The s-Block Elements

- Group 1 elements: lithium, sodium, potassium, rubidium, caesium, and francium
- Called alkali metals because their hydroxides are alkaline in nature

Electronic Configuration

- General outer electronic configuration is ns^1.

Physical Properties

- Silvery white, soft, and light metals
- Density increases down the group. (exception: Na is denser than K)
- Low melting and boiling points − Reason: weak metallic bond due to the presence of only one electron in the outermost shell

Alkali metals form oxides on combustion.

The stability of peroxides or superoxides increases on moving down the group (i.e., with increase in size).

Reasons − Larger the cations, more is the stabilization of large anions through lattice energy.

Hydroxides obtained by reaction of oxides with water − White crystalline solids

Hydroxides of alkali metals are the strongest of all the bases.

Halides

Preparation:

By the reaction of aqueous hydrohalic acid (HX) with appropriate oxide, hydroxide or carbonate

For a given metal −

Fluoride → Chloride → Bromide → Iodide

$\Delta_f H^{\ominus}$ becomes less negative →

Salts of oxo-acids

An oxo-acid is an acid which contains oxygen has at least one hydrogen atom attached to oxygen contains at least one other atom has the hydroxyl group and the oxo group attached to the same atom

Reasons for anomalous properties of lithium:

- Exceptionally small size of its atom/ion
- High polarizing power (charge/radius ratio)

Some important compounds of sodium:

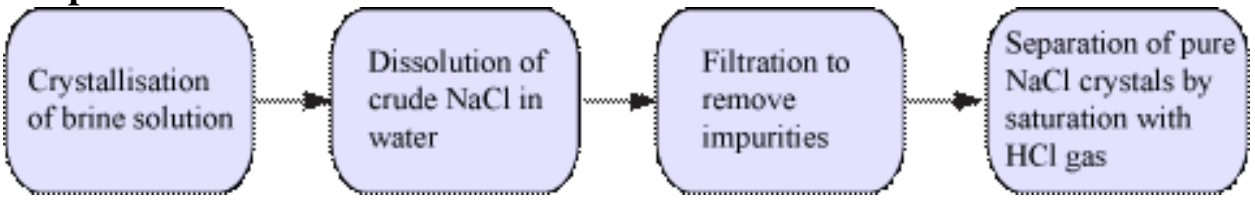
- Sodium carbonate (Washing soda), $Na_2CO_3 \cdot 10H_2O$

a. Prepared by Solvay process
b. Solvay process is not applied to manufacture K_2CO_3.

($KHCO_3$ is soluble in water and hence, cannot be precipitated.

- Sodium chloride, NaCl

a. **Preparation**



- Sodium hydroxide (Caustic soda), NaOH

a. **Preparation**

By electrolysis of NaCl in Castner – Kellner cell

At anode: $Cl^- \rightarrow \frac{1}{2}Cl_2 + e^-$

At cathode: $Na^+ + e^- \xrightarrow{Hg} Na - \text{amalgam}$

$$2Na - \text{amalgam} + 2H_2O \longrightarrow 2NaOH + 2Hg + H_2$$

- Sodium hydrogencarbonate (Baking soda), $NaHCO_3$

a. **Preparation**

$$Na_2CO_3 + H_2O + CO_2 \longrightarrow 2NaHCO_3$$

- Physical properties

1. Generally, silvery white, lustrous, and relatively soft, but harder than the alkali metals

   Be, Mg → greyish

1. Melting points and boiling points are higher than those of corresponding alkali metals → due to smaller sizes
2. Flame test → Be and Mg do not give any colour to flame (because their electrons are strongly bound).

- Atomic and ionic radii

1. Smaller than those of corresponding alkali metals in the same period (due to increased nuclear charge)
2. Increase on moving down the group

- Ionisation enthalpy

1. Low in a particular period
2. Decreases on moving down the group
3. First ionisation enthalpy → Higher than those of corresponding alkali metals (due to smaller size)
4. Second ionisation enthalpy → Smaller than those of corresponding alkali metals

- Hydration enthalpy

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

  (Decreases with increase in ionic size)

