10 The s-Block Elements

Facts that Matter

• General Electronic Configuration of s-Block Elements

For alkali metals [noble gas] ns^1
For alkaline earth metals [noble gas] ns^2

• Group 1 Elements: Alkali metals

Electronic Configuration. ns^1, where n represents the valence shell.
These elements are called alkali metals because they readily dissolve in water to form soluble hydroxides, which are strongly alkaline in nature.

Element	*Symbol*	*Electronic configuration*
Lithium	Li	$1s^2 2s^1$
Sodium	Na	$1s^2 2s^2 2p^6 3s^1$
Potassium	K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Rubidium	Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$
Caesium	Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$
		$4p^6 4d^{10} 5s^2 5p^6 6s^1$ or $[Xe]\ 6s^1$
Francium	Fr	$[Rn]\ 7s^1$

• Atomic and Ionic Radii

Atomic and ionic radii of alkali metals increase on moving down the group *i.e.,* they increase in size going from Li to Cs. Alkali metals form monovalent cations by losing one valence electron. Thus cationic radius is less as compared to the parent atom.

• Ionization Enthalpy

The ionization enthalpies of the alkali metals are generally low and decrease down the group from Li to Cs.
Reason: Since alkali metals possess large atomic sizes as a result of which the valence s-electron (ns^1) can be easly removed. These values decrease down the group because of decrease in the magnitude of the force of attraction with the nucleus on account of increased atomic radii and screening effect.

• Hydration Enthalpy

Smaller the size of the ion, more is its tendency to get hydrated hence more is the hydration enthalpy.

Hydration enthalpies of alkali metal ions decrease with increase in ionic sizes.

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

• Physical Properties

(*i*) All the alkali metals are silvery white, soft and light metals.

(*ii*) They have generally low density which increases down the group.

(*iii*) They impart colour to an oxidising flame. This is because the heat from the flame excites the outermost orbital electron to a higher energy level. When the excited electron comes back to the ground state, there is emission of radiation in the visible region.

• Chemical Properties of Alkali Metals

(*i*) Reaction with air:

When exposed to air surface of the alkali metals get tarnished due the formation of oxides and hydroxides.

Alkali metals combine with oxygen upon heating to form different oxides depending upon their nature.

$$4Li + O_2 \xrightarrow{\text{Heat}} 2Li_2O$$

$$2Na + O_2 \xrightarrow{\text{Heat}} \underset{\text{Sodium peroxide}}{Na_2O_2}$$

$$K + O_2 \xrightarrow{\text{Heat}} \underset{\text{Potassium superoxide}}{KO_2}$$

↓ Reactivity with oxygen increases from Li to Cs.

(*ii*) Reaction with water:

Alkali metals react with water to form hydroxide and dihydrogen

$$2M + 2H_2O \longrightarrow 2M^+ + 2OH^- + H_2$$

(M = an alkali metal)

(*iii*) Reaction with hydrogen:

The alkali metals combine with hydrogen at about 673 K (lithium at 1073 K) to form hydrides.

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

The ionic character of hydrides increases from Li to Cs.

(*iv*) Reaction with halogens:

Alkali metals combine with halogens directly to form metal halides.

$$2M + X_2 \longrightarrow 2MX$$

They have high melting and boiling points.

Order of reactivity of M:

$$Li \xrightarrow[\text{increases}]{} Cs$$

Order of reactivity of X_2: $F_2 > Cl_2 > Br_2 > I_2$

(*v*) Reducing nature:

The alkali metals are strong reducing agents. In aqueous solution it has been observed that the reducing character of alkali metals follows the sequence Na < K < Rb < Cs < Li, Li is the strongest while sodium is least powerful reducing agent. This can be explained in terms of electrode potentials ($E°$). Since the electrode potential of Li is the lowest. Thus Li is the strongest reducing agent.

(*vi*) **Solubility in liquid ammonia:**

The alkali metals dissolve in liquid ammonia to give deep blue solution. The solution is conducting in nature.

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

When light falls on the ammoniated electrons, they absorb energy corresponding to red colour and the light which emits from it has blue colour. In concentrated solution colour changes from blue to bronze. The blue solutions are paramagnetic while the concentrated solutions are diamagnetic.

• Uses of Alkali Metals

Uses of Lithium

(*i*) Lithium is used as deoxidiser in the purification of copper and nickel.
(*ii*) Lithium is used to make both primary and secondary batteries.
(*iii*) Lithium hydride is used as source of hydrogen for meteorological purposes.
(*iv*) Lithium aluminium hydride ($LiAlH_4$) is a good reducing agent.
(*v*) Lithium carbonate is used in making glass.

Uses of Sodium

(*i*) Used as sodium amalgum in laboratory (synthesis of organic compounds).
(*ii*) Sodium is used in sodium vapour lamp.
(*iii*) In molten state, it is used in nuclear reactors.
(*iv*) An alloy of sodium-potassium is used in high temperature thermometres.

Uses of Potassium

(*i*) Salts of potassium are used in fertilizers.
(*ii*) Used as reducing agent.

Uses of Cesium

(*i*) In rocket propellent
(*ii*) In photographic cells.

• Group 2 Elements: Alkaline Earth Metals

Alkaline Earth Metals: They were named alkaline earth metals since they were alkaline in nature like alkali metals oxides and they were found in the earth's crust.
Example, Be (Beryllium), Ca, Mg, Sr etc.

• Electronic Configuration

Their general electronic configuration is represented as [noble gas] ns^2.

Element	*Symbol*	*Electronic configuration*
Beryllium	Be	$1s^2 2s^2$
Magnesium	Mg	$1s^2 2s^2 2p^6 3s^2$
Calcium	Ca	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Strontium	Sr	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$ $4s^2 4p^6 5s^2$
Barium	Ba	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$ $4p^6 4d^{10} 5s^2 5p^6 6s^2$ or $[Xe]6s^2$
Radium	Ra	$[Rn]7s^2$

• Atomic and Ionic Radii

Atomic and ionic radii of alkaline earth metals one comparatively smaller than alkali metals. Within the group atomic and ionic radii increases with the increase in atomic number.

