na^+$ ion is larger than $Mg^{2+}$ ion and $S^{2-}$ ion is larger than $Cl^-$ ion, which of the following will be least soluble in water?
(a) Sodium chloride (b) Sodium sulphide
(c) Magnesium chloride
(d) Magnesium sulphide

**13.** Alkali metals form hydrated compounds. The hydration enthalpies of alkali metals is in the sequence
(a) $Rb^+ > Li^+ > Na^+ > K^+ > Cs^+$
(b) $Cs^+ > Rb^+ > K^+ > Na^+ > Li^+$
(c) $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$
(d) $K^+ > Na^+ > Li^+ > Rb^+ > Cs^+$

**14.** The alkali metals have low melting points. Which of the following alkali metals is expected to melt if the room temperature rises to 30°C?
(a) Na (b) K
(c) Rb (d) Cs

**15.** $Mg \xrightarrow[\text{Heat}]{\text{Air}} X + Y$

$Y \xrightarrow[\text{Colourless gas}]{H_2O} Z \xrightarrow{H_2O}$ Solution $\xrightarrow[\text{Blue coloured solution}]{CuSO_4}$ (*A*)

Substances *X*, *Y*, *Z* and *A* are respectively
(a) $Mg_3N_2$, MgO, $NH_3$, $CuSO_4 \cdot 5H_2O$
(b) $Mg(NO_3)_2$, MgO, $H_2$, $CuSO_4 \cdot 5H_2O$
(c) MgO, $Mg_3N_2$, $NH_3$, $[Cu(NH_3)_4]SO_4$
(d) $Mg(NO_3)_2$, MgO, $H_2O_2$, $CuSO_4 \cdot 5H_2O$

**16.** Fill up the blanks with appropriate choices.
Lithium and magnesium react slowly with water. Their hydroxides are __(i)__ soluble in water. Carbonates of Li and Mg __(ii)__ easily on heating. Both LiCl and $MgCl_2$ are __(iii)__ in ethanol and are __(iv)__ . They crystallise from their aqueous solutions as __(v)__ .

| | (i) | (ii) | (iii) | (iv) | (v) |
|---|---|---|---|---|---|
| (a) | more | do not decompose | soluble | hygro-scopic | hydrates |
| (b) | less | decompose | soluble | delique-scent | hydrates |
| (c) | freely | sublime | insoluble | delique-scent | anhy-drous |
| (d) | freely | decompose | soluble | hygro-scopic | crystals |

**17.** Which of the following has correct increasing basic strength?
(a) MgO < BeO < CaO < BaO
(b) BeO < MgO < CaO < BaO
(c) BaO < CaO < MgO < BeO
(d) CaO < BaO < BeO < MgO

**18.** Match the column I with column II and mark the appropriate choice.

| | **Column I** | | **Column II** |
|---|---|---|---|
| (A) | Li | (i) | Violet flame |
| (B) | K | (ii) | Golden yellow flame |
| (C) | Na | (iii) | Photoelectric cell |
| (D) | Cs | (iv) | Carbonate decomposes on heating |

(a) (A) → (iv), (B) → (i), (C) → (iii), (D) → (ii)
(b) (A) → (i), (B) → (iii), (C) → (ii), (D) → (iv)
(c) (A) → (iii), (B) → (ii), (C) → (i), (D) → (iv)
(d) (A) → (iv), (B) → (i), (C) → (ii), (D) → (iii)

**19.** In the given reactions,
$Be_2C + H_2O \longrightarrow BeO + X$
$CaC_2 + H_2O \longrightarrow Ca(OH)_2 + Y$
$Mg_2C_3 + H_2O \longrightarrow Mg(OH)_2 + Z$
*X*, *Y* and *Z* are respectively
(a) $CH_4$, $C_2H_2$, $C_3H_8$ (b) $CH_4$, $C_2H_6$, $C_3H_8$
(c) $CH_4$, $C_2H_2$, $C_3H_4$ (d) $C_2H_2$, $C_2H_6$, $C_3H_4$

**20.** Which of the following statements are wrong?
1. Barium is more reducing than magnesium.
2. $Ba(OH)_2$ is more basic than $Be(OH)_2$.
3. $Mg^{2+}$ is precipitated as $MgCO_3$ by ammonium carbonate in presence of ammonium chloride.
4. $MgCl_2$ gives colouration in flame test.

(a) 1 and 2 (b) 3 and 4
(c) 1 and 3 (d) 1, 2 and 3

**21.** Lattice energies of $BeF_2$, $MgF_2$, $CaF_2$ and $BaF_2$ are –2906, – 2610, – 2459 and –2367 kJ $mol^{-1}$ respectively. Hydration energies of $Be^{2+}$, $Mg^{2+}$, $Ca^{2+}$, $Ba^{2+}$ and $F^-$ are –2494, –1921, – 1577, –1305 and – 457 kJ $mol^{-1}$ respectively. Which of the fluorides is soluble in water?
(a) $BeF_2$ (b) $MgF_2$
(c) $CaF_2$ (d) $BaF_2$

**22.** The element *A* burns in nitrogen to give an ionic compound *B*. The compound *B* reacts with water to give *C* and *D*. A solution of *C* becomes milky on bubbling carbon dioxide. What is the nature of compound *D*?

(a) Acidic (b) Basic
(c) Amphoteric (d) Neutral

**23.** The charge/size ratio of a cation determines its polarizing power. Which one of the following sequence represents the increasing order of the polarizing power of the cationic species, $K^+$, $Ca^{2+}$, $Mg^{2+}$, $Be^{2+}$ ?
(a) $Mg^{2+}$, $Be^{2+}$, $K^+$, $Ca^{2+}$
(b) $Be^{2+}$, $K^+$, $Ca^{2+}$, $Mg^{2+}$
(c) $K^+$, $Ca^{2+}$, $Mg^{2+}$, $Be^{2+}$
(d) $Ca^{2+}$, $Mg^{2+}$, $Be^{2+}$, $K^+$

**24.** Metals form basic hydroxides. Which of the following metal hydroxide is the least basic?
(a) $Mg(OH)_2$ (b) $Ca(OH)_2$
(c) $Sr(OH)_2$ (d) $Ba(OH)_2$

**25.** Which of the following is incorrect?
(a) Both $BeCl_2$ and $AlCl_3$ have bridged chloride structure.
(b) Both $BeCl_2$ and $AlCl_3$ are strong Lewis acids.
(c) Both $BeCl_2$ and $AlCl_3$ are covalent compounds.
(d) $BeCl_2$ is weak Lewis acid while $AlCl_3$ is strong Lewis acid.

**26.** Some of the group 2 metal halides are covalent and soluble in organic solvents. Among the following metal halides, the one which is soluble in ethanol is
(a) $BeCl_2$ (b) $MgCl_2$
(c) $CaCl_2$ (d) $SrCl_2$

**27.** The hydroxide, which is best soluble in water, is
(a) $Sr(OH)_2$ (b) $Ba(OH)_2$
(c) $Ca(OH)_2$ (d) $Mg(OH)_2$

**28.** The activity of alkaline earth metals as reducing agents
(a) decreases from Be to Ba
(b) increases from Be to Ba
(c) increases from Be to Ca and decreases from Ca to Ba
(d) decreases from Be to Ca and increases from Ca to Ba.

