CHAPTER

4

The *s*-Block Elements

Recap Notes

GROUP 1 ELEMENTS (ALKALI METALS)

- **General characteristics :**
  - **Electronic configuration :** Alkali metals are *s*-block elements. These contain only one electron in the *s*-orbital of the valence shell outside the noble gas core and hence their valence shell electronic configuration is ns^1.
  - **Atomic and ionic radii :** Alkali metal atoms are largest in their corresponding period in periodic table. Atomic as well as ionic size increases from Li to Fr due to the presence of an extra shell of electrons. Atomic volume (At.wt./density) also increases in moving down from Li to Cs.

Element	Atomic radii (pm)	Ionic radii (pm)	Atomic volume (mL/mol)
Li	152	76	13.0
Na	186	102	23.7
K	227	138	44.4
Rb	248	152	55.8
Cs	265	167	69.3

  - **Ionization energy :** 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.
  - **Hydration of ions :** The salts of alkali metals are ionic and soluble in water. The solubility is due to the fact that cations get hydrated by water molecules.

$$M^+ + aq. \longrightarrow \underset{\text{Hydrated cation}}{[M(aq.)]^+}$$

The smaller the cation, the greater is the degree of its hydration. Thus, the degree of hydration of M^+ ions decreases from Li^+ to Cs^+. Consequently, the radii of the hydrated ion decreases from Li^+ to Cs^+.

- **Physical properties :**
  - **Flame colouration :** Alkali metals and their salts, when introduced into the flame impart characteristic colour to the flame.

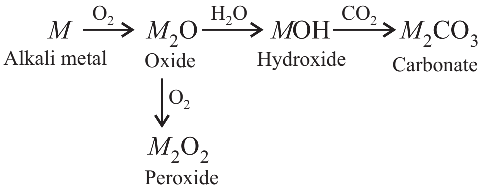
Metal	Li	Na	K	Rb	Cs
Colour	Crimson red	Yellow	Pale violet	Red violet	Blue

Atomic and physical properties of the alkali metals

Property	Lithium Li	Sodium Na	Potassium K	Rubidium Rb	Caesium Cs	Francium Fr
Atomic number	3	11	19	37	55	87
Atomic mass (g mol^{-1})	6.94	22.99	39.10	85.47	132.91	223
Electronic configuration	[He] $2s^1$	[Ne] $3s^1$	[Ar] $4s^1$	[Kr]$5s^1$	[Xe] $6s^1$	[Rn] $7s^1$

Ionization enthalpy / kJ mol^{-1}	520	496	419	403	376	~375
Hydration enthalpy/kJ mol^{-1}	–506	–406	–330	–310	–276	–
Metallic radius / pm	152	186	227	248	265	–
Ionic radius (M^+) / pm	76	102	138	152	167	180
Melting point / K	454	371	336	312	302	–
Boiling point / K	1615	1156	1032	961	944	–
Density / g cm^{-3}	0.53	0.97	0.86	1.53	1.90	–
Standard potential $E°/V$ for (M^+ / M)	–3.04	–2.714	–2.925	–2.930	–2.927	–

- **Chemical properties :**
  - **Action with air :** On exposure to moist air, all alkali metals except lithium tarnish quickly. The effect of atmosphere increases from Li to Cs. These are, therefore, always kept under kerosene oil to protect them from air.



  - Lithium forms lithium oxides, sodium forms sodium oxide and peroxide. Potassium, rubidium and caesium form oxides, peroxides and superoxides due to increase in the size of cations on moving down the group.
  - **Action of water :** The alkali metals react readily with water forming hydroxides and evolving hydrogen. For example,
    $2Li + 2H_2O \longrightarrow 2LiOH + H_2$
    - The reactivity with water increases down the group from lithium to caesium. The basic nature of hydroxides increases down the group.
  - **Action with hydrogen :** Alkali metals react with hydrogen to form ionic hydrides of type M^+H^-.
    $2M + H_2 \longrightarrow 2MH$
    - The electrolysis of the fused hydride (LiH) yields hydrogen at anode.
    - The ionic character of the bonds in these hydrides increases from LiH to CsH and their stability decreases in the same order. They are powerful reducing agents especially at high temperatures.
  - **Action with halogens :** The alkali metals combine readily with halogens forming the halides. These are ionic compounds and are represented as M^+X^-. For example,
    $2Li + Cl_2 \longrightarrow 2Li^+Cl^-$
  - **Solubility in liquid NH_3 :** The alkali metals dissolve in liquid ammonia without the evolution of hydrogen. The colour of the dilute solutions is blue. The metal atom loses electron and it combines with ammonia molecule.
    $M \longrightarrow M^+ \text{ (in liquid ammonia)} + e^-$
    $M + (x + y)\,NH_3 \longrightarrow [M(NH_3)_x]^+ + \underset{\text{(ammoniated electron)}}{[e(NH_3)_y]^-}$
    On heating its blue colour changes to bronze.
    It is ammoniated electron which is responsible for colour.

