The s-Block Elements

Contents

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10.1 Introduction:

The long form of the periodic table has been conveniently divided into four blocks i.e., s, p, d and f-blocks, depending upon the filling of a particular shell. The s-block elements are those in which the last electron enters the outermost s-orbital.as the s-orbital can accommodate only two electrons, this block includes only two groups; group 1 (or 1A) and group 2 (or II A). They have one or two electrons in their outermost s-subshell respectively.
The elements of group 1 are called alkali metals and those of group 2 are called alkaline earth metals. Therefore, **alkali metals and alkaline earth metals constitute s-block.**

10.2 Abundance and occurrence:

The alkali and alkaline earth metals are the most violently active of all the metals. These are readily oxidized and therefore, are not found in the free state in nature. These are found in the combined form with halide, sulphate, carbonate, silicate ions, etc.
The abundance of alkali and alkaline earth metals in the earth's crust shows a good range. Among these, **calcium** is the fifth most abundant element in the earth's crust and hence, the third most abundant metal after aluminum and iron. Vast sedimentary deposits of $CaCO_3$ occur over large parts of the earth's surface. **Magnesium** is the sixth and **sodium** is the seventh most abundant element in the crustal rocks. These are fourth and fifth most abundant metals after Al, Fe and Ca. **Potassium** (eighth) is the next most abundant element after sodium. There is unlimited supplies of NaCl in natural brines and oceanic waters. The occurrence of other metals in the earth's crust is very poor. Strontium and barium have much lower abundances. *Francium and radium are radioactive* and have rare terrestrial abundance. Francium is highly radioactive and its longest lived isotope ^{223}Fr has a half life of only 21 minutes. Beryllium is rare and radium is the rarest of all comprising only 10^{-10} per cent of igneous rocks.

10.3 Anomalous properties of first element in each group:

The elements in a group show similar physical and chemical properties. However, the first member of each group differs from its succeeding member (called **congeners**). For example, lithium shows an anomalous behavior as compared to sodium and rest of the family members of the alkali metal family. It is mainly due to the following reasons:

- **(i)** Small size of the atom and its ion.
- **(ii)** High ionization enthalpy and electronegativity.
- **(iii)** Non-availability of d-orbitals.

Let us study the effect of above factors on the chemistry of first element as compared to other elements (specially second). The first member differs from its succeeding members in some of the properties as given below:

- **(i)** As we go down a group, the size goes on increasing, therefore, the first member of each group has the smallest size in its group.
- **(ii)** Because of small size the first member has largest ionization enthalpy and ionization enthalpy decreases down the group.
- **(iii)** All the elements of second period has abnormally low electron affinity than the second member. For example, the electron affinity decreases as we move down a group but the first member has abnormally lower electron affinity than the second because of its small size.
- **(iv)** The small size of the atom results in relatively high cohesive properties associated with relatively strong intermetallic bonding. On the other hand, large atoms usually form weak bonds, therefore, the bond strengths of the compounds decrease as we move down the group. For example, lithium has relatively high heat of atomization, melting and boiling points, density and hardness.
- **(v)** The first member has higher electronegativity as compared to other members of the group. Therefore, it has greater tendency to form covalent bonds. Forexample,

lithium halides are covalent while halides of other members of group 1 are ionic in nature.

(vi) Because of very small size of the cation of second period (e.g., Li^+, Be^{2+}), it has high hydration energy and as a result, it acts as an excellent reductant in aqueous solution.

(vii) The first member of the group has no vacant d, orbitals in its valence shell. Therefore, it can have maximum of four co-ordination number i.e., only four electron pair. On the other hand, the elements of third period (second member of the group) have vacant 3d-orbitals in its valence shell. Therefore, these can have co-ordination number more than four. In other words, elements of second period cannot extend their octets while the elements of higher periods can extend their octets. For example, beryllium forms BeF_4^{2-} whole aluminium forms AlF_6^{3-} in solution.

The specific characteristics of distinction between **lithium** and **sodium** (group 1) and **beryllium and magnesium** (group 2) are discussed later in discussion of groups.

10.4 Diagonal relationship:

It has been observed that some elements of second period show similarities with the elements of the third period present diagonally to each other, though belonging to different groups. This is called **diagonal relationship.**

Thus, the similarity in properties of elements present diagonally is called diagonal relationship. This is shown below:

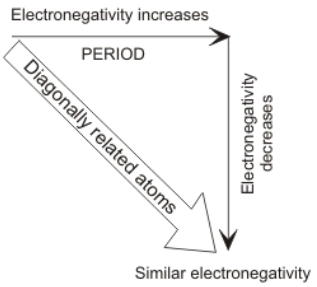
	Group 1	**Group 2**	**Group 13**	**Group 14**
Second period	**Li**	**Be**	**B**	**C**
Third period	**Na**	**Mg**	**Al**	**Si**

For example, lithium resemble with magnesium, the elements of group 2.

