CHEMISTRY

CLASS NOTES FOR CBSE

Chapter 10. S-Block Elements

01. Introduction

S.No.	Properties	Alkaline earth metals	Alkali metals
(i)	Electronic configuration	Two electrons are present in the valency shell. The configuration is ns^2.	One electron is present in the valency shell. The configuration of ns^1.
(ii)	Valency	Bivalent	Monovalent
(iii)	Electropositive nature	Less electropositive	More electropositive.
(iv)	Hydroxides	Weak bases less soluble and decompose on heating.	Strong bases, highly soluble and stable towards heat.
(v)	Bicarbonates	These are not known in free state. Exist only in solution.	These are known in solid state.
(vi)	Carbonates	Insoluble in water. Decompose on heating.	Soluble in water. Do not decompose on heating (Li_2CO_3 is an exception).
(vii)	Action of nitrogen	Directly combine with nitrogen and form nitrides.	Do not directly combine with nitrogen.
(viii)	Action of carbon	Directly combine with carbon and form carbides.	Do not directly combine with carbon.
(ix)	Nitrates	Decompose on heating evolving a mixture of NO_2 and oxygen.	Decompose on heating evolving only oxygen.
(x)	Solubility of salts	Sulphates, phosphates, fluorides, chromates, oxalates, etc. are insoluble in water.	Sulphates, phosphates, fluorides, chromates, oxalates, etc. are soluble in water.
(xi)	Physical properties	Comparatively harder. High melting points. Diamagnetic.	Soft, low melting points, Paramagnetic.
(xii)	Hydration of compounds	The compounds are extensively hydrated $MgCl_2.\ 2H_2O$, $BaCl_2.2H_2O$ are hydrated chlorides.	The compounds are less hydrated, NaCl, KCl, RbCl form non-hydrated chlorides.
(xiii)	Reducing power	Weak as ionisation potential values are high and oxidation potential values are low.	Strong, as ionisation potential values are low and oxidation potential values are high.

When alkali metals are dissolved in liquid ammonia, a blue coloured solution is obtained. It is paramagnetic in nature. However, on increasing the concentration of alkali metal a bronze-coloured solution is obtained. This solution is diamagnetic in nature. These solutions are meta-stable and the alkali metal solwly reacts with NH_3 to give amide and H_2.

$$\mathbf{2M^+\ (am) + 2e^-\ (am) + 2NH_3(\mathit{l}) \rightarrow 2MNH_2(am) + H_2(\mathit{g})}$$

(M = alkali metal, *am* = ammonia)

$$\textbf{\textit{e.g.,}}\quad \mathbf{2Na^+\ (am) + 2e^-\ (am) + 2NH_3(\mathit{l}) \rightarrow 2NaNH_2(am) + H_2(\mathit{g})}$$

Heavier alkali metal hydroxides form numerous hydrates *e.g.* NaOH, nH_2O (where n = 1, 2, 3, 4, 5 and 7). Although Li has the most negative $E^\circ_{Li^+/Li}$ (*i.e.* highest oxidation potential) its reaction with water is considerably less vigorous than that of sodium which has the least negative E° among the alkali metals. Other metals react explosively with water.

All the alkali metal halides have high enthalpies of formation (ΔH°_f).

- ΔH°_f vlaues for fluorides become less negative as we move down the group.
- ΔH°_f values for fluorides, bromides and iodides become more negative as we move down the group, *e.g.,* ΔH°_f(NaCl) = −411.15 kJ mol^{-1}, ΔH°_f(KCl) = −435.9 kJ mol^{-1},
- For a given alkali metal, ΔH°_f values always become less negative from F^- to I^-. ΔH°_f(kJ mol^{-1}); NaF (−573.65); NaCl (−411.15); NaBr (−361.06); NaI (−287.78).

Melting and boiling points of halides always follow the trend

fluoride > chloride > bromide > iodide.

In water, the low solubility of LiF is due to its high lattice energy and the low solubility of CsI in water is due to smaller hydration energies of its two ions. Other halides of Li (Cl^-, Br^- and I^-) are soluble in ethyl alcohol, acetone and ethyl acetate. LiCl is soluble in pyridine. Lithium is least reactive but strongest reducing agent among all the alkali metals. Lithium unlike other alkali metals forms no acetylide on reaction with acetylene.