- **General characteristics of the compounds of the alkali metals :**
  - **Nature of alkali metal oxides :** Alkali metal oxides are basic in nature and their basic character increases gradually on moving down the group.
    $Li_2O + H_2O \rightarrow 2LiOH$
    $Na_2O + H_2O \rightarrow 2NaOH$

$Na_2O_2 + 2H_2O \rightarrow 2NaOH + H_2O_2$
$2K_2O_2 + 2H_2O \rightarrow 4KOH + O_2$
$2KO_2 + 2H_2O \rightarrow 2KOH + H_2O_2 + O_2$

- **Nature of alkali metal hydroxides:** Alkali metal hydroxides, form the strongest bases. The basic character of the alkali metal hydroxides increases from LiOH to CsOH.
  LiOH < NaOH < KOH < RbOH < CsOH
  This can be explained on the basis of increase in internuclear distances between the oxygen of the OH group and the metal ion, which increases with the increase in the cation size (Li^+ to Cs^+). The thermal stability, solubility, reactivity, m.pt. and b.pt. of hydroxides also increases on moving from Li to Cs.
- **Nature of alkali metal halides :** Alkali metals halides are ionic compounds except LiBr and LiI. All alkali halides are freely soluble in water except LiF. LiF is soluble in non-polar solvents. The insolubility of LiF in water is attributed to high lattice energy as Li^+ and F^- ions have small size and large charge. The CsI has also low solubility due to smaller hydration energy of its two ions.
  - Melting point and boiling point of halides of a particular alkali metal follow the order :
    Fluorides > Chlorides > Bromides > Iodides.
- **Salts of oxoacids :** Alkali metal hydroxides being basic in nature react with oxoacids (such as H_2CO_3, H_3PO_4, HNO_3, HNO_2, H_2SO_4, etc.) to form different salts such as metal carbonates, bicarbonates, sulphates, nitrates, etc.
- **Nature of carbonates and bicarbonates :** Alkali metal carbonates and bicarbonates are highly stable towards heat and their stability increases down the group, since electropositive character increases from Li to Cs. However Li_2CO_3 is less stable and readily decomposes to form oxide because Li^+ ion exerts a strong polarising action and distorts the electron cloud of the nearby oxygen atom of CO_3^{2-} ion. This result in the weakening of the C—O bond and strengthening of Li—O bond. This ultimately facilitate the decomposition of Li_2CO_3 into Li_2O and CO_2.
  $Li_2CO_3 \longrightarrow Li_2O + CO_2$
  - Alkali metal bicarbonates on heating decompose to give respective carbonates.
    $2MHCO_3 \xrightarrow{\Delta} M_2CO_3 + CO_2 + H_2O$
  - All carbonates and bicarbonates are water soluble. Their solubility and stability decreases in the following order :
    $Cs_2CO_3 > Rb_2CO_3 > K_2CO_3 > Na_2CO_3 > Li_2CO_3$
- **Anomalous behaviour of Li :** Due to very small size, comparatively high ionization enthalpy, non-availability of *d*-orbitals and high polarizing power of Li^+.

All alkali metals	Except
Do not react directly with N_2 or C.	Li which forms Li_3N or Li_2C_2
Form amide (MNH_2) with ammonia.	Li which forms Li_2NH
Nitrates are thermally stable.	$LiNO_3$
Carbonates are thermally stable.	Li_2CO_3
Form double salts (alums) from their sulphates.	Li_2SO_4
Form acetylides with acetylene.	Li

- **Diagonal relationship of Li with Mg :** Lithium resembles magnesium mainly due to the similarity in sizes of their atoms.
  - Both Li and Mg are quite hard. They are harder and have high m.pt. than other elements in their respective groups.
  - Carbonates, nitrates and hydroxides decompose on heating :
    $Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2$
    $MgCO_3 \xrightarrow{\Delta} MgO + CO_2$
    $4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2$
    $2Mg(NO_3)_2 \xrightarrow{\Delta} 2MgO + 4NO_2 + O_2$
    $2LiOH \xrightarrow{\Delta} Li_2O + H_2O$
    $Mg(OH)_2 \xrightarrow{\Delta} MgO + H_2O$

- Combine directly with N_2 :
  $6Li + N_2 \longrightarrow 2Li_3N$
  $3Mg + N_2 \longrightarrow Mg_3N_2$
- Combine with carbon on heating :
  $2Li + 2C \longrightarrow Li_2C_2$
  $Mg + 2C \longrightarrow MgC_2$
- Form monoxide on heating with oxygen :
  $4Li + O_2 \longrightarrow 2Li_2O$
  $2Mg + O_2 \longrightarrow 2MgO$
- Their periodic properties like atomic radii, ionic radii, atomic volume are quite comparable.
- $MgSO_4$ like Li_2SO_4 does not form alums.
- $LiHCO_3$ and $Mg(HCO_3)_2$ do not exist in solid state, they exist in solution only.

- **Biological importance of sodium and potassium :** Sodium and potassium are present in biological fluids. The most remarkable feature of Na^+ and K^+ ions is the development and functional features of nerve cells. In the resting state, a nerve cell shows a potential corresponding to the potassium ion concentration across the membrane. During activation of nerve cells, a chemical, acetyl-choline is released near its end plate and the membrane potential is discharged. This discharge is transmitted through the length of the nerve cell by an electric pulse. This action illustrates the importance of Na^+ and K^+ ions.

GROUP 2 ELEMENTS (ALKALINE EARTH METALS)

- **General characteristics :**
  - **Electronic configuration :** The electronic configuration of alkaline earth metals are ns^2.
  - **Atomic and ionic radii :** The atoms and ions are large but are smaller than the corresponding group 1 elements since the extra charge on the nucleus draws the valence electrons. The size of atom increases gradually from Be to Ra.
  - **Ionisation energies and electropositive character :** With increasing size, the ionization energy decreases from Be to Ra. Although the second ionization energy of these atoms to produce M^{2+} ions is greater than the first to produce M^+ ions, the extra energy to remove second electron is offset by the lattice energy in crystals or the solvation energy in solutions. The electropositive character increases from Be to Ra. The oxidation potentials corresponding to the reaction,
    $M_{(s)} \longrightarrow M^{2+} + 2e^-$
    are relatively high which indicate that in solutions these elements are good reducing agents.
  - **Hydration of ions and hydration energy :** The ions are hydrated in aqueous solutions. The extent of hydration of the ions decreases with increase in atomic number as shown by their heats of hydration (ΔH) in kJ mol^{-1}.

Be^{2+}	Mg^{2+}	Ca^{2+}	Sr^{2+}	Ba^{2+}
–2406	–1929	–1632	–1485	–1276

  - **Flame test :** In flame the electrons are excited to higher energy states, and when these drop back to the original state, extra energy is emitted in the form of visible light. This accounts for the characteristic colours given by Ca, Sr, and Ba in the flame test.
    Calcium—brick red, Strontium—crimson red
    Barium—apple green, Radium—crimson red

- **Chemical properties :**
  - **Action of air :** Alkaline earth metals being less electropositive than the alkali metals, are less reactive. These are only slowly oxidised on exposure to air. Their activity increases down the group from Be to Ra.
  - **Action of water :** With the exception of beryllium and magnesium, the alkaline earth metals react vigorously with cold water to form hydrogen and corresponding metal hydroxide.
    $Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$
    The alkaline earth metal hydroxides are less basic than the alkali metal hydroxides. Beryllium hydroxide is amphoteric in nature, whereas the basic

character of other hydroxides increases down the group.

- **Action of hydrogen :** All the alkaline earth metals except Be, combine with H_2 directly on heating to form metal hydrides of the general formula, MH_2.

  $M + H_2 \xrightarrow{\Delta} MH_2$

  (where, M = Mg, Ca, Sr or Ba)

  BeH_2 can, however be prepared by reducing $BeCl_2$ with $LiAlH_4$.

  $2BeCl_2 + LiAlH_4 \rightarrow 2BeH_2 + LiCl + AlCl_3$

  BeH_2 is covalent, MgH_2 is partly ionic but the hydrides of Ca, Sr, and Ba are ionic solids. Further, BeH_2 and MgH_2 are polymeric while others are monomeric.
- **Reaction with oxygen :** The alkaline earth metals burn in oxygen to form oxides or peroxides. Beryllium, magnesium and calcium form oxides, strontium and barium form peroxides with excess of oxygen.

  $2M + O_2 \longrightarrow 2MO$ (M = Be, Mg or Ca)

  $M + O_2 \longrightarrow MO_2$ (M = Ba or Sr)

  The reactivity of the alkaline earth metals towards oxygen increases down the group.
- **Action of halogens :** Alkaline earth metals combine with halogens at high temperatures to form halides, MX_2. Beryllium halides are covalent, hygroscopic and fumes, in air due to hydrolysis. The other halides are electrovalent and readily soluble in water.
- **Action of acids :** All of them easily dissolve in acids with the liberation of hydrogen. Beryllium differs from other members of its group in being soluble in alkalies also.

  $Be + 2NaOH + 2H_2O \longrightarrow \underset{\text{Sodium beryllate}}{Na_2[Be(OH)_4]} + H_2$
- **Action with nitrogen :** All the alkaline earth metals burn in nitrogen to form nitrides of the type M_3N_2.

  $3M + N_2 \longrightarrow M_3N_2$

  The ease of formation of nitrides decreases from Be to Ba.

- **Characteristics of compounds of the alkaline earth metal :**
  - **Nature of oxides and peroxides :** Oxides of group 2 elements are basic in nature. Their basic strength increases down the group.

    $$\underset{\text{Strongly basic}}{BaO} > \underset{\text{Strongly basic}}{SrO} > \underset{\text{Basic}}{CaO} > \underset{\text{Weakly basic}}{MgO} > \underset{\text{Amphoteric}}{BeO}$$
  - **Nature of hydroxides :** $Be(OH)_2$ is amphoteric, but the hydroxides of Mg, Ca, Sr and Ba are basic. The basic strength increases from Mg to Ba.
  - **Nature of halides :** Due to small size and high charge density, all beryllium halides are essentially covalent and are soluble in organic solvents.

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

    The halides of all other alkaline earth metals are ionic. Their ionic character, however, increases as the size of the metal ion increases.

    They are hygroscopic, and fume in air due to hydrolysis. On hydrolysis, they produce acidic solution.
  - **Carbonates and bicarbonates :** Carbonates of alkaline earth metals can be produced by passing CO_2 through their hydroxides.

    $M(OH)_{2(aq)} + CO_{2(g)} \longrightarrow MCO_{3(s)} + H_2O_{(l)}$

    Alkaline earth metal carbonates are ionic but beryllium carbonate is prone to hydrolysis. It contains the hydrated ion $[Be(H_2O)_4]^{2+}$ rather than Be^{2+}. The solubility of carbonates decreases down the group from Be to Ba.

    Bicarbonates of alkaline earth metals exist in solution only. On heating, bicarbonates decompose to carbonates with evolution of CO_2.

    $M(HCO_3)_2 \xrightarrow{\Delta} MCO_3 + CO_2 + H_2O$

    Bicarbonates of Ca and Mg cause temporary hardness of water whereas their chlorides and sulphates are responsible for permanent hardness of water.
  - **Sulphates :** The sulphates of alkaline earth metals (MSO_4) are prepared by the action of sulphuric acid on metals, metal oxides, hydroxides and carbonates.