Reason: Because these elements have only two valence electrons and the magnitude of the force of attraction with the nucleus is quite small.

• Ionization Enthalpies

These metals also have low ionization enthalpies due to fairly large size of atoms. As the atomic sizes increase down the group ionization enthalpies are expected to decrease in the same manner.

Due to their small size in comparison to alkali metals first ionization enthalpies of alkaline earth metals is higher than that of alkali metals.

• Hydration Enthalpies

The hydration enthalpies of alkaline earth metal ions are larger than those of the alkali metals. Thus alkaline earth metals have more tendency to become hydrate. The hydration enthalpies decreases down the group since the cationic size increases.

$$Be^{2+} > Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$$

Metallic character: They have strong metallic bonds as compared to the alkali metals in the same period. This is due to the smaller kernel size of alkaline earth metal and two valence electrons present in the outermost shell.

• Physical Properties

(*i*) They are harder than alkali metals.

(*ii*) M.P and B.P are higher than the corresponding alkali metals due to their small size.

(*iii*) The electropositive character increases down the group.

(*iv*) Except Be and Mg, all these metals impart characteristic colour to the flame.

(*v*) The alkaline earth metals possess high thermal and electrical conductivity.

• Chemical Properties

1. **Reaction with oxygen.** Beryllium and magnesium are kinetically inert to oxygen because of the formation of a thin film of oxide on their surface.

   Reactivity towards oxygen increases as going down the group.

2. **Reaction with water.** Since these metals are less electropositive than alkali metals, they are less reactive towards water.

   Magnesium reacts with boiling water or steam. Rest of the members reacts even with cold water.

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

3. **Reaction with halogens.** They combine with the halogens at appropriate temperature to form corresponding halides MX_2.

$$M + X_2 \longrightarrow MX_2 \ (X = F, Cl, Br, I)$$

   Thermal decomposition of $(NH_4)_2 BeF_4$ is used for the preparation of BeF_2.

4. **Reaction with hydrogen.** These metals except B*e* combine with hydrogen directly upon heating to form metal hydrides.

$$M + H_2 \xrightarrow{\text{Heat}} \underset{\text{Metal hydride}}{MH_2}$$

$$BeCl_2 + LiAlH_4 \xrightarrow{\text{Ether}} 2BeH_2 + LiCl + AlCl_3$$

• General Characteristics of Compounds of Alkaline Earth Metals

Oxides and Hydroxides

(*i*) The alkaline earth metals burn in oxygen to form MO (monoxide).
(*ii*) These oxides are very stable to heat.
(*iii*) BeO is amphoteric in nature while oxides of other elements are ionic.
(*iv*) Exept BeO, they are basic in nature and react with water to form sparingly soluble hydroxides.

$$MO + H_2O \longrightarrow M(OH)_2$$

(*v*) Hydroxides of alkaline earth metals are less stable and less basic than alkali metal hydroxides.
(*vi*) Beryllium hydroxide is amphoteric in nature.

Halides

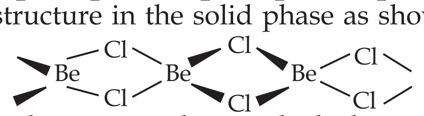
The alkaline earth metals combine directly with halogens at appropriate temperatures forming halides, MX_2.
They can also be prepared by the action of halogen acids (*HX*) on metals, metal oxides, metal hydroxides.

$$M + 2\,HX \longrightarrow MX_2 + H_2$$
$$MO + 2HX \longrightarrow MX_2 + H_2O$$
$$M\,(OH)_2 + 2HX \rightarrow MX_2 + 2H_2O$$

(*i*) Except beryllium halides, all other halides of alkaline earth metals are ionic in nature.
(*ii*) Except $BeCl_2$ and $MgCl_2$ other chloride of alkaline earth metals impart characteristic colours to flame.

$$\underset{\text{Brick red}}{CaCl_2} \qquad \underset{\text{Crimson}}{SrCl_2} \qquad \underset{\text{Grassy green}}{BaCl_2}$$

(*iii*) The tendency to form halide hydrates decreases down the group.
For example, ($MgCl_2 \cdot 8\,H_2O$, $CaCl_2 \cdot 6\,H_2O$, $SrCl_2 \cdot 6\,H_2O$, $BaCl_2 \cdot 2\,H_2O$)
(*iv*) $BeCl_2$ has a chain structure in the solid phase as shown below.



In vapour phase the compound exist as a dimer which decomposes at about 1000K to give monomer in which Be atom is in sp hybridisation state.

Sulphates

(*i*) The sulphates of alkaline earth metals are white solids and quite stable to heat.
(*ii*) $BeSO_4$ and $MgSO_4$ are readily soluble in water. Solubility decreases from $BeSO_4$ to $BaSO_4$.

Reason. Due to greater hydration enthalpies of Be^{2+} ions and Mg^{2+} ions they overcome the lattice enthalpy factor. Their sulphates are soluble in water.

Carbonates

Carbonates of alkaline earth metals are thermally unstable and decompose on heating.

$$MCO_3 \underset{}{\overset{heat}{\rightleftharpoons}} MO + CO_2$$

The thermal stability increases from $BeCO_3$ to $BaCO_3$.

$BeCO_3$ is unstable and kept only in the atmosphere of CO_2.

Solubility in water. $BeCO_3$ is highly soluble in water whereas $BaCO_3$ is almost insoluble.

Nitrates

(*i*) Nitrates of alkaline earth metals are prepared by treating the corresponding metal carbonates with dilute HNO_3

For example,

$$\underset{\substack{\text{Metal}\\\text{carbonate}}}{MCO_3} + 2HNO_3 \longrightarrow M(NO_3)_2 + H_2O + CO_2$$

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

(*ii*) All these metal nitrates decompose on heating to give the oxide

$$2M\,(NO_3)_2 \longrightarrow 2MO + 4NO_2 + O_2$$

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

• Some Important Compounds of Calcium

(*i*) Calcium Oxide (Quick Lime) CaO

Preparation: It is prepared by heating limestone in a rotery kiln at 1070 – 1270 K.

$$CaCO_3 \overset{heat}{\rightleftharpoons} CaO + CO_2$$

Since the reaction is reversible, the CO_2 is removed as soon as it is formed to enable the reaction to proceed to completion.

Properties:

(*i*) It is a white amorphous solid. M.P = 2870 K.

(*ii*) It reacts with water to become slacked lime. The reaction is highly exothermic and produce hissing sound.

$$CaO + H_2O \longrightarrow Ca\,(OH)_2$$

(*iii*) On exposure to atmosphere, it absorbs moisture and carbon dioxide.

$$CaO + H_2O \longrightarrow Ca\,(OH)_2$$
$$CaO + CO_2 \longrightarrow CaCO_3$$

(*iv*) At high temperature it combines with acidic oxides.

$$CaO + SiO_2 \longrightarrow CaSiO_3$$
$$6\,CaO + P_4O_{10} \longrightarrow 2\,Ca_3\,(PO_4)_2$$

Uses:

(*i*) In the manufacture of cement, sodium carbonate, calcium carbide etc.

(*ii*) Used in the purification of sugar.

(*iii*) In the manufacture of dye stuffs.

(*ii*) Calcium Hydroxide (slacked lime), $Ca(OH)_2$

Preparation: It is prepared by adding water to quick lime, CaO.

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Properties:

(*i*) It is a white amorphous powder.

(*ii*) When it is passed through dry Cl_2 bleaching powder is formed.

$$Ca(OH)_2 + Cl_2 \longrightarrow CaOCl_2 + H_2O$$

(*iii*) When it is treated with CO_2 the solution becomes milky due to the formation of calcium carbonate.

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

Uses:

(*i*) It is used in the manufactuning of building material.

(*ii*) Used in white-wash as a disinfectant.

(*iii*) Used to detect CO_2 gas in the laboratory.

(*iii*) Calcium Carbonate or Limestone ($CaCO_3$)

Preparation: Calcium carbonate occurs in nature in different forms like limestone, marble, chalk etc. It can be prepared by passing CO_2 through slaked lime in limited amount.

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

It can also prepared by the reaction of a solution of sodium carbonate with calcium chloride.

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

Properties:

(*i*) It is a white fluffy powder and is sparingly soluble in water.

(*ii*) Upon heating to 1200 K, it decomposes to evolve carbon dioxide.

$$CaCO_3 \xrightarrow{1200\text{ K}} CaO + CO_2$$

(*iii*) It reacts with dilute acids to form corresponding chloride, sulphate, water and CO_2 gas is evolved.

$$CaCO_3 + 2\,HCl \longrightarrow CaCl_2 + H_2O + CO_2$$
$$CaCO_3 + H_2SO_4 \longrightarrow CaSO_4 + H_2O + CO_2$$

Uses:

(*i*) In the manufacturing of Quick Lime.

(*ii*) With $MgCO_3$ used as flux in the extraction of metals.

(*iii*) Used as an antacid.

(*iv*) In the manufacture of high quality paper.

(*iv*) Calcium Sulphate (Plaster of Paris) $CaSO_4\cdot 1/2H_2O$

Preparation: It is obtained when gypsum $CaSO_4\cdot 2\,H_2O$ is heated to 393 K

$$2(CaSO_4\cdot 2H_2O) \longrightarrow 2(CaSO_4)\cdot H_2O + 3H_2O$$

Above 393 K anhydrous $CaSO_4$ is formed, which is called 'dead burnt plaster'.

Properties:

(*i*) It is a white atmosphous powder.

(*ii*) When it is mixed in adequate quantity of water it forms a plastic hard mass within 15 minutes.

Uses:

(*i*) Commonly used in making pottery, ceramics etc.

(*ii*) Used in the surgical bandages for setting the fractured bone or sprain.

(*iii*) For making statues, ornamental work, decorative material etc.

(*v*) Cement

Preparation: Prepared by combining a material rich in CaO with other material such as clay, which contains SiO_2 along with the oxides of aluminium, iron and magnesium.

$$\text{Clay + limestone} \xrightarrow[\text{strong heating}]{\Delta} \underset{\text{clinker}}{\text{Cement}}$$

$$\text{Cement clinker + 2 to 3\% gypsum} \downarrow \text{Cement}$$

Important Ingredients of portland cement:
(Ca_2SiO_4) dicalcium silicate 26%
(Ca_3SiO_5) Tricalcium silicate 51%
($Ca_3Al_2O_6$) Tricalcium Aluminate 11%
Uses:
In plastering and in construction purposes.

Words that Matter

- s-block elements constitute Group I and II elements.
- General electronic configuration of
  Group I = [Noble gas] ns^1
  Group II = [Noble gas] ns^2
- **Diagonal Relationship**

  Second period Li Be B C

  Third period Na Mg Al Si

  The first three elements of second period (Li, Be, B) show diagonal similarity with the elements (Mg, Al, Si) of third period. Such similarities are termed as diagonal relationship.
- The alkali metals are silvery-white soft metals. They are highly reactive. Their aqueous solutions are strongly alkaline in nature. Their atomic and ionic sizes increase on moving down the group and ionization enthalpies decrease systematically down the group.
- **Alkaline earth metals.** They are much similar to alkali metals but due to small size some differences are there. Their oxides and hydroxides are less basic than the alkali metals.
- Sodium hydroxide (NaOH) is prepared by the electrolysis of aq NaCl in Castner-Kellner cell.
- Slaked lime $Ca(OH)_2$ is formed by the action of quick lime on water.
- Gypsum is $CaSO_4 \cdot 2\ H_2O$. On heating upto 390 K $CaSO_4 1/2 H_2O$ (plaster of paris) is formed.