Cause of diagonal relationship. The cause of diagonal relationship is the similarity in properties such as electronegativity, ionization enthalpy, size or charge/radius ratio, etc. between the diagonal elements. For example, on moving from left to right across a period, the electronegativity increases and while moving down a group electronegativity decreases. Therefore, on moving diagonally, the two opposing tendencies almost cacel out and the elecronegativity values remain almost same as we move diagonally. Thus, the diagonal pairs have many similar properties.

Three important diagonal pairs. The following pairs exhibit diagonal similarity:

(i) Lithium - magnesium



(ii) Beryllium – aluminium

(iii) Boron – silicon.

The specific characteristics of diagonally related pairs such as Li – Mg and Be-Al are discussed later in the discussion of groups.

10.5 Group 1; Elements-The Alkali Metals

10.5.1 The Electronic Configurations of Alkali Metals:

Element	Symbol	Atomic number	Electronic Configuration	Abundance in earth's crust (ppm)
Lithium	Li	3	[He] $2s^1$	65
Sodium	Na	11	[Ne]$3s^1$	28,300
Potassium	K	19	[Ar] $4s^1$	25,900
Rubidium	Rb	37	[Kr] $5s^1$	310
Caesium	Cs	55	[Xe] $6s^1$	7
Francium	Fr	87	[Rn] $7s^1$	-

- All these elements are metallic in nature and are known as **alkali metals** because their hydroxides are strong bases or alkalies.

10.5.2 General Physical Characteristics of Alkali Metals:

- **Atomic Radii**. The **atomic radii** of alkali metals are largest in their respective periods. The atomic radii increase on moving down the group from top to bottom.
- **Ionic Radii.** Alkali metals change into positively charged ions by losing their outermost electron. These ions are considerably smaller than the parent atoms. Ionic radii increase on moving down the group.
- **Density.** Alkali metals have **low densities** due to their large atomic size. Densities increases on going down the group from top to bottom. Potassium is, however, lighter than sodium.
- **Ionization Energy.** The first **ionization energies** of alkali metals are very low as compared with the other elements of the same period.
- **Electropositive Character.** On account of their low ionisation energies, these metals have a strong tendency to lose their valence electrons and thus change into positive ions. Consequently, **alkali metals are strongly electropositive or metallic in character.** As this tendency for losing electrons increases down the group, the electropositive character increases.
- **Oxidation State.** Alkali metals exhibit an **oxidation state** of + 1.
- **Reducing Character.** Alkali metals are very **good reducing agents** because of their great tendency to lose electrons. The reducing character increases from Na to Cs. Li is however, stronger reducing agent than Na due to greater hydration energy.
- **Cohesive Forces.** Because of weak metallic bond, alkali metals have very **weak cohesive forces** and can be cut with the help of knife. Softness of alkali metals increases with increase in atomic number.
- **Melting and Boiling Points.** Alkali metals have very **low melting and boiling points** due to weak cohesive forces.
- **Photoelectric Effect.** Alkali metals exhibit **photoelectric effect** due to their low ionization energies. Because of this property potassium and caesium are used as cathodes in photoelectric cells.
- **Flame Colouration.** Alkali metals impart characteristic colours to the flame when they are heated in a Bunsen burner flame.

- **The compounds of alkali metals are diamagnetic.** Superoxides of alkali metals are, however, paramagnetic.
- **Hydration.** All the simple salts of alkali metals dissolve in water. In solutions alkali metal ions are hydrated.

10.5.3 Chemical Properties Of Alkali Metals

- **The alkali metals are highly reactive elements.** The reactivity of alkali metals is due to

  -low value of ionization energy; and

  -low heat of atomization.

- **The reactivity of alkali metals increases from Li to Cs.**

1. Hydrides. Alkali metals react with dry hydrogen to form hydrides.

$$2M + H_2 \xrightarrow{Heat} 2\overset{+}{M}\bar{H}$$

-These hydrides are ionic in nature and exist as crystalline solids.