**29.** For alkali metals, which one of the following trends is incorrect?
(a) Hydration energy : Li > Na > K > Rb
(b) Ionization energy : Li > Na > K > Rb
(c) Density : Li < Na < K < Rb
(d) Atomic size : Li < Na < K < Rb

**30.** The first ionization energies of alkaline earth metals are higher than those of the alkali metals. This is because
(a) there is an increase in the nuclear charge of the alkaline earth metals
(b) there is a decrease in the nuclear charge of the alkaline earth metals
(c) there is no change in the nuclear charge
(d) none of the above.

**31.** Match the column I with column II and mark the appropriate choice.

| | **Column I** | | **Column II** |
|---|---|---|---|
| (A) | Na | (i) | Crimson red |
| (B) | K | (ii) | Yellow |
| (C) | Sr | (iii) | Apple green |
| (D) | Ba | (iv) | Violet |

(a) (A) → (i), (B) → (ii), (C) → (iii), (D) → (iv)
(b) (A) → (ii), (B) → (iv), (C) → (i), (D) → (iii)
(c) (A) → (iv), (B) → (iii), (C) → (ii), (D) → (i)
(d) (A) → (iii), (B) → (iv), (C) → (i), (D) → (ii)

**32.** Group 2 elements form compounds in +2 oxidation state though total energy required to produce $M^{2+}$ ions of these elements is quite high. This is because
(a) these elements predominantly form ionic compounds
(b) lattice enthalpy of bivalent compounds of group 2 elements more than compensates the energy requirement for the formation of +2 ions
(c) their hydration energy is quite high
(d) none of these.

**33.** Which of the following is true for magnesium?
(a) It is more electropositive than sodium.
(b) It is manufactured by electrolysis of aqueous magnesium chloride.
(c) It is a strong reducing agent.
(d) It resembles, in chemical properties, with its diagonally placed element boron in group 13 of the periodic table.

**34.** Choose the incorrect statement in the following.
(a) BeO is almost insoluble but $BeSO_4$ is soluble in water.
(b) BaO is soluble but $BaSO_4$ is insoluble in water.
(c) LiI is more soluble than KI in ethanol.
(d) Both Li and Mg form solid hydrogen carbonates.

**35.** A metal salt solution forms a yellow precipitate with potassium chromate in acetic acid, a white precipitate with dilute sulphuric acid but does not give precipitate with sodium chloride or iodide. The white precipitate obtained when sodium carbonate is added to the metal salt solution will consist of

(a) lead carbonate (b) basic lead carbonate

(c) barium carbonate

(d) strontium carbonate.

**36.** Beryllium and aluminium exhibit many properties which are similar. But the two elements differ in

(a) maximum covalency in compounds

(b) exhibiting amphoteric nature in their oxides

(c) forming covalent halides

(d) forming polymeric hydrides.

**37.** Halides of alkaline earth metals form hydrates such as $MgCl_2 \cdot 6H_2O$, $CaCl_2 \cdot 6H_2O$, $BaCl_2 \cdot 2H_2O$ and $SrCl_2 \cdot 2H_2O$. This shows that halides of group 2 elements

(a) are hygroscopic in nature

(b) act as dehydrating agents

(c) can absorb moisture from air

(d) all of the above.

**38.** Magnesium reacts with an element ($X$) to form an ionic compound. If the ground state electronic configuration of ($X$) is $1s^2\ 2s^2\ 2p^3$, the simplest formula for this compound is

(a) $Mg_2X_3$ (b) $MgX_2$

(c) $Mg_2X$ (d) $Mg_3X_2$

**39.** The correct order of solubility of the sulphates of alkaline earth metals in water is Be > Mg > Ca > Sr > Ba. This is due to

(a) decreasing lattice energy

(b) high heat of solvation for smaller ions like $Be^{2+}$

(c) increase in melting points

(d) increasing molecular weight.

**40.** Magnesium forms $Mg^{2+}$ and not $Mg^{+}$ because

(a) magnesium(II) carbonate is insoluble in water

(b) generally higher oxidation states are preferred by metals

(c) ionic radius of Mg(II) is smaller than of Mg(I)

(d) hydration energy of divalent magnesium ion is higher.

## Case Based MCQs

**Case I :** Read the passage given below and answer the following questions from 41 to 43.

Alkali metals have the lowest ionization energy in their corresponding period in periodic table because they have large size which results in a large distance between the nucleus and the outermost electron. Ionization energy of alkali metals decreases from Li to Cs due to increase in atomic size. First ionization energy of alkali metals is very low but they have very high value of second ionization energy.

$$\underset{[\text{noble gas}]\ ns^1}{M_{(g)}} \xrightarrow{1^{st}\ I.E.} \underset{[\text{noble gas}]}{M^+ + e^-}$$

| Metal | Ionization Energy (kJ mol$^{-1}$) | |
|---|---|---|
| | $IE_1$ | $IE_2$ |
| Li | 520.1 | 7296 |
| Na | 495.7 | 4563 |
| K | 418.6 | 3051 |
| Rb | 402.9 | 2633 |
| Cs | 375.6 | 2230 |

**41.** Alkali metals are characterised by

(a) good conductors of heat and electricity

(b) high melting points

(c) low oxidation potentials

(d) high ionisaiton potentials.

**42.** Metals dissolve in liquid ammonia giving coloured solutions which are conducting in nature. The colour of the solution and reason of its conductance is

(a) yellow, $NH_4^+$

(b) blue, ammoniated metals

(c) orange, $[M(NH_3)_x]^+$

(d) blue, ammoniated electron.

**43.** Alkali metals displace hydrogen from water forming bases due to the reason that

(a) they are far above the hydrogen in electrochemical series based on oxidation potential

(b) they are far below the hydrogen in electrochemical series based on oxidation potential

(c) their ionization potential is less than that of other elements.
(d) they contain only one electron in their outermost shell.

**Case II :** Read the passage given below and answer the following questions from 44 to 46.

Alkaline earth metals are less reactive with water as compared to alkali metals. Their reactivity with water increases down the group. Be does not react with water at all, magnesium reacts only with hot water while other metals Ca, Sr and Ba react with cold water.

Order of the reactivity with water :

Ba > Sr > Ca > Mg

$Be(OH)_2$ is amphoteric, but the hydroxides of Mg, Ca, Sr and Ba are basic. The basic strength increases from Mg to Ba.

| $Be(OH)_2$ | $Mg(OH)_2$ | $Ca(OH)_2$ | $Sr(OH)_2$ | $Ba(OH)_2$ |
|---|---|---|---|---|
| Amphoteric | Weakly basic | Basic | Strongly basic | Strongly basic |

**44.** Which of the following statements is false?
(a) Strontium decomposes water readily than beryllium.
(b) Barium carbonate melts at a higher temperature than calcium carbonate.
(c) Barium hydroxide is more soluble in water than magnesium hydroxide.
(d) Beryllium hydroxide is more basic than barium hydroxide.

**45.** Chemical compound '*A*' is used as flocculant in water and sewage treatment. It reacts with $Na_2CO_3$ to generate caustic soda. When $CO_2$ is bubbled through '*A*', it turns cloudy. What is the chemical formula of '*A*' ?
(a) $CaCO_3$ (b) CaO
(c) $Ca(OH)_2$ (d) $Ca(HCO_3)_2$

**46.** What is *X* in the following reaction?

$$MgCl_2 + 2H_2O \longrightarrow X + 2HCl + H_2O$$

(a) MgO
(b) Mg
(c) $Mg(OH)_2$
(d) Mg(OH)Cl

**Case III :** Read the passage given below and answer the following questions from 47 to 50.

All alkali metals dissolve and form blue solution in liquid ammonia. When alkali metals are dissolved in liquid ammonia, there is a considerable expansion in total volume hence such solutions are called expanded metals. The blue solution of an alkali metal in ammonia shows certain characteristic properties which are explained on the basis of formation of ammoniated (solvated) metal cations and ammoniated electrons in the metal ammonia solution in the following way :

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

$$\text{Thus, } M + (x+y)NH_3 \rightarrow \underset{\text{Ammoniated metal cation}}{[M(NH_3)_x]^+} + \underset{\text{Ammoniated electron}}{[e(NH_3)_y]^-}$$

The blue solution is paramagnetic and has high electrical conductivity due to the presence of unpaired electron in the cavities in ammoniacal solution.

**47.** A metal *M* reacts with $N_2$ to give a compound '*A*' ($M_3N$). '*A*' on heating at high temperature gives back '*M*' and '*A*' on reacting with $H_2O$ gives a gas '*B*'. '*B*' intensifies colour of $CuSO_4$ solution when passes through it. *M* and *B* can be
(a) Mg and $NH_3$
(b) Na and $NH_3$
(c) Li and $NH_3$
(d) Al and $NH_3$

**48.** Sodium dissolves in liquid $NH_3$ to give a deep blue solution. This is due to
(a) ammoniated $Na^+$
(b) ammoniated $Na^-$
(c) formation of $Na^+/Na^-$ pair
(d) ammoniated electrons.

**49.** The increasing order of the density of alkali metals is
(a) Li < K < Na < Rb < Cs
(b) Li < Na < K < Rb < Cs
(c) Cs < Rb < Na < K < Li
(d) Cs < Rb < K < Na < Li

**50.** The reaction between sodium and water can be made less vigorous by
(a) lowering the temperature
(b) adding a little alcohol
(c) amalgamating sodium
(d) adding a little acetic acid.

## Assertion & Reasoning Based MCQs

**For question numbers 51-60, a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.**

*(a) Assertion and reason both are correct statements and reason is correct explanation for assertion.*
*(b) Assertion and reason both are correct statements but reason is not correct explanation for assertion.*
*(c) Assertion is correct statement but reason is wrong statement.*
*(d) Assertion is wrong statement but reason is correct statement.*

**51. Assertion :** The carbonate of lithium decomposes easily on heating to form lithium oxide and $CO_2$.

**Reason :** Lithium being very small in size polarises large carbonate ion leading to the formation of more stable $Li_2O$ and $CO_2$.

**52. Assertion :** Generally alkali and alkaline earth metals form superoxides.

**Reason :** There is a single bond between O and O in superoxides.

**53. Assertion :** $BeSO_4$ is soluble in water while $BaSO_4$ is not.

**Reason :** Hydration energy decreases down the group from Be to Ba and lattice energy remains almost constant.

**54. Assertion :** Lithium salts are hydrated.

**Reason :** Lithium has higher polarising power than other alkali metal group members.

**55. Assertion :** Beryllium carbonate is kept in the atmosphere of carbon dioxide.

**Reason :** Beryllium carbonate is unstable and decomposes to give beryllium oxide and carbon dioxide.

**56. Assertion :** *s*-block elements are highly electropositive.

**Reason :** The valence electrons present in *s*-orbital are loosely held.

**57. Assertion :** The ionization enthalpies of alkali metals increase progressively as we move down the group form Li to Cs.

**Reason :** The ionic radii increase as we move down the group from Li to Cs.

**58. Assertion :** The solubility of the alkaline earth metal hydroxides in water increases with increase in atomic number down the group.

**Reason :** The alkaline earth metal hydroxides are basic in character. The only exception is $Be(OH)_2$ which is amphoteric.

**59. Assertion :** Electrode potential is a measure of the tendency of an element to lose electrons in the aqueous solution. So, the reducing property can be correlated in terms of electrode potentials ($E^\circ$) of alkali metals.

**Reason :** More negative is the electrode potential, higher is the tendency of the element to loose electrons and hence, stronger is the reducing agent.

**60. Assertion :** The first element (Li) of group I differs considerably from the rest of the elements of the same group.

**Reason :** Lithium has small ionic and atomic radii, high electronegativity and ionization enthalpy, high polarising power of its cation and absence of *d*-electrons in its valence shell compared to rest of the elements of the same group.

# SUBJECTIVE TYPE QUESTIONS

## Very Short Answer Type Questions (VSA)

**1.** Which has a higher melting point, sodium or potassium?

**2.** Why are the first ionization enthalpies of group 1 elements low?

**3.** Why are group 2 elements harder than group 1 elements?

**4.** $BeCl_2$ in aqueous solution exists as $[Be(OH)_4]^{2+}$. Explain.

**5.** Find out the oxidation state of sodium in $Na_2O_2$.

**6.** Halides of Be dissolve in organic solvents while those of Ba do not. Why is it so?

**7.** What will be formed when lithium nitrate is heated?

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

**9.** Explain why Cs is used in photoelectric cell?

**10.** Potassium carbonate cannot be prepared by Solvay process. Why ?

## Short Answer Type Questions (SA-I)

**11.** On combustion Li forms $Li_2O$; sodium gives the peroxide, $Na_2O_2$; and potassium, rubidium and caesium give superoxides, $MO_2$. Why Li does not form a peroxide?

**12.** Complete the reactions :

(a) $Be(OH)_2 + 2OH^- \longrightarrow$

(b) $BaO + H_2O \longrightarrow$

**13.** (i) Why are alkali metal used in photo-electric cells?

(ii) What happens when K burns in air? Write equation.

**14.** What happens when :

(a) sodium metal is dropped in water?

(b) sodium metal is dissolved in liquid ammonia?

**15.** Why is $Li_2CO_3$ decomposed at a lower temperature whereas $Na_2CO_3$ at higher temperature?

**16.** What happens when :

(i) sodium peroxide dissolved in hot water?

(ii) sodium metal is heated in free supply of air?

**17.** Draw the structure of (i) $BeCl_2$ (vapour) (ii) $BeCl_2$ (solid).

**18.** (i) Name the chief factors responsible for the anomalous behaviour of lithium.

(ii) Complete the following reactions :

(a) $4LiNO_3 \xrightarrow{\Delta}$ (b) $2NaNO_3 \xrightarrow{\Delta}$

**19.** Account for the following:

(i) Magnesium does not show any flame colouration.

(ii) Group 1 elements have low melting and boiling points.

**20.** Alkaline earth metals always form divalent cations even though their second ionization enthalpies are almost double than their first ionization enthalpies. Explain.