NCERT TEXTBOOK QUESTIONS SOLVED

Q1. *What are the common physical and chemical features of alkali metals?*

Ans. **Physical properties of alkali metals**

(*i*) Alkali metals have low ionization enthalpies.
(*ii*) Alkali metals are highly electropositive in nature.
(*iii*) Alkali metals exhibit +1 oxidation states in their compounds.
(*iv*) Alkali metals impart characteristic colours to the flame.

Chemical properties of alkali metals:

(*i*) Alkali metals are highly reactive in nature.
(*ii*) Alkali metals hydroxides are highly basic in nature.
(*iii*) Alkali metals dissolve in liquid ammonia to form blue and conducting solution.

Q2. *Discuss the general characteristics and gradation in properties of alkaline earth metals.*

Ans. (*i*) Atomic size goes on increasing down the group.

(*ii*) Ionisation energy goes on decreasing down the group.
(*iii*) They are harder than alkali metals.
(*iv*) They are less electropositive than alkali metals.
(*v*) Electropositive character increases on going down the group.

Q3. *Why are alkali metals not found in nature?*

Ans. Alkali metals are highly reactive in nature. That's why they always exist in combined state in nature.

Q4. *Find out the oxidation state of sodium in Na_2O_2.*

Ans. Let x be the oxidation state of Na in Na_2O_2

$$\begin{aligned} 2x + 2(-1) &= 0 \\ 2x - 2 &= 0 \\ 2x &= 2 \\ x &= +1. \end{aligned}$$

Q5. *Explain why is sodium less reactive than potassium.*

Ans. It is because ionization enthalpy ΔH_i of potassium = 419 kJ mol^{-1}.
Ionization enthalpy of sodium = 496 kJ/mol. Since Ionization enthalpy of potassium is less than that of sodium, potassium is more reactive than sodium.

Q6. *Compare the alkali metals and alkaline earth metals with respect to (i) ionization enthalpy, (ii) basicity of oxides, (iii) solubility of hydroxides.*

Ans. (*i*) **Ionization enthalpy.** Because of high nuclear charge the ionization enthalpy of alkaline earth metals are higher than those of the corresponding alkali metals.
(*ii*) **Basicity of oxides.** Basicity of oxides of alkali metals are higher than that of alkaline earth metals.
(*iii*) Solubility of hydroxides of alkali metals are higher than that of alkaline earth metals. Alkali metals due to lower ionization enthalpy are more electropositive than the corresponding group 2 elements.

Q7. *In what ways lithium shows similarities to magnesium in its chemical behaviour?*

Ans. (*i*) Both react with nitrogen to form nitrides.
(*ii*) Both react with O_2 to form monoxides.
(*iii*) Both the elements have the tendency to form covalent compounds.
(*iv*) Both can form complex compounds.

Q8. *Explain why can alkali and alkaline earth metals not be obtained by chemical reduction method.*

Ans. Alkali and alkaline earth metals are themselves better recucing agents, and reducing agents better than alkali metals are not available. That is why these metals are not obtained by chemical reduction methods.

Q9. *Why are potassium and caesium, rather than lithium used in photoelectric cells?*

Ans. Potassium and caesium have much lower ionization enthalpy than that of lithium. As a result, these metals easily emit electrons on exposure to light. Due to this, K and Cs are used in photoelectric cells rather than lithium.

Q10. *When alkali metal dissolves in liquid ammonia, the solution can acquire different colours. Explain the reason for this type of colour change.*

Ans. Alkali metals dissolve in liquid ammonia and give deep blue solutions which are conducting in nature because ammoniated electrons absorb energy in the visible region of light and impart blue colour.

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

Q11. *Beryllium and magnesium do not give colour to flame whereas other alkaline earth metals do so. Why?*

Ans. Due to small size, the ionization enthalpies of Be and Mg are much higher than those of other alkaline earth metals. Therefore, a large amount of energy is needed to excite their valence electron, and that's why they do not impart colour to the flame.

Q12. *Discuss the various reactions that occur in the Solvay process.*

Ans.

$$CaCO_3\ (s) \xrightarrow{\text{heat}} CaO + CO_2$$

$$NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$$

$$NaCl + NH_4OH + CO_2 \longrightarrow NaHCO_3 + NH_4Cl$$

$$2\ NaHCO_3 \xrightarrow{\text{heat}} Na_2CO_3 + CO_2 + H_2O$$

$$Na_2CO_3 + 10\ H_2O \longrightarrow Na_2CO_3 \cdot 10\ H_2O$$

Q13. *Potassium carbonate cannot be prepared by Solvay process. Why?*

Ans. Potassium carbonate being more soluble than sodium bicarbonate does not get precipitated when CO_2 is passed through a concentrated solution of KCl saturated with ammonia.

Q14. *Why is Li_2CO_3 decomposed at a lower temperature whereas Na_2CO_3 at higher temperature?*

Ans. Li_2CO_3 is a covalent compound whereas Na_2CO_3 is an ionic compound. Therefore, Lattice energy of Na_2CO_3 is higher than that of Li_2CO_3. Thus, $LiCO_3$ is decomposed at a lower temperature.

Q15. *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

Ans. (*a*) Nitrates of both group 1 and group 2 elements are soluble in water because hydration energy is more than the lattice energy.

Nitrates of both group 1 and group 2 elements are thermally unstable but they decompose differently except $LiNO_3$ *e.g.*

$$2NaNO_3 \xrightarrow{\text{heat}} 2NaNO_2 + O_2$$

$$2KNO_3 \xrightarrow{\Delta} 2KNO_2 + O_2$$

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

$$2Mg(NO_3)_2 \xrightarrow{\Delta} 2MgO + 4\ NO_2 + O_2$$

(*b*) Carbonates of group 1 elements are soluble in water except Li_2CO_3

They are also thermally stable except Li_2CO_3

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

Group 2 carbonates are insoluble in water because their Lattice energy are higher than hydration energy.