-These hydrides of alkali metals react with water to form corresponding hydroxides and hydrogen gas.

$$LiH + H_2O \rightarrow LiOH + H_2$$

$$NaH + H_2O \rightarrow NaOH + H_2$$

-These hydrides are ***strong reducing agents*** and their reducing nature **increases down the group.**

-Alkali metals also form complex hydrides such as $LiAlH_4$ and $NaBH_4$, which are also good reducing agents.

$$4LiH + AlCl_3 \xrightarrow{\text{Dry ether}} Li[AlH_4]$$

$$4NaH + \underset{Trimethylborate}{B(OCH_3)_3} \xrightarrow{500-550K} Na\left[BH_4\right] + 3NaOCH_3$$

2. Oxides. Alkali metals on reaction with air form different kinds of oxides. For example, the alkali metals on reaction with limited quantity of oxygen form normal oxides of the formula, M_2O.

$$4M + O_2 \rightarrow 2M_2O$$

(Where M = Li, Na, K, Rb, Cs).

However, when heated with excess of air, lithium forms **normal oxide,** Li_2O, sodium forms *peroxide*, Na_2O_2, whereas potassium, rubidium and caesium form **superoxides** having general formula MO_2.

$$4Li + O_2 \rightarrow 2Li_2O$$

Lithium oxide

$$2Na + O_2 \rightarrow Na_2O_2$$

Sodium peroxide

$$K + O_2 \rightarrow KO_2$$

Potassium superoxide

The **normal oxides** of alkali metals dissolve readily in water to form hydroxides and a large amount of heat is produced.

$$M_2O(s) + H_2O(l) \rightarrow 2M^+ (aq) + 2OH^- (aq) + Heat$$

The **peroxides** and **superoxides** are also readily hydrolysed by water

$$Na_2O_2(s) + 2H_2O(l) \rightarrow 2Na^+ (aq) + 2HO^-(aq) + H_2O_2(l)$$

$$2KO_2(s) + 2H_2O(l) \rightarrow 2K^+(aq) + 2OH^-(aq) + H_2O_2(l) + O_2(g)$$

- **Basic nature of Alkali Metal Hydroxides.** The hydroxides of alkali metals are strong bases, all of which are highly soluble in water as well as in alcohol and are stable to heat.

$$2LiOH \xrightarrow{Heat} Li_2O + H_2O$$

The **basic character of alkali metal hydroxides increases in going down the group**. This can be explained in terms of decreasing ionization energies in moving down the group. The decrease in ionization energies leads to weakening of the bond between alkali metal and hydroxide ion. This result in the increased concentration of hydroxyl ions in solution, i.e., **increased basic character**. Thus, the basic character of alkali metal hydroxides is in the order

LiOH < NaOH < KOH < RbOH < CsOH

- **Reaction with Water.** Alkali metals react with water and other compounds containing acidic hydrogen atoms such as hydrogen halides (HX) and acetylene (C_2H_2) and liberate hydrogen gas.

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

All the alkali metals when exposed to atomosphere react with oxygen and moisture to form oxides and hydroxides and therefore their surface gets tarnished. In order to protect from atmosphere oxygen and *water, these metals are stored under kerosene oil.*

- **Reaction with Halogens.** Alkali metals react with halogens to form metal halides, which are ionic crystalline solids having general formula M^+X^-.

$$2M + X_2 \rightarrow 2\overset{+}{M}\bar{X}$$

All the halides of alkali metals **except lithium fluoride** are freely soluble in water. The low solubility of lithium fluoride is attributed to greater force of attraction between lithium ions and fluoride ions in the crystal lattice.

- **Reaction with Non-Metals**. Alkali metals, on heating react with non-metals such as sulphur and phosphorous to form sulphides and phosphides respectively.

$$2M + S \rightarrow M_2S$$

$$3M + P \rightarrow M_3P$$

- **Solubility in Liquid Ammonia.** All alkali metals dissolve in liquid ammonia giving *dark blue* solutions which are conducting in nature. These solutions contain ammoniated cations and ammoniated electrons as shown below.

$$M + (x + y)\ NH_3 \rightarrow M^+ (NH_3)x + e^{-1} (NH_3)y$$

- **Complex Compounds.** The alkali metal ions form fewer complex compounds than any other group of metal ions. This is due to their large atomic size and weak effective nuclear charge. The complex forming ability decreases in the order.

Li > Na > K > Rb > Cs.

Anomalous Behaviour Of Lithium And Its Diagonal Relationship With Magnesium

The properties of lithium are quite different from the properties of other alkali metals. On the other hand, it shows greater resemblance with magnesium, which is diagonally opposite element of group 2. The main reasons for the anomalous behavior of lithium as compared to other alkali metals are:

-The extremely small size of lithium atom and its ion.

-Greater polarizing power of lithium ion (Li^+), due to its small size which results in the greater covalent character in the compounds.

-Least electