Unit

10 The s-Block Elements

- Elements in which the last electron enters the s-orbital are called **s-block elements.** Since s-subshell has only one orbit which can accommodate only two electrons, therefore there are only two groups of s-block elements.
- **Electronic configuration**

Alkali metals

Element		Atomic No.	Electronic configuration
Name	Symbol		
Lithium	Li	3	[He] $2s^1$
Sodium	Na	11	[Ne] $3s^1$
Potassium	K	19	[Ar] $4s^1$
Rubidium	Rb	37	[Kr] $5s^1$
Cesium	Cs	55	[Xe] $6s^1$
Francium	Fr	87	[Rn] $7s^1$

Alkaline earth metals

Element		Atomic No.	Electronic configuration
Name	Symbol		
Beryllium	Be	4	[He] $2s^2$
Magnesium	Mg	12	[Ne] $3s^2$
Calcium	Ca	20	[Ar] $4s^2$
Strontium	Sr	38	[Kr] $5s^2$
Barium	Ba	56	[Xe] $6s^2$
Radium	Ra	88	[Rn] $7s^2$

- **Characteristic properties of s-block elements**
  - All s-block elements have high tendency to lose electrons to form positive ions so metallic in character. They possess very low electronegativity.
  - They all form basic oxides and hydroxides.
  - Predominantly form ionic bond with oxidation state +1 (alkali metals) and +2 (alkaline earth metals).
  - They have very weak tendency to form complexes.
  - Most of s-block elements show flame colouration due to low ionisation energy.

Group I		Group II	
Li	Deep red	Be	–
Na	Golden yellow	Mg	–
K	Pale violet	Ca	Brick red
Rb	Red violet	Sr	Blood red
Cs	Violet	Ba	Apple green
Fr	–	Ra	Crimson

- **Reaction with liquid ammonia :** Alkali metals and alkaline earth metals dissolve in liquid ammonia (solubility may be as high as 5 M). These dilute solutions are blue in colour and are conducting in nature.

$$M + (x + y)NH_3 \longrightarrow \underset{\text{Ammoniated cation}}{[M(NH_3)_x]^+} + \underset{\substack{\text{Ammoniated electron} \\ \text{(responsible for blue colour)}}}{[e(NH_3)_y]^-}$$

  - Ammoniated cations and electrons are responsible for conductivity.
  - In alkali metals, as concentration increases, the ammoniated metal ions may get bound to free unpaired electrons, known as expanded metals and solution turns bronze from blue.
  - In alkaline earth metals, evaporation of ammonia gives hexammoniates of the metals, which slowly decompose to give amides.

    $M(NH_3)_6 \longrightarrow M(NH_2)_2 + 4NH_3\uparrow + H_2\uparrow$

    Concentrated solutions of these metals in ammonia are bronze coloured due to formation of metal clusters.
  - The blue coloured solutions are paramagnetic while bronze coloured solutions are diamagnetic

GROUP-1 ELEMENTS (ALKALI METALS)

Occurrence

- Sodium and potassium are the most abundant (7th and 8th of all elements respectively) alkali metals found in earth's crust. Rubidium and cesium are found in very less amount. Francium being radioactive is found only in trace amount.
- Because of their high reactivity, they are never found in the free state. Alkali metals mostly occur as halides, oxides silicates, borates and nitrates in nature.

Element	Ore
Li	$LiAl(SiO_3)_2$: Spodumene, $LiAl(Si_2O_5)_2$: Petalite, $Li(AlF)PO_4$: Amblygonite
Na	NaCl : Common salt, $NaNO_3$: Chile Saltpetre, Na_3AlF_6 : Cryolite
K	$KCl.MgCl_2.6H_2O$: Carnalite, K_2CO_3 : Potash, KCl : Sylvine

Physical Properties

- **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.

Element	Atomic radii (pm)	Ionic radii (pm)	Atomic volume (mL)
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

- **Melting and boiling points :** The melting and boiling points of alkali metals are very low because of weak intermetallic bonding. Weak intermetallic bonding can be explained by their larger size and only one valence electron. There is a decrease in boiling point and melting point on moving down the group form Li to Cs.
- **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.
  - First ionization energy of alkali metals is very low but they have very high value of second ionization energy.

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

- **Hydration of ions :** The solubility of alkali metals salts in water is due to the fact that the cations get hydrated by water molecules. The degree of hydration depends upon the size of the cation. Smaller the size of a cation, greater is its charge density and hence greater is its tendency to draw electrons from molecules which are thus polarized.

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

(Relative ionic radii)

$Li^+_{(aq)} > Na^+_{(aq)} > K^+_{(aq)} > Rb^+_{(aq)} > Cs^+_{(aq)}$

(Relative ionic radii in water or relative degree of hydration)

- **Reducing property :** Since alkali metals have strong tendency to lose electrons they act as strong reducing agents. Tendency of an element to lose electrons in solution is measured in term of its standard oxidation potential ($E°_{\text{oxid.}}$).

Metal	Li/Li$^+$	Na/Na$^+$	K/K$^+$	Rb/Rb$^+$	Cs/Cs$^+$
$E°_{\text{oxid.}}$	+3.05	+2.71	+2.93	+2.93	+2.92

This shows that Li is the strongest reducing agent in solution.

Chemical Properties

- Alkali metals are highly reactive because of their low ionization energy. As the ionisation energy decreases with increase in atomic number, their reactivity also increases from Li to Cs.
- **Action of air and oxygen :** The alkali metals tarnish in air due to the formation of oxides, hydroxides and carbonates at their surface. They react violently with atmospheric water hence they are stored in kerosene oil or paraffin wax. When burnt in oxygen, lithium forms lithium oxide (Li_2O), sodium form sodium peroxide (Na_2O_2) and other alkali metals form superoxides (MO_2, where M = K, Rb or Cs).

$$M \xrightarrow{O_2} \underset{\text{Oxide}}{M_2O} \xrightarrow{H_2O} \underset{\text{Hydroxide}}{MOH} \xrightarrow{CO_2} \underset{\text{Carbonate}}{M_2CO_3}$$

$$M_2O \xrightarrow{O_2} \underset{\text{Peroxide}}{M_2O_2}$$

- **Action of water :** Alkali metals react readily and vigorously with water to form hydroxides with the liberation of hydrogen.

$$2M + 2H_2O \rightarrow 2MOH + H_2$$

(where M = Li, Na, K, Rb or Cs)

The reactivity with water increases on descending the group from Li to Cs, due to the increase in electropositive character in the same order : Li < Na < K < Rb < Cs.

  - *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

- **Action of hydrogen :** Alkali metals react with hydrogen forming ionic hydrides, M^+H^-. Again since electropositive character increases from Li to Cs, reactivity of alkali metals with hydrogen to form hydrides increases from Li to Cs.

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

Thermal stability : LiH > NaH > KH > RbH > CsH

Reactivity : LiH < NaH < KH < RbH < CsH

Heat of formation

ΔH_f (kcal/mol) : 21.6 → 13.9 → 14.4 → 13.5 → 13.0

- **Action with halogens :** The alkali metals combine readily with halogens (X_2) forming halides.

$$2M + X_2 \rightarrow 2M^+X^-$$

The ease of formation of alkali metal halides increases from Li to Cs because electropositive character increases from top to bottom in the group.

Li < Na < K < Rb < Cs : Reactivity toward halogens.

Reactivity of halogens towards a particular alkali metal follows the order : $F_2 > Cl_2 > Br_2 > I_2$.