Thermal stability of carbonates of group 2 increases down the group because Lattice energy goes no increasing due to increase in ionic character.

(*c*) Sulphates of group 1 are soluble in water except Li_2SO_4. They are thermally stable.

Solubility of sulphates of group 2 decreases down the group because Lattice energy dominates over hydration energy.

Sulphates of group 2 elements are thermally stable and increasing down the group due to increases in Lattice energy.

Q16. *Starting with sodium chloride how would you proceed to prepare.*

(i) Sodium metal *(ii) Sodium hydroxide*

(iii) Sodium peroxide *(iv) Sodium carbonate?*

Ans. (*i*) **Sodium metal** is manufactured by electrolysis of a fused mass of NaCl 40% and $CaCl_2$ 60% in Down's cell at 873 K, using iron as cathode and graphite as anode. Na is liberated at the cathode.

At cathode:

$$Na^+ + e^- \longrightarrow Na\ (l)$$

At anode:

$$2Cl^-\ (melt) \longrightarrow Cl_2\ (g) + 2e^-$$

(*ii*) **Sodium hydroxide** is manufactured by electrolysis of an aqueous solution of NaCl (brine) in Castner-Kellner cell.

At cathode:

$$Na^+ + e^- \longrightarrow \underset{\text{Sodium}}{Na}$$

$$2Na + Hg \longrightarrow Na - Hg$$

$$2Na - Hg + 2H_2O \longrightarrow 2NaOH + H_2 + Hg$$

At anode:

$$Cl^- - e^- \longrightarrow Cl$$

$$Cl + Cl \longrightarrow Cl_2$$

(*iii*) **Sodium peroxide:**

$$4Na + O_2 \longrightarrow 2Na_2O$$

$$2Na_2O + O_2 \longrightarrow 2Na_2O_2$$

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

$$NaCl + NH_3 + CO_2 + H_2O \longrightarrow NaHCO_3 \downarrow + NH_4Cl$$

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

Q17. *What happens when (i) magnesium is burnt in air, (ii) Quick lime is heated with silica (iii) chlorine reacts with slaked lime (iv) calcium nitrate is heated?*

Ans. (*i*) $2Mg(s) + O_2(g) \xrightarrow{\Delta} 2MgO(s)$

(*ii*) $CaO(s) + SiO_2(s) \xrightarrow{\Delta} CaSiO_3(s)$

(*iii*) $2Ca(OH)_2 + 2Cl_2 \longrightarrow CaCl_2 + Ca(OCl)_2 + 2H_2O$

(*iv*) $2Ca(NO_3)_2(s) \xrightarrow{\Delta} 2CaO(s) + 4NO_2(g) + O_2(g)$

Q18. *Describe two important uses of each of the following:*

(i) caustic soda (ii) sodium carbonate (iii) quick lime

Ans. (*i*) **Caustic soda**

(*a*) It is used in the manufacturing of soap paper, artificial silk etc.

(*b*) It is used in textile industries.

(*ii*) **Sodium carbonate**

(*a*) Used in the softening of water, for laundry and cleaning purposes.

(*b*) It is used in glass manufacturing.

(*iii*) **Quick lime**

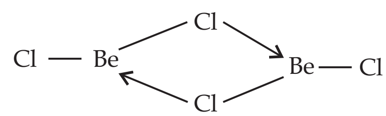
(*a*) It is used in the preparation of bleaching powder.

(*b*) Used in the purification of sugar and in the manufacturing of cement.

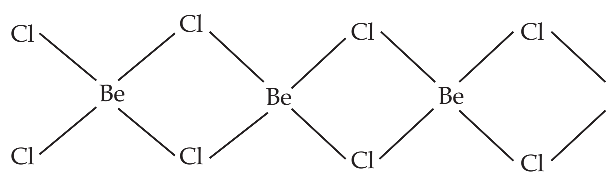
Q19. *Draw the structure of (i)* $BeCl_2$ *(vapour), (ii)* $BeCl_2$ *(solid).*

Ans. $BeCl_2$ (vapour)

In the vapour state, it exists as a chlorobridged dimer.



$BeCl_2$ (solid)



Q20. *The hydroxides and carbonates of sodium and potassium are easily soluble in water while the corresponding salts of magnesium and calcium are sparingly soluble in water. Explain.*

Ans. Since group 1 hydroxides and carbonates due to large size contain higher hydration energy than the lattice energy so, they are easily soluble in water. Whereas, in magnesium and calcium due to small size their lattice energy dominates over hydration energy they are sparingly soluble in water.

Q21. *Describe the importance of the following:*

(i) Limestone (ii) Cement (iii) Plaster of Paris.

Ans. Limestone:

(*i*) Extensively used in the manufacturing of high quality paper.

(*ii*) Used as mild abrasive in toothpaste.

(*iii*) As a filler in cosmetics.

(*iv*) Used as an antacid.

Cement:

(*i*) An important building material.

(*ii*) Used in concrete and reinforced cement.

Plaster of Paris:

(*i*) Used in plasters.

(*ii*) In dentistry, in ornamental work for making statues.

Q22. *Why are lithium salts commonly hydrated and those of the other alkali metal ions usually anhydrous?*

Ans. Due to smallest size, Li^+ can polarize water molecules easily than the other alkali metal ions.

Q23. *Why is LiF almost insoluble in water whereas LiCl soluble not only in water but also in acetone?*

Ans. It is due to high lattice energy of LiF as compared to LiCl.

LiCl is soluble in water because its hydration energy is higher than its lattice energy.