Diagonal Relationship between Lithium and Magnesium

- Both LiCl and $MgCl_2$ are soluble in ethanol.
- Both lithium perchlorate and magnesium perchlorate are extremely soluble in alcohol.
- Both $LiCl_2$ and $MgCl_2$ are deliquescent and crystallise from aqueous solution as hydrates Licl.$2H_2O$. $MgCl_2.8H_2O$.
- Solid bicarbonates are not formed by Li and Mg.

The lesser solubility of $NaHCO_3$ in soda solvay process is partly due to common ion effect of the Na^+ ions from the excess NaCl present.

Larger cation stabilises larger anion e.g., For example,

- Li forms Li_2O, Na forms mostly Na_2O_2. K, Rb and Cs mostly form superoxides (KO_2, RbO_2, CsO_2)
- LiI reacts with KF to give LiF and KI

$$\mathbf{LiI + KF \rightarrow LiF + KI}$$

- The following reaction proceed better with KF than with NaF.

$$\mathbf{-C\text{–}C + MF \rightarrow -C\text{–}F + MCl}$$

Baking soda is $NaHCO_3$. Baking powder is a mixture of $NaHCO_3$, corn starch and an acid salt. Glauber's salt is Na_2SO_4. $10H_2O$.

- The density of potassium is lesser than that of sodium contrary to the expectations. This is probably because of the abnormal increase in atomic size on moving from Na (186 pm) to K (227 pm).

Hence, potassium is lighter than sodium. Lithium is lightest known metal (density = 0.534 g/cc). The second ionisation energies of all the alkali metals are very large because when one electron is lost from these elements, the resulting ions acquire noble gas configuration which are very stable and have high effective nuclear charge. Thus. large amount of energies are required to remove the second electron. Caesium is the most electropositive of all the alkali metals. Since alkali metals are highly electropositive (a) they (except Li) mostly form ionic compounds.

(b) their electronegativity values are low. Due to their low ionization energies, alkali metals (except Li) show photoelectric effect. Cs can emit photoelectrons even with red light. All the alkali metals are silvery white, soft and light metals. Relative ionic radii :

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

Relative ionic radii in water or Relative degree of hydration :

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

Metal	Li	Na	K	Rb	Cs
Colour	Crimson	Yellow	Violet	Red violet	Blue violet
Wave length* (nm)	670.8	589.0 589.6	404.4 766.5 769.9	420.2 421.5 780.0	455.6 459.3 697.5

Alkali metals are light metals. Their densities are as follows:

	Li	Na	K	Rb	Cs
Density (g/cs)	**0.54**	**0.97**	**0.86**	**1.53**	**1.87**

Li is the lightest metal at room temperature. Potassium is lighter than sodium. Lithium can't be stored in kerosene as it floats to the surface, due to its very low density. Therefore, lithium is generally kept wrapped in paraffin wax. In fact lithium is the lightest metal known. It is the least fusible, least dense and least soft of all the alkali metals. Although Li has the highest ionisation energy, yet Li is the strongest reducing agent, because of its large heat of hydration, which is sufficient to overcome its ionisation energy. Metallic nature of elements increases from Li to Cs because IE decreases in the same order. Ionic radii of alkali metals in water follows the order. $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$ Thus, in aqueous solution, due to larger ionic radius, Li^+ has lowest mobility.

KOH is a better absorber of CO2 than NaOH because potassium carbonate is more soluble and does not separate out. Specific heat of alkali metals decreases from Li to Cs. Na and K are 6^{th} and 7^{th} most abundant elements in the earth crust. When a pellet of NaOH is kept exposed to moist air for longer periods a white powder is formed which is sodium carbonate. Smaller the size and higher the charge on the ion, greater is its tendency to distort the electron cloud of anion so greater is its polarizing power (Fajan rule). The most metallic elements in the periodic table are caesium and francium. When burnt in air Li reacts with O_2 as well as N_2.

$$4Li(s) + O_2(g) \longrightarrow 2Li_2O(s)$$

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

Other alkali metals do not react with N_2.

Superoxide ion O^-_2 contains a three electron bond. The peroxides and superoxides are strong oxidizing agents. Peroxides react with water to give O_2.