## Short Answer Type Questions (SA-II)

**21.** 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

**22.** (a) Write balanced equations for the reactions:

(i) alkali metal and water

(ii) alkali metal and dihydrogen.

(b) The $E$ for $Cl_2/Cl^-$ is +1.36, for $I_2/I^-$ is + 0.53, for $Ag^+/Ag$ is +0.79, for $Na^+/Na$ is – 2.71 and for $Li^+/Li$ is – 3.04. Arrange the following ionic species in decreasing order of reducing strength:

$I^-$, Ag, $Cl^-$, Li, Na

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

**24.** Give reason :

(i) Alkali metals are strong reducing agents

(ii) Alkali metals are soft metals

(iii) Alkali metals tarnish in air easily.

**25.** Complete the reactions:

(i) $Na + O_2 \longrightarrow$

(ii) $BeCl_2 + LiAlH_4 \longrightarrow$

(iii) $NaNO_3 \xrightarrow{\Delta}$

**26.** Explain term Diagonal Relationship in context of Be and Al. Give two points of similarities between them.

**27.** Give reason:

(a) Sodium is less reactive than potassium.

(b) Be and Mg do not give colour to the flame whereas other alkaline earth metals do so.

(c) Sodium is stored under kerosene oil.

**28.** An element of group 2 forms covalent oxide which is amphoteric in nature and dissolves in water to give an amphoteric hydroxide. Identify the element and write chemical reactions of the hydroxide of the element with an alkali and an acid.

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

**30.** Arrange the following in order of property mentioned against each :

(i) $BeCl_2$, $MgCl_2$, $CaCl_2$, $BaCl_2$ [Increasing ionic character]

(ii) $Mg(OH)_2$, $Ca(OH)_2$, $Ba(OH)_2$, $Sr(OH)_2$ [Increasing solubility in water]

**31.** Given reason:

(i) Alkali metals impart colours to Bunsen Burner flame.

(ii) $KO_2$ is paramagnetic.

(iii) Alkaline earth metals give-blue solutions, when dissolved in liquid ammonia.

**32.** (a) Among the given alkali metals, which has least melting point? Na, K, Rb, Cs

(b) Which one of the following alkali metals gives hydrated salts? Li, Na, K, Cs

**33.** How would you explain the following observations?

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

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

**34.** Comment on the following observations :

(a) Why are alkali metals not found in nature?

(b) The mobilities of the alkali metal ions in aqueous solution are $Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$

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

## Long Answer Type Questions (LA)

**36.** Present a comparative account of the alkali and alkaline earth metals with respect to the following characteristics :

(i) Tendency to form ionic/ covalent compounds

(ii) Nature of oxides and their solubility in water

(iii) Formation of oxosalts

(iv) Solubility of oxosalts

(v) Thermal stability of oxosalts

**37.** (a) (i) What are the products formed when alkali metal oxide ($M_2O$), peroxide ($M_2O_2$) and superoxide ($MO_2$) hydrolysed by water?

(ii) Why sodium forms peroxide but potassium forms superoxide?

(b) (i) Can we store sodium in water or not? Why.

(ii) Write balanced equations for the reactions between :

(I) $Na_2O_2$ and $H_2O$

(II) $KO_2$ and $H_2O$.

**38.** (a) Explain why can alkali and alkaline earth metals not be obtained by chemical reduction methods.

(b) Explain the following :

(i) Be and Mg do not impart any colour to the flame.

(ii) What happens when Mg is burnt in air?

(c) Account for the following:

(i) Magnesium does not show any flame colouration.

(ii) Group 1 elements have low melting and boiling points.

**39.** The *s*-block elements are characterised by their larger atomic sizes, lower ionisation enthalpies, invariable +1 oxidation state and solubilities of their oxosalts. In the light of these features describe the nature of their oxides, halides and oxosalts.

**40.** (a) When a metal of group 1 was dissolved in liquid ammonia, the following observations were obtained:

(i) Blue solution was obtained initially.

(ii) On concentrating the solution, blue colour changed to bronze colour.

How do you account for the blue colour of the solution? Give the name of the product formed on keeping the solution for some time.

(b) The stability of peroxide and superoxide of alkali metals increase as we go down the group. Explain giving reason.

# ANSWERS

## OBJECTIVE TYPE QUESTIONS

**1. (a) :** Among alkali metal ions, $Li^+$ is the smallest. Therefore it has highest charge density and hence attracts water molecules more strongly than any other alkali metal cations.

**2. (d) :** In a group thermal stability of carbonates increases down the group from Li to Cs as the ionic radius of the cation increases and its polarising power decreases.

$Cs_2CO_3 > Rb_2CO_3 > K_2CO_3 > Na_2CO_3 > Li_2CO_3$

**3. (c) :** Lithium has highest hydration enthalpy which accounts for its high negative $E^\circ$ value and its high reducing power.

**4. (d) :** Li is strongest reducing agent in aqueous solution because the standard electrode potential of lithium is the lowest, *i.e.*, –3.04 volts.

**5. (d) :** Alkali metal hydrides are ionic in character.

**6. (c) :** All alkali metals are soft and light.

**7. (b) :** LiF is least soluble in water due to its high lattice enthalpy.

**8. (d) :** The carbonates of both the metals are unstable.

**9. (a)**

**10. (a) :** $Be(OH)_2$ reacts with NaOH to give beryllate ion becoming soluble in it.

$$Be(OH)_2 + 2OH^- \longrightarrow [Be(OH)_4]^{2-}$$

**11. (a) :** (i) → (C,D), (ii) → (B), (iii) → (D), (iv) → (A, E)

**12. (d) :** Higher the lattice energy lower the solubility. Out of the four combinations possible, the lattice energy of MgS is higher than those of $Na_2S$, $MgCl_2$ and NaCl and hence MgS is the least soluble.

**13. (c) :** Smaller the cation, greater is the degree of its hydration. As the degree of hydration decreases from $Li^+$ to $Cs^+$, the hydration energy of alkali metals also decreases from $Li^+$ to $Cs^+$.

**14. (d) :** Cs melts at a temperature of 30°C since melting point decreases down the group.

**15. (c) :** $5Mg + \underbrace{O_2 + N_2}_{\text{Air}} \longrightarrow \underset{(X)}{MgO} + \underset{(Y)}{Mg_3N_2}$

$Mg_3N_2 + 6H_2O \longrightarrow 3Mg(OH)_2 + \underset{(Z)}{2NH_3}$

$NH_3 + H_2O \longrightarrow \underset{\text{(Solution)}}{NH_4OH}$

$CuSO_4 + 4NH_4OH \longrightarrow \underset{(A)}{[Cu(NH_3)_4]\,SO_4} + 4H_2O$

**16. (b)**

**17. (b) :** The basic strength of the oxides increases gradually forms BeO to BaO.

$$\underset{\text{Amphoteric}}{BeO} < \underset{\text{Weakly basic}}{MgO} < \underset{\text{Basic}}{CaO} < \underbrace{SrO < BaO}_{\text{Strongly basic}}$$

**18. (d) :** (A) → (iv), (B) → (i), (C) → (ii), (D) → (iii)

**19. (c)**

**20. (b) :** Statements (3) and (4) are wrong.

$MgCO_3$ is prepared by adding sodium bicarbonate to a hot solution of magnesium salt like $MgSO_4$.