Q24. *Explain the significance of sodium, potassium, magnesium and calcium in biological fluids.*

Ans. **Sodium ions:**

(*i*) Na^+ ions participate in the transmission of nerve signals, in regulating the flow of water across cell membranes.

(*ii*) In the transport of sugars and amino acids into cell.

Potassium ions:

(*i*) They activate many enzymes.

(*ii*) Participate in the oxidation of glucose to produce ATP.

Magnesium ions:

(*i*) All enzymes that utilise ATP in phosphate transfer require magnesium as a cofactor.

(*ii*) Mg is the main pigment for the absorption of light in plants.

Calcium:

(*i*) Ca^{2+} ions are present in bones.

(*ii*) plays important roles in neuromuscular function.

Q25. *What happens when*

(i) Sodium metal is dropped in water?

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

(iii) Sodium peroxide dissolves in water?

Ans. (*i*) $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

(*ii*) $2Na + O_2 \longrightarrow Na_2O_2$

(*iii*) $Na_2O_2 + 2H_2O \longrightarrow 2NaOH + H_2O_2$

Q26. *Comment on each of the following observations:*

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

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

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

Ans. (*a*) Smaller the size of the ion, more highly it is hydrated and hence greater is the mass of the hydrated ion and thus the ionic mobility become lesser. The extent of hydration decreases in the order.

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

Thus the mobility of Cs^+ will be the highest.

(*b*) Due to its smaller size lithium can form nitride directly.

(*c*) It is because reduction potential depends upon sublimation energy, ionisation energy and hydration energy. Their resultant is almost constant for these ions.

Q27. *State as to why*

(a) a solution of Na_2CO_3 *is alkaline?*

(b) alkali metals are prepared by electrolysis of their fused chlorides?

(c) Sodium is found to be move useful than potassium?

Ans. (*a*) Na_2CO_3 is a salt of a weak acid, carbonic acid (H_2CO_3) and a strong base NaOH. Thus it undergoes hydrolysis to produce strong base NaOH and its aqueous solution is alkaline in nature.

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

(*b*) Because the discharge potential of alkali metals is much higher than that of hydrogen, therefore when the aqueous solution of any alkali metal chloride is subjected to electrolysis, H_2, instead of the alkali metal, is produced at the cathode. Therefore alkali metals are prepared by electrolysis of their fused chlorides.

(*c*) Since potassium is move reactive than sodium and it is found in nature to a less extent than Na, sodium is found to be more useful.

Q28. *Write balanced equations for reactions between.*

(a) Na_2O_2 and water

(b) KO_2 and water

(c) Na_2O and CO_2

Ans. (*a*) $Na_2O_2 + 2H_2O \longrightarrow 2NaOH + H_2O_2$

(*b*) $2KO_2 + 2H_2O \longrightarrow 2KOH + O_2 + H_2O_2$

(*c*) $Na_2O + CO_2 \longrightarrow Na_2CO_3$

Q29. *How would you explain the following observations?*

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

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

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

Ans. (*i*) Lattice energy of BeO is compartively higher than the hydration energy. Therefore, it is almost insoluble in water. Whereas $BeSO_4$ is ionic in nature and its hydration energy dominates the lattice energy.

(*ii*) Both BaO and $BaSO_4$ are ionic compounds but the hydration energy of BaO is higher than the lattice energy therefore it is soluble in water.

(*iii*) Since the size of Li^+ ion is very small in comparison to K^+ ion, it polarises the electron cloud of I^- ion to a great extent. Thus LiI dissolves in ethanol more easily than the KI.

Q30. *Which of the alkali metal is having least melting point?*

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

Ans. Size of Cs is the biggest thus, its melting point is the lowest. (*d*) is correct.

Q31. *Which one of the following alkali metals give hydrated salts?*

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

Ans. Li^+ is the smallest. Thus, it has the highest charge density and hence attracts the water molecules more strongly.

Q32. *Which one of the following alkaline earth metal carbonates is thermally most stable?*

(a) $MgCO_3$ (b) $CaCO_3$ (c) $SrCO_3$ (d) $BaCO_3$

Ans. (*d*) $BaCO_3$.

MORE QUESTIONS SOLVED

I. VERY SHORT ANSWER TYPE QUESTIONS

Q1. *Name the alkali metal which shows diagonal relationship with magnesium?*

Ans. Li.

Q2. *Why alkali and alkaline earth metals cannot be obtained by chemical reduction method?*

Ans. Because alkali and alkaline earth metals are themselves stronger reducing agents than the majority of other reducing agents.

Q3. *Name the compounds used for the manufacture of washing soda by Solvay process.*

Ans. NaCl, $CaCO_3$ and NH_3.

Q4. *Which electrolyte is used to obtain sodium in Castner's process?*

Ans. Fused NaOH.

Q5. *What happens when crystals of washing soda are exposed to air?*

Ans. Monohydrate ($Na_2CO_3 \cdot H_2O$) is formed as a result of efflorescence.

Q6. *Name the alkaline earth metals whose salt do not impart colour to a non-luminous flame.*

Ans. Beryllium does not impart colour to a non-luminous flame.

Q7. *What is dead burnt plaster?*

Ans. It is anhydrous calcium sulphate ($CaSO_4$).

Q8. *What is Quick lime? What happens when it is added to water?*

Ans. CaO is quick lime. When it is added to water, $Ca(OH)_2$ is formed.

Q9. *Arrange the following in the increasing order of solubility in water.*
$MgCl_2$, $CaCl_2$, $SrCl_2$, $BaCl_2$

Ans. $BaCl_2 < SrCl_2 < CaCl_2 < MgCl_2$

Q10. *Give the chemical formula of Epsom salt.*

Ans. $MgSO_4 \cdot 7H_2O$

Q11. *How would you prepare sodium silicate from silica?*

Ans. $$Na_2CO_3 + SiO_2 \xrightarrow{\text{Fuse}} \underset{\text{Sod. Silicate}}{Na_2SiO_3} + CO_2$$

Q12. *What happens when sodium metal is heated in free supply of air?*

Ans. Sodium peroxide is formed.

$$2Na + O_2 \longrightarrow Na_2O_2$$

Q13. *What is the general name for elements of group 1?*

Ans. Alkali metals.