$M + H_2SO_4 \rightarrow MSO_4 + H_2$

$MO + H_2SO_4 \rightarrow MSO_4 + H_2O$

$M(OH)_2 + H_2SO_4 \rightarrow MSO_4 + 2H_2O$

$MCO_3 + H_2SO_4 \rightarrow MSO_4 + CO_2 + H_2O$

- **Nitrates :** Alkaline earth metal nitrates are prepared in solution and can be crystallized as hydrated salts by the action of HNO_3 on oxides, hydroxides and carbonates.

  $MO + 2HNO_3 \longrightarrow M(NO_3)_2 + H_2O$

  $M(OH)_2 + 2HNO_3 \longrightarrow M(NO_3)_2 + 2H_2O$

  $MCO_3 + 2HNO_3 \longrightarrow M(NO_3)_2 + CO_2 + H_2O$

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

  Magnesium nitrate crystallizes as $Mg(NO_3)_2.6H_2O$ while $Ba(NO_3)_2$ crystallises as anhydrous salt.

- **Anomalous behaviour of Be :** It is due to small size, high ionization enthalpy, absence of *d*-orbitals.

All alkaline earth metals	**Except**
Form ionic non volatile nitrides.	Be_3N_2 is covalent and volatile.
Form ionic carbonates.	$BeCO_3$ is unstable.
Form basic oxides.	BeO is amphoteric.

- **Diagonal relationship of beryllium and aluminium :** Both have same electronegativity, polarising power and charge / radius ratio of their ions.
  - Both have tendency to form covalent compounds.
  - Both $BeCl_2$ and $AlCl_3$ have chlorine bridged structure in vapour phase.
  - Both the metals dissolve in strong alkalies.

Practice Time

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

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

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	hygroscopic	hydrates
(b)	less	decompose	soluble	deliquescent	hydrates
(c)	freely	sublime	insoluble	deliquescent	anhydrous
(d)	freely	decompose	soluble	hygroscopic	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 :

$$M \longrightarrow M^+ + e^-$$

$$M^+ + xNH_3 \longrightarrow [M(NH_3)_x]^+$$

$$e^- + yNH_3 \longrightarrow [e(NH_3)_y]^-$$

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

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 ($E°$) 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) $BeCl_2 + LiAIH_4 \longrightarrow$

(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 $E°$ 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

ΔH solution = – 3408 – (– 2906) = – 502 kJ mol^{-1}

MgF_2 :

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

Lattice energy = – 2610

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

CaF_2 :

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

Lattice energy = – 2459

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

BaF_2 :

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

Lattice energy = – 2367

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

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

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

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 $\propto \dfrac{\text{charge on ion}}{\text{size of cation}}$

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.

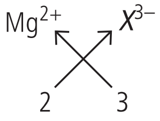
35. (c) : Pb and Ba both give yellow ppt. with K_2CrO_4 in acetic acid and white precipitate with dil. H_2SO_4. But lead salts give precipitate with sodium chloride or iodide also whereas barium does not. It means given salt is barium salt and with sodium carbonate it gives white precipitate of barium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

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

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

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

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

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

$$NH_3 + H_2O \rightarrow NH_4OH$$

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

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 $(:\ddot{O}\overset{\cdots}{—}\ddot{O}:)^{-}$.

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)

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

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

$\therefore$ 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. $$M \xrightarrow{O_2} \underset{\substack{\text{Normal oxide}\\(Li_2O)}}{M_2O} \xrightarrow{O_2} \underset{\substack{\text{Peroxide}\\(Na_2O_2)}}{M_2O_2} \xrightarrow{O_2} \underset{\substack{\text{Super oxide}\\(KO_2,\ RbO_2,\ CsO_2)}}{MO_2}$$

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.

$$\underset{\text{Oxide ion}}{O^-}\qquad \underset{\text{Peroxide ion}}{[—O—O—]^{2-}}\qquad \underset{\text{Superoxide}}{[O—O]^-}$$

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

(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.

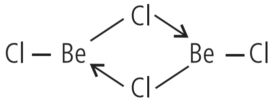
16. (i) Sodium hydroxide and hydrogen peroxide will be formed.

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

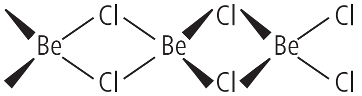
(ii) Sodium peroxide is formed.

$$2Na + O_2 \longrightarrow Na_2O_2$$

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 1st and 2nd 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.

☺☺☺