- Chemical properties

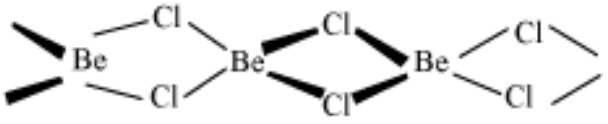
  Less reactive than alkali metals

1. Reactivity towards air and water:
2. React with air to form oxide and nitride
3. React with water to form hydroxides

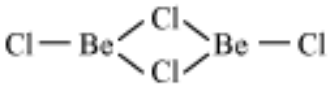
1. Reactivity towards halogens:

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

1. Beryllium halides are covalent in nature.



$BeCl_2$ in solid state



$BeCl_2$ in vapour phase

1. Reactivity towards hydrogen:

   Form hydrides, MH_2 (except Be)

1. Reactivity towards acids → liberate H_2

   $M + 2HCl \rightarrow MCl_2 + H_2$

1. Reducing nature:

   Strong reducing agent (but less stronger than corresponding alkali metals)

1. Solutions in liquid ammonia:

Dissolve in liquid ammonia to give deep blue black solutions

1. Salts of Oxoacids

- Thermal stability of Carbonates: $BaCO_3$> $SrCO_3$> $CaCO_3$> $MgCO_3$> $BeCO_3$
- Thermal stability of Sulphates: BaSO4> SrSO4> CaSO4> MgSO4> BeSO4

Compounds of alkaline earth metals are ionic in nature (but less ionic than compounds of corresponding alkali metals) because of greater nuclear charge and smaller size.

Oxides

(i) Burn in oxygen forming metal monoxide (MO)

(ii) MO has rock salt structure. (Exception − BeO)

(iii) Thermally stable − Because of high enthalpies of formation

(iv)Ionic in nature (Except BeO is amphoteric)

Hydroxides

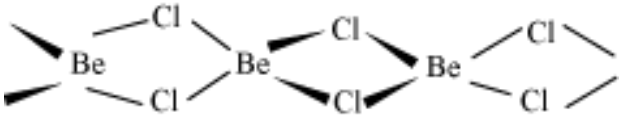
Oxides react with water to form hydroxides, which are sparingly soluble.

$$\xrightarrow[\text{(Solubility, thermal stability, basic character) increases}]{Mg(OH)_2,\ Ca(OH)_2,\ Sr(OH)_2,\ Ba(OH)_2}$$

Halides

(i) Ionic in nature (Exception − BeX_2 is covalent and soluble in organic solvents)

(ii) Structure of $BeCl_2$ in solid state (chain structure) is



(iii) In vapour phase, $BeCl_2$ forms chloro-bridged dimer. At high temperature (1200 K), this dimer undergoes dissociation to form linear monomer.

(iv) Tendency to form halide hydrates decreases down the group.

Salts of oxo-acids

Carbonates

(i) Insoluble in water

(ii) Precipitated by adding sodium or ammonium carbonate

(iii) Solubility decreases down the group.

(iv) When heated, decompose to given oxide and CO_2

(v) Thermal stability increases down the group. $BeCO_3$ is unstable and hence, kept in the atmosphere of CO_2.

Sulphate

(i) White solids

(ii) Thermally stable

(iii) Solubility decreases down the group.

(iv) $BeSO_4$ and $MgSO_4$ are readily soluble in water. Reason − High hydration enthalpy of Be^{2+} and Mg^{2+}

Nitrates

(i) They are prepared by dissolving carbonates in dilute nitric acid.

(ii) Tendency to form hydrates decreases down the group due to increase in size and decrease in hydration enthalpy.

- Small atomic and ionic size
- Does not exhibit coordination number more than four

- Oxide and hydroxide of Be (unlike others) are amphoteric in nature.

Some important compounds of calcium:

- Calcium oxide (Quick lime), CaO
- Calcium hydroxide (Slaked lime), $Ca(OH)_2$
- Calcium carbonate, CaCO3
- Calcium sulphate (Plaster of Paris), $CaSO4 \cdot \frac{1}{2} H2O$
- Calcium sulphate (Plaster of Paris), $CaSO_4 \cdot \frac{1}{2} H_2O$

Biological Importance of Mg and Ca:

- Adult body contains 25 g of Mg, 1200 g of Ca.
- About 99% of body Ca is present in bones and teeth.

Biological fluids: contain Na^+, K^+, Mg^{2+}, Ca^{2+}

- They perform maintenance of ion balance.
- They perform nerve impulse conduction.