Q14. *Why are alkali metals soft?*

Ans. Since the atoms of alkali metals have bigger kernels and smaller number of valence electrons, the metallic bonds in them are very weak and hence are soft.

Q15. *What do you mean by diagonal relationship in periodic table?*

Ans. The resemblance of the first element of second period with diagonally situated element of neighbouring element is called diagonal relationship.

Q16. *Why is $BeCl_2$ soluble in organic solvent?*

Ans. Since $BeCl_2$ is a covalent compound it is soluble in organic solvent.

Q17. *Why do alkali metals give characteristic flame colouration?*

Ans. Alkali metals due to low ionization energy absorb energy from visible region to radiate complementary colour.

Q18. *Why is the solution of alkali metals in liquid ammonia conducting in nature?*

Ans. Due to ammoniated electrons and cations.

Q19. *Which is more basic NaOH or $Mg(OH)_2$?*

Ans. NaOH is more basic.

Q20. *Which alkaline earth metals do not impart colour to the flame?*

Ans. Be and Mg.

Q21. *What is soda ash?*

Ans. Soda ash is anhydrous sodium carbonate (Na_2CO_3).

II. SHORT ANSWER TYPE QUESTIONS

Q1. *Why are alkali metals always univalent? Which alkali metal ion forms largest hydrated ion in aqueous solution?*

Ans. They are always univalent because after losing one electron, they aquire nearest inert gas configuration.

Li^+ forms largest hydrated cations because it has the highest hydration energy.

Q2. *What is the effect of heat on the following compounds (Give equations for the reactions)?*

(i) $CaCO_3$ *(ii)* $CaSO_4 \cdot 2H_2O$

Ans. (*i*) $$CaCO_3 \xrightarrow{\text{heat}} CaO + CO_2$$

(*ii*) $$CaSO_4 \cdot 2H_2O \xrightarrow{\text{heat}} \underset{\text{Dead burnt plaster}}{CaSO_4} + 2H_2O$$

Q3. *Explain the following:*

(a) Lithium iodide is more covalent than lithium fluoride.

(b) Lattice enthalpy of LiF is maximum among all the alkali metal halides.

Ans. (*a*) According to Fazan's rule, Li^+ ion can polarise I^- ion more than the F^- ion due to bigger size of the anion. Thus LiI^- has more covalent character than LiF.

(*b*) Smaller the size (internuclear distance), more is the value of Lattice enthalpy since internuclear distance is expected to be least in the LiF.

Q4. *Write the chemical formula of the following compounds.*

(i) Chile salt petre *(ii) Marble* *(iii) Brine*

Ans. (*i*) $NaNO_3$ (*ii*) $CaCO_3$ (*iii*) NaCl.

Q5. *Explain the following:*

(a) Why Cs is considered as the most electropositive element?

(b) Lithium cannot be used in making photoelectric cells.

(c) Lithium does not form alums.

Ans. (*a*) Due to its lowest ionization energy, Cs is considered as the most electropositive element.

(*b*) Lithium cannot be used in making photoelectric cells because out of all the alkali metals it has highest ionization energy and thus cannot emit electrons when exposed to light.

(*c*) Due to small size, lithium does not form alums.

Q6. *(a) What makes lithium to show properties uncommon to the rest of the alkali metals?*

(b) When is a cation highly polarising? Which alkali metal cation has the highest polarising power?

Ans. (*a*) The unusual properties of lithium as compared to other alkali metals is due to its exceptionally small size of atom and its ion and its high polarising power.

(*b*) A cation is highly polarising if its charge/size ratio is very high.

Li^+ ion has the highest polarising power.

Q7. *Why are ionic hydrides of only alkali metals and alkaline earth metals are known? Give two examples.*

Ans. Alkali metals and alkaline earth metals are most electropositive due to low ionization ethalpy therefore they form ionic hydrides. *e.g.* NaH, KH and CaH_2

Q8. *Why does the solution of alkali metals becomes blue in liquid ammonia? Give the chemical equation also.*

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

$$Na^+(am) + e^-(am) + NH_3(l) \longrightarrow NaNH_2(am) + \frac{1}{2} H_2(g)$$

Q9. *Give the important uses of the following compounds.*

(i) $NaHCO_3$ *(ii) NaOH*

Ans. (*i*) **Uses of $NaHCO_3$**

- It is used in fire extinguisher.
- It is mild antiseptic for skin infections.
- It is used as antacid.

(*ii*) **Uses of NaOH**

- It is used in soap industry.
- It is used in textile industry.
- It is used as reagent in laboratory.
- It is used in absorbing poisonous gases.

Q10. *What is the mixture of CaC_2 and N_2 called? How is it prepared?*

Ans. It is called Nitrolime.

It is prepared by heating CaC_2 with N_2 at high temperature.

$$CaC_2 + N_2 \longrightarrow \underset{\text{Nitrolime}}{CaCN_2 + C}$$

III. LONG ANSWER TYPE QUESTIONS

Q1. *(a) Compare four properties of alkali metals and alkaline earth metals.*

(b) What happens when alkali metals are dissolved in ammonia?

(c) $MgCl_2$ is electrolysed.

Ans. (*a*)

Alkali metals	*Alkaline earth metals*
(*i*) They are soft metals.	(*i*) They are harder than alkali metals.
(*ii*) Alkali metals show +1 oxidation state.	(*ii*) Alkaline earth metals show +2 oxidation state.
(*iii*) Their carbonates are soluble in water except Li_2CO_3.	(*iii*) Their carbonates are insoluble in water.
(*iv*) Except Li, alkali metals do not form complex compounds.	(*iv*) They can form complex compounds.