$MgSO_4 + 2NaHCO_3 \rightarrow MgCO_3 + Na_2SO_4 + H_2O + CO_2$

Mg salts do not impart any colour to the flame as due to very small size, the electrons are held tightly and hence their excitation is difficult.

Statements (1) and (2) are correct.

Oxidation potential of Ba is 2.92 and that of Mg is 2.36 *i.e.* Ba is a better reducing agent than Mg.

$Be(OH)_2$ is amphoteric in nature but $Ba(OH)_2$ is basic as basicity of hydroxides increases with increasing electropositive character.

**21. (d) :** $BeF_2$ :

Hydration energy = – 2494 + 2 × (– 457) = – 3408

Lattice energy = – 2906

$\Delta H$ solution = – 3408 – (– 2906) = – 502 kJ $mol^{-1}$

$MgF_2$ :

Hydration energy = – 1921 + 2 × (– 457) = –2835

Lattice energy = – 2610

$\Delta H_{solution}$ = – 2835 – (– 2610) = – 225 kJ $mol^{-1}$

$CaF_2$ :

Hydration energy = – 1577 + 2 × (– 457) = –2491

Lattice energy = – 2459

$\Delta H_{solution}$ = – 2491 – (– 2459) = – 32 kJ $mol^{-1}$

$BaF_2$ :

Hydration energy = – 1305 + 2 × (– 457) = –2219

Lattice energy = – 2367

$\Delta H_{solution}$ = – 2219 – (– 2367) = + 148 kJ $mol^{-1}$

$\Delta H_{solution}$ is +ve in $BaF_2$, hence $BaF_2$ is easily soluble in water.

**22. (b) :** $A \xrightarrow{\text{burns in nitrogen}} \underset{\text{ionic compound}}{B} \xrightarrow{H_2O} C + D$, $C \xrightarrow{+CO_2}$ milky solution

Since the solution of *C* becomes milky on bubbling $CO_2$ into it, *C* is lime water ($Ca(OH)_2$).Thus, the element *A* is calcium and *B* is the compound of calcium with nitrogen. Nitrides on reaction with water give the corresponding hydroxide and ammonia. Therefore, *A* is calcium (Ca). *B* is calcium nitride ($Ca_3N_2$). *C* is lime water $Ca(OH)_2$. *D* is ammonia ($NH_3$).

**23. (c) :** Polarizing power $\propto \frac{\text{charge on ion}}{\text{size of cation}}$

**24. (a) :** $Mg(OH)_2$ is least basic since basic character of hydroxides increases from $Mg(OH)_2$ to $Ba(OH)_2$.

**25. (d) :** Both $BeCl_2$ and $AlCl_3$ are strong Lewis acids.

**26. (a) :** $BeCl_2$ is a covalent compound due to small size of $Be^{2+}$ (Fajan's rule) hence, soluble in ethanol.

**27. (b) :** The hydroxides of alkaline earth metals are less soluble. The solubility increases with increase in size of the metal and this is due to decrease in their lattice energy and increase in hydration energy.

**28. (b) :** As the reduction potential values decreases down a group, the reducing character increases down the group.

**29. (c) :** Density gradually increases on moving down from Li to Cs. Potassium is, however, lighter than sodium. The abnormal value of potassium is due to unusual increase in atomic size, *i.e.*, atomic volume.

**30. (a) :** 1st ionisation energy of alkaline earth metals is higher than those of alkali metals due to increased nuclear charge.

**31. (b)** **32. (b)**

**33. (c) :** Magnesium is a strong reducing agent.

**34. (d) :** Lithium does not form solid bicarbonate though it does exist in solution. Bicarbonates of alkaline earth metals are stable only in solution and have not isolated in the pure state.

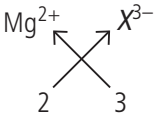
**35. (c) :** Pb and Ba both give yellow ppt. with $K_2CrO_4$ in acetic acid and white precipitate with dil. $H_2SO_4$. But lead salts give precipitate with sodium chloride or iodide also whereas barium does not. It means given salt is barium salt and with sodium carbonate it gives white precipitate of barium carbonate.

**36. (a) :** Be exhibits maximum covalency of four whereas Al shows maximum covalency of six.

**37. (d) :** Halides of alkaline earth metals are hygroscopic in nature, act as dehydrating agent and can absorb moisture from air.

**38. (d) :** Electronic configuration of $X$ is $1s^2, 2s^2, 2p^3$. So, valency of $X$ will be 3.

Magnesium ion = $Mg^{2+}$




Formula : $Mg_3X_2$

**39. (b) :** As we move down the group from $BeSO_4$ to $BaSO_4$, the enthalpy of hydration of the cations become smaller due to increase in ionic size. Salts of heavier metal ions are less soluble than those of lighter ions.

**40. (d) :** Hydration energy for $Mg^{2+}$ > hydration energy for $Mg^+$, hence Mg forms dipositive ion.

**41. (a) :** Alkali metals are good conductors of heat and electricity but have low melting points and ionisation potentials.

**42. (d) :** The metal atom loses electron and combines with ammonia molecule.

$$M + (x + y)\,NH_3 \rightarrow [M(NH_3)_x]^+ + \underset{\text{Solvated electron}}{[e(NH_3)_y]^-}$$

These ammoniated electrons are responsible for blue colour and conductivity of the solution.

**43. (b) :** In electrochemical series based on oxidation potential alkali metals are far below hydrogen, thus they displace hydrogen from water.

**44. (d) :** Basic strength of alkali metal hydroxides increases down the group. Hence $Ba(OH)_2$ is more basic than $Be(OH)_2$.

**45. (c) :** $Ca(OH)_2$ is used as flocculant.

$$\underset{(A)}{Ca(OH)_2} + Na_2CO_3 \longrightarrow \underset{\text{Caustic soda}}{2NaOH} + CaCO_3\downarrow$$

$$\underset{(A)}{Ca(OH)_2} + CO_2 \longrightarrow \underset{\text{Cloudy}}{CaCO_3\downarrow} + H_2O$$

**46. (a) :** $MgCl_2 + 2H_2O \longrightarrow MgO + 2HCl + H_2O$

**47. (c) :** Reaction can be summarized as :

$$M + N_2 \rightarrow \underset{(A)}{M_3N}$$

$$M_3N \xrightarrow{\Delta} M$$

$$M_3N + H_2O \longrightarrow B \xrightarrow[\text{Intensify its colour}]{\text{Passed through } CuSO_4}$$

$$\underset{(M)}{6Li} + N_{2(g)} \rightarrow \underset{(A)}{2Li_3N}$$

$$\underset{(A)}{Li_3N} + 3H_2O \rightarrow 3LiOH + \underset{(B)}{NH_3}$$

$$NH_3 + H_2O \rightarrow NH_4OH$$

$$4NH_4OH + CuSO_4 \rightarrow \underset{\text{Blue coloured solution}}{Cu(NH_3)_4SO_4} + 4H_2O$$

**48. (d) :** $M + (x + y)NH_3 \rightarrow [M(NH_3)_x]^+ + [e(NH_3)_y]^-$

The deep blue colour of the solution is due to the ammoniated electrons which absorbs energy in the visible region of light and thus imparts blue colour to the solution.

