Chemistry Study Materials for Class 11

(NCERT Based Revision Notes of Chapter- 10)

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s- block element

• 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.	$1s^2 2s^2 2p^6 3s^2 3p^6 3rf^{10} 4s^2$
Francium	Fr	$4p^6 4d^{w} 5s^2 5p^6 6s^1$ or *[Xe]* $6s^1$ *[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.

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

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

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

↓ 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.

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

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.

$M + (x + y)\ NH_3 \longrightarrow [M(NH_3)X]^+ + [e(NH3)y]^-$

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.

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