(*b*) They form blue coloured solution. The solution is paramagnetic in nature.

(*c*) $$MgCl_2 \xrightarrow{\text{electrolysis}} Mg^{2+} + 2Cl^-$$

At cathode:

$$Mg^{2+}(aq) + 2e^- \longrightarrow Mg(s)$$

At anode:

$$2Cl^-(aq) - 2e^- \longrightarrow Cl_2(g)$$

Q2. *State as to why*

(*a*) *Alkali metals show only +1 oxidation state.*

(*b*) *Na and K impart colour to the flame but Mg does not.*

(*c*) *Lithium on being heated in air mainly forms the monoxide and not the peroxide.*

(*d*) *Li is the best reducing agent in aqueous solution.*

Ans. (*a*) Alkali metals have low ionization enthalpies.

They have a strong tendency to lose 1 electron to form unipositive ions. Thus they show an oxidation state of +1 and are strongly electropositive.

(*b*) Valence electrons of alkali metals like Na and K easily absorb energy from the flame and are excited to higher energy levels. When these electrons return to the ground state, the energy is emitted in the form of light.

Magnesium atom has small size so electrons are strongly bound to the nucleus. Thus they need large amount of energy for excitation of electrons to higher energy levels which is not possible in bunsen flame.

(*c*) Due to the small size of Li^+ it has a strong positive field which attracts the negative charge so strongly that it does not permit the oxide ion, O^{2-} to combine with another oxygen atom to form peroxide ion.

(*d*) Since, among alkali metals, lithium has the most negative electrode potential ($E^0 = -3.04$ V) so, it is the strongest reducing agent in the aqueous solution.

IV. MULTIPLE CHOICE QUESTIONS

1. The reducing property of alkali metals follows the order

(*a*) Na < K < Rb < Cs < Li (*b*) K < Na < Rb < Cs < Li

(*c*) Li < Cs < Rb < K < Na (*d*) Rb < Cs < K < Na < Li

2. Which of the following is the least thermally stable?

(*a*) $MgCO_3$ (*b*) $CaCO_3$ (*c*) $SrCO_3$ (*d*) $BeCO_3$

3. When heated to 800°C, $NaNO_3$ gives

(*a*) $Na + N_2 + O_2$ (*b*) $NaNO_2 + O_2$ (*c*) $Na_2O + O_2 + N_2$ (*d*) $NaN_3 + O_2$

4. Lithium shows a diagonal relationship with

(*a*) sodium (*b*) silicon (*c*) nitrogen (*d*) magnesium

5. In the Solvay process

(*a*) an ammoniacal brine solution is carbonated with CO_2, forming $NaHCO_3$ which on decomposition at 150°C produces Na_2CO_3

(*b*) a sodium amalgum reacts with water to produce NaOH which gives Na_2CO_3 on reacting with CO_2

(*c*) A brine solution is made to react with $BaCO_3$ to produce Na_2CO_3

(*d*) all of the above

6. The oxide of which of the following metals is amphoteric?

(*a*) Pb (*b*) Mg (*c*) Ca (*d*) Al

7. Alkaline earth metals are

(*a*) more reactive (*b*) less reducing

(*c*) more oxidizing (*d*) less basic than alkali metals

8. Which of the following is not a peroxide?

(*a*) KO_2 (*b*) CrO_5 (*c*) Na_2O_2 (*d*) BaO_2

9. Hydrides as well as halides of alkaline earth metals tend to polymerize

(*a*) Sr (*b*) Ca (*c*) Be (*d*) Mg

10. Which of the following is used in photoelectric cells?

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

Ans. 1. (*a*) 2. (*d*) 3. (*c*) 4. (*d*) 5. (*a*)

6. (*a*) and (*d*) 7. (*b*) and (*d*) 8. (*a*) 9. (*c*) 10. (*d*)

V. HOTS QUESTIONS

Q1. *Why are alkali metals soft and have low melting points?*

Ans. Alkali metals have only one valence electron per metal atom. As a result, the binding energy of alkali metal ions in the close-packed metal lattices are weak. Therefore, these are soft and have low melting point.

Q2. *Which out of the following can be used to store an alkali metal?*

H_2O, C_2H_5OH and Benzene

Ans. Benzene can be used to store an alkali metal because other substances react with alkali metal as:

$$Na + H_2O \rightarrow NaOH + 1/2H_2$$

$$Na + C_2H_5OH \rightarrow C_2H_5ONa + 1/2H_2$$

Q3. *Potassium carbonate cannot be prepared by Solvay process. Why?*

Ans. This is due to the reason that potassium bicarbonate ($KHCO_3$) formed as an intermediate (when CO_2 gas is passed through ammoniated solution of potassium chloride) is highly soluble in water and cannot be separated by filtration.

Q4. *The hydroxides and carbonates of sodium and potassium are easily soluble in water while the corresponding salts of magnesium and calcium are sparingly soluble in water. Explain.*

Ans. All the compounds are crystalline solids and their solubility in water is guided by both lattice enthalpy and hydration enthalpy. In case of sodium and potassium compounds, the magnitude of lattice enthalpy is quite small as compared of sodium and potassium that are mentioned, readily dissolve in water. However, in case of corresponding magnesium and calcium compounds, the cations have smaller sizes and more magnitude of positive charge. This means that their lattice enthalpies are more as compared to the compounds of sodium and potassium. Therefore, the hydroxides and carbonates of these metals are only sparingly soluble in water.

Q5. *Why is LiF almost insoluble in water whereas LiCl soluble not only in water but also in acetone?*

Ans. The low solubility of LiF in water is due to its very high lattice enthalpy (F^- ion is very small in size). On the other hand, in lithium chloride (LiCl) the lattice enthalpy is comparatively very small. This means that the magnitude of hydration enthalpy is quite large. Therefore lithium chloride dissolves in water. It is also soluble in acetone due to dipolar attraction. (Acetone is polar in nature).

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