**49. (a) :** Density gradually increases in moving down from Li to Cs. K is, however, lighter than Na.

Densities of alkali metals are as follows :

| **Element** | Li | Na | K | Rb | Cs |
|---|---|---|---|---|---|
| **Density (g/mL)** | 0.54 | 0.97 | 0.86 | 1.53 | 1.90 |

*i.e.*, Li < K < Na < Rb < Cs

**50. (c)** **51. (a)**

**52. (c) :** Only K, Rb and Cs from alkali metals form superoxides and superoxides possess three electron bond $(:\ddot{O} \therefore \ddot{O}:)^-$.

**53. (a)**

**54. (c) :** $Li^+$ has maximum degree of hydration due to its small size and for this reason lithium salts are mostly hydrated.

**55. (a)**

**56. (a) :** The loosely held *s*-electron in the outermost valence shell of these elements makes them the most electropositive metals which readily give ions, $M^+$ or $M^{2+}$.

**57. (d) :** The ionization enthalpies of alkali metals decrease progressively as we move down the group from Li to Cs.

**58. (b) :** The solubility of the alkaline earth metal hydroxides in $H_2O$ increases with increase in atomic number down the group. This is due to the reason that both lattice energy and hydration energy decrease down the group as the size of the cation increases but lattice energy decreases more rapidly than the hydration energy and hence, their solubility increases down the group.

**59. (a)** **60. (a)**

## SUBJECTIVE TYPE QUESTIONS

**1.** Sodium has higher melting point than potassium because of stronger metallic bonding.

**2.** The ionization enthalpy, is the energy required to completely remove an electron from an isolated gaseous atom or ion. The closer and more tightly an electron is bound to its nucleus, the more difficult it will be to remove the electron and higher will be the ionization enthalpy. Group 1 elements have low first ionization enthalpies because the loss of an electron takes place readily to form a stable octet.

**3.** This is because the smaller atomic size causes the electrons to be packed more closely, thereby forming strong metallic bonds.

**4.** Due to small size and high ionisation enthalpy, Be forms coordination compound.

**5.** Let oxidation state of Na = $x$

The oxidation state of O = −1 (present as peroxide)

$\therefore \quad 2x + 2(-1) = 0 \implies 2x - 2 = 0$

$\implies \quad 2x = 2 \implies x = +1$

$\therefore$ Oxidation state of Na in $Na_2O_2$ is +1.

**6.** This is because halides of Be are covalent while those of Ba are ionic.

**7.** Lithium nitrate on heating decomposes to give $Li_2O$, $NO_2$ and $O_2$

$$4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2$$

**8.** 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 $LiCl \cdot 2H_2O$ but sodium chloride as NaCl.

**9.** Calcium is used in photoelectric cell due to low value of ionization potential.

**10.** Solvay process cannot be extended to the manufacture of $K_2CO_3$ because $KHCO_3$ is too soluble to be precipitated by the addition of ammonium hydrogen carbonate to a saturated solution of potassium chloride.

**11.** $M \xrightarrow{O_2}$ $M_2O$ (Normal oxide ($Li_2O$)) $\xrightarrow{O_2}$ $M_2O_2$ (Peroxide ($Na_2O_2$)) $\xrightarrow{O_2}$ $MO_2$ (Super oxide ($KO_2$, $RbO_2$, $CsO_2$))

The normal oxides contain $O^{2-}$ ion, the peroxides contain $O_2^{2-}$ ion and superoxides contain $O_2^-$ ion. The peroxides and superoxides become more stable with increase in atomic number of the alkali metal. The formation and stability of these oxides can be explained on the basis of lattice energy effects. $Li^+$ ion being a small ion has a strong positive field around it and can stabilise only a small anion, $O^{2-}$ whereas $Na^+$ being a large cation can stabilise a large ion and so on.

$O^-$ (Oxide ion) $\quad [— O — O —]^{2-}$ (Peroxide ion) $\quad [O — O]^-$ (Superoxide)

**12.** (a) $Be(OH)_2 + 2OH^- \longrightarrow [Be(OH)_4]^{2-}$ (Beryllate ion)

(b) $BaO + H_2O \longrightarrow Ba(OH)_2$

**13.** (i) Alkali metals are used in photoelectric cells and because of their low ionization energies they lose electrons very easily on irradiation.

(ii) Potassium reacts with oxygen to form superoxide.

$K + O_2 \rightarrow KO_2$

**14.** (a) Sodium hydroxide and hydrogen gas will be formed which will catch fire.

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

(b) Sodium metal dissolves in liquid ammonia and gives deep blue solution which is conducting in nature.

$$Na + (x + y)\ NH_3 \rightarrow [Na(NH_3)_x]^+ + [e(NH_3)_y]^-$$

The blue colour of the solution is due to the ammoniated electron which absorbs energy in the visible region of light and thus imparts blue colour to the solution.

**15.** $Li_2CO_3$ decomposes on heating because the $Li^+$ ion exerts a strong polarising action and distorts the electron cloud of the nearby oxygen atom of the large $CO_3^{2-}$ ion. This results in the weakening of the C—O bond and strengthening of the Li—O bond. This ultimately facilitates the decomposition

of $Li_2CO_3$ into $Li_2O$ and $CO_2$. The lattice energy of $Li_2O$ is higher than the lattice energy of $Li_2CO_3$. This also favours decomposition of $Li_2CO_3$. $Na^+$ ion being bigger in size, the lattice energy of $Na_2O$ is less stable than that of $Na_2CO_3$. Therefore, $Na_2CO_3$ does not decompose on heating.

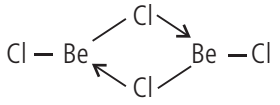
**16.** (i) Sodium hydroxide and hydrogen peroxide will be formed.

$$Na_2O_2 + H_2O_{(l)} \longrightarrow 2NaOH_{(aq)} + H_2O_2$$

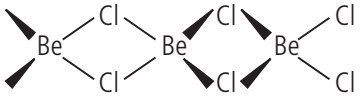
(ii) Sodium peroxide is formed.

$$2Na + O_2 \longrightarrow Na_2O_2$$

**17.** (i) $BeCl_2$ exists as a dimer in vapour phase, which dissolves into the linear monomer Cl–Be–Cl at 1200 K.




(ii) $BeCl_2$(solid): Beryllium chloride has a chain structure in solid state :




**18.** (i) Chief factors responsible for the anomalous behaviour of lithium are :

(I) its very small size (II) high electronegativity

(III) high ionization enthalpy and

(IV) absence of vacant *d*-atomic orbital in the valence shell.

(ii) (a) $4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2$

(b) $2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2$

**19.** (i) In Mg, the electrons are tightly held and excitation by mere flame is rather difficult, thus, Mg does not show flame colouration.

(ii) As group 1 elements have large size, the intermetallic bonds in them are quite weak. Hence they have low melting and boiling points.

**20.** In the solid state, higher enthalpy of lattice formation by $M^{2+}$ ions (as compared to $M^+$ ions) more than compensates the higher second ionization enthalpies of metals and in aqueous solution, higher enthalpy of hydration of $M^{2+}$ ions (as compared to $M^+$ ions) more than compensate at the higher second ionization enthalpy of metals.

**21.** (a) **Nitrates :** Nitrates of alkali metals are of the type $MNO_3$, they are soluble in water and do not undergo hydrolysis. Except $LiNO_3$, other nitrates of this group decompose to nitrites and oxygen upon heating.

$$2MNO_3 \longrightarrow 2MNO_2 + O_2\uparrow$$

Nitrates of alkaline earth metals are of the type $M(NO_3)_2$. They are soluble in water and upon heating they decompose into their corresponding oxides with evolution of a mixture of $NO_2$ and oxygen.

$$2M(NO_3)_2 \xrightarrow{\Delta} 2MO + 4NO_2 + O_2\uparrow$$

(b) **Carbonates :** All alkali metals form carbonates of the type $M_2CO_3$. Their carbonates are highly stable towards heat and readily soluble in water. The stability and solubility of the carbonates increases in the same order:

$$Cs_2CO_3 > Rb_2CO_3 > K_2CO_3 > Na_2CO_3 > Li_2CO_3.$$

$Li_2CO_3$ decomposes on heating and is insoluble in water. Carbonates of alkaline earth metals ($MCO_3$) are insoluble in water, they dissolve in water only in presence of $CO_2$ due to the formation of a hydrogen carbonate.

$$MCO_3 + H_2O + CO_2 \longrightarrow M(HCO_3)_2$$

Solubility of carbonates decreases as we descend the group and stability increases due to increase in electropositive character of the metal.

(c) **Sulphates :** Alkali metal sulphates are of the type $M_2SO_4$. Except $Li_2SO_4$, all other sulphates are soluble in water. Alkaline earth metal sulphates are of the type $MSO_4$. The solubility of sulphates decreases on moving down the group. $CaSO_4$ is sparingly soluble, $SrSO_4$, $BaSO_4$ and $RaSO_4$ are almost insoluble. The solubilities of $BeSO_4$ and $MgSO_4$ are due to high energy of solvation of smaller $Be^{2+}$ and $Mg^{2+}$ ions. The order of solubility is :

$$BaSO_4 < SrSO_4 < CaSO_4 < MgSO_4 < BeSO_4.$$

The sulphates decompose on heating to give the corresponding oxide ($MO$).

$$2MSO_4 \xrightarrow{\Delta} 2MO + 2SO_2 + O_2$$

The stability increases as the basic nature of the metal increases.

$$SrSO_4 > CaSO_4 > MgSO_4 > BeSO_4.$$

**22.** (a) (i) $2M + 2H_2O \rightarrow 2M^+ + 2OH^- + H_2$

(ii) $2M + H_2 \longrightarrow 2M^+H^-$

(b) The order is : $Li > Na > I^- > Ag > Cl^-$

**23.** (i) **Ionisation Enthalpy :** The ionisation energies of alkaline earth elements are higher than those of alkali metals due to higher nuclear charge and smaller radii.

(ii) **Basicity of oxides :** Oxides of alkali metals are stronger bases as compared to those of alkaline earth metals present in the same period. *e.g.*, when $Na_2O$ is dissolved in water, NaOH formed is a stronger base than when MgO is dissolved in water to form $Mg(OH)_2$.This is due to higher ionisation energies of alkaline earth metals.

(iii) **Solubility of hydroxides :** Alkali metal hydroxides are more soluble in water as compared to the hydroxides of alkaline earth metals present in the same period. This is due to higher lattice energy of the hydroxides of alkaline earth elements as compared to those of alkali metals.

**24.** (i) The alkali metals have only one electron in their valence shell which they lose easily, owing to their low ionization enthalpies, hence they are strong reducing agents.

(ii) Due to the presence of weak metallic bonding, alkali metals are soft and can be cut with a knife.

(iii) Lithium, sodium and potassium are all soft metals that are easily cut with a knife. The freshly cut surface is a shiny, silver colour, but this tarnishes quickly to a dull grey as the metal reacts with oxygen and water present in the air.

**25.** (i) $2Na + O_2 \longrightarrow 2Na_2O_2$

(ii) $2BeCl_2 + LiAlH_4 \longrightarrow 2BeH_2 + LiCl + AlCl_3$

(iii) $2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2$

**26.** The charge/radius ratio of $Be^{2+}$ is nearly the same as that of $Al^{3+}$ ion. Hence, beryllium resembles aluminium. This is known as diagonal relationship.

(i) Like aluminium, beryllium is not readily attacked by acids because of the presence of an oxide film on the surface of the metal.

(ii) Beryllium hydroxide dissolves in excess of alkali to give a beryllate ion, $[Be(OH)_4]^{2-}$ just as aluminium hydroxide gives aluminate ion, $[Al(OH)_4]^-$.

**27.** (a) This is mainly due to high ionisation enthalpy of sodium as compared to potassium. Therefore, potassium is more electropositive, reactive and a stronger reducing agent than sodium.

(b) In Be and Mg, the electrons are tightly held and excitation by mere flame is rather difficult, thus, they do not show flame colouration.

(c) Sodium is stored in kerosene oil because in air, sodium is easily oxidised to oxide which may dissolve in the moisture to form hydroxide.

**28.** The element is beryllium.

$$2Be + O_2 \xrightarrow{\Delta} \underset{\text{Amphoteric}}{2BeO}$$

$$BeO + H_2O \longrightarrow Be(OH)_2$$

$$Be(OH)_2 + 2OH^- \longrightarrow \underset{\text{Beryllate ion}}{[Be(OH)_4]^{2-}}$$

$$Be(OH)_2 + 2HCl \longrightarrow BeCl_2 + 2H_2O$$

**29.** (a) Both Li and Mg are harder and lighter than the other metals in their respective groups.

(b) Li like Mg decomposes water slowly to liberate hydrogen.

$2Li + 2H_2O \longrightarrow 2LiOH + H_2\uparrow$

$Mg + 2H_2O \longrightarrow Mg(OH)_2 + H_2\uparrow$

The hydroxides so formed are weak bases and decompose on heating.

$2LiOH \longrightarrow Li_2O + H_2O$;

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

(c) Both form nitrides upon direct combination with nitrogen gas.

$6Li + N_2 \longrightarrow 2Li_3N$; $3Mg + N_2 \longrightarrow Mg_3N_2$

(d) Both form monoxides when heated in oxygen.

$4Li + O_2 \longrightarrow 2Li_2O$; $2Mg + O_2 \longrightarrow 2MgO$

The oxides do not combine with excess oxygen to give superoxides.

(e) Carbonates of Li and Mg decompose on heating.

$Li_2CO_3 \longrightarrow Li_2O + CO_2$; $MgCO_3 \longrightarrow MgO + CO_2$

(f) Both LiCl and $MgCl_2$ are soluble in ethanol, they are deliquescent and crystallise from aqueous solution as hydrates, $LiCl.2H_2O$ and $MgCl_2.6H_2O$.

**30.** (i) $BeCl_2 < MgCl_2 < CaCl_2 < BaCl_2$

(ii) $Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$

**31.** (i) Alkali metals and their salts impart colour to flame. It is because their loosely held valence electrons get excited to higher energy level. When they return back they release visible light of characteristic colour to flame.

(ii) In $KO_2$, superoxide ion $O_2^-$ is present. $O_2^-$ ion has one unpaired electron in its antibonding molecular orbital. Hence, it is paramagnetic in nature.

(iii) Like alkali metals the alkaline earth metals dissolve in liquid ammonia to give deep blue black solution forming ammoniated ions.

$M + (x + y)NH_3 \rightarrow [M(NH_3)_x]^{2+} + 2[e(NH_3)_y]^-$

**32.** (a) Atomic size increases as we move down the group. As a result, the strength of metallic bonding decreases on moving down a group in the periodic table. This causes a decrease in the melting point. Among the given metals, the size of Cs is the largest and thus it has the least melting point.

(b) Smaller the size of an ion, the more highly it is hydrated. Among alkali metal ions, $Li^+$ ion is smallest in size. Also , it has the highest charge density and highest polarizing power. Hence, it attracts water molecules more strongly than the other alkali metal ions. As a result, it forms hydrated salts such as $LiCl{\cdot}2H_2O$. The other alkali metal ions are larger in size than $Li^+$ and have weaker charge densities. Hence, they usually do not form hydrated salts.

**33.** (i) BeO is almost insoluble in water as it is covalent in nature and tightly held together in the solid state while $BeSO_4$ is highly soluble in water due to high energy of solvation of smaller $Be^{2+}$ ion.

(ii) LiI having much more covalent character than KI because of small size of $Li^+$ as compared to $K^+$. Hence, LiI is more soluble in ethanol.

**34.** (a) All the alkali metals have one valence electron, $ns^1$, outside the noble gas core. The loosely held *s*-electron in the outermost valence shell of these elements makes them the most electropositive metals. They readily lose an electron to give monovalent $M^+$ ions. Thus, due to the reason cited above, alkali metals are never found free in nature but are always found in combined state.

(b) This is attributed to the hydration of the cation in water. As a result, size of the cation increases and its mobility decreases. Due to the smallest size, $Li^+$ ion is hydrated to the maximum and has least mobility while $Cs^+$ ion due to least hydration has maximum mobility.

**35.** (i) **Electronic configuration :** The valence electronic configuration of atoms of the group II A elements is $ns^2$, where '$n$' is the period number.

(ii) **Atomic and ionic sizes :** The size of the atom increases gradually from Be to Ra. Their ions are also large and size of the ion increases from $Be^{2+}$ to $Ra^{2+}$.

(iii) **Ionisation enthalpy :** The $1^{st}$ and $2^{nd}$ ionisation energies of these metals decrease from Be to Ba as size increases.

**36.** (i) Tendency to form ionic/covalent compounds :

(a) All common compounds of alkali metals are generally ionic in nature. Halides and oxides are ionic in nature with exception of lithium compounds which are generally covalent in nature.

(b) Alkaline earth metals form ionic oxides and halides except Be which forms covalent compounds.

(ii) Nature of oxides and their solubility in water :

(a) Alkali metals form oxides, peroxides and superoxides. These oxides are basic in nature and basic character increases down the group. Oxides dissolve in water to give hydroxides. These hydroxides are strong bases.

(b) Alkaline earth metal oxides are basic in nature but less basic than alkali metal oxides. BeO is amphoteric while other oxides are basic and form sparingly soluble hydroxides.

(iii) Formation of oxosalts :

(a) Alkali metals form sulphates, carbonates and bicarbonates. (b) Alkaline earth metals form sulphates, carbonates and nitrates.

(iv) Solubility of oxosalts : (a) Alkali metal oxosalts are generally soluble in water. Solubility increases down the group. (b) Carbonates and sulphates of alkaline earth metals become insoluble as we move down the group.

(v) Thermal stability of oxosalts : (a) Alkali metal oxosalts are thermally stable and stability increases down the group. $Li_2CO_3$ and $Li_2SO_4$ decompose on heating. (b) Alkaline earth metals carbonates decompose on heating while their thermal stability increases down the group.

**37.** (a) (i) Alkali metal oxide, peroxide and superoxide are easily hydrolysed by water to form the hydroxides according to the following reactions:

$M_2O + H_2O \longrightarrow 2M^+ + 2OH^-$

$M_2O_2 + 2H_2O \longrightarrow 2M^+ + 2OH^- + H_2O_2$

$2MO_2 + 2H_2O \longrightarrow 2M^+ + 2OH^- + H_2O_2 + O_2$

(ii) Sodium ion is larger in size and has comparatively weaker +ve field around it which can not prevent oxide ion to combine with another oxygen atom to form peroxide ion. While $K^+$ is larger than $Na^+$ and has still weaker +ve field which can not prevent even peroxide ion to combine with another oxygen atom to form superoxide.

(b) (i) Sodium will react with water to form sodium hydroxide and hydrogen gas which will catch fire. For this reason, sodium cannot be stored in water.

(ii) (I) $Na_2O_2 + 2H_2O \rightarrow 2NaOH + H_2O_2$

(II) $2KO_2 + 2H_2O \rightarrow 2KOH + H_2O_2 + O_2$

**38.** (a) Alkali metals cannot be extracted by the reduction of their oxides and other compounds as they are strong reducing agents themselves and no such reducing agents are there which can reduce them to get pure metal.

(b) (i) In Be and Mg, the electrons are tightly held and excitation by mere flame is rather difficult, thus, they do not show flame colouration.

(ii) Mg burns in air to form MgO and $Mg_3N_2$.

$2Mg + O_2 \longrightarrow 2MgO$

$3Mg + N_2 \longrightarrow Mg_3N_2$

(c) (i) In Mg, the electrons are tightly held and excitation by mere flame is rather difficult, thus, Mg does not show flame colouration.

(ii) As group 1 elements have large size, the intermetallic bonds in them are quite weak. Hence they have low melting and boiling points.

**39.** (i) Nature of oxides – Alkali metals form $M_2O$, $M_2O_2$ and $MO_2$ types of oxides. The stability of the peroxide or superoxide increases as the size of metal cation increases, this is due to stabilisation of large anions by larger cations.

(ii) Nature of halides – Alkali metal halides have general formula $MX$. All halides are soluble in water. LiF is very less soluble in water due to its high lattice energy. Their m.pt. and b.pt. follow the trend – fluoride > chloride > bromide > iodide due to increase in size of halide ion, the lattice energy increases.

(iii) Oxosalts – Oxosalts of alkali metals are generally soluble in water and thermally stable. As electropositive character increases down the group, stability of carbonates and bicarbonates increases.

**40.** (a) (i) Alkali metals dissolve in liquid ammonia giving deep blue solutions which are conducting in nature.

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

Blue colour is due to ammoniated electron.

(ii) On standing, hydrogen is slowly liberated and amide is formed. The colour changes to bronze. On keeping for sometime.

$$M^+ + e^- + NH_3 \longrightarrow MNH_2 + \frac{1}{2} H_2$$

(b) The increase in stability of peroxides and superoxides of alkali metals from Li to Cs can be explained by stabilisation of larger anions by larger cations through higher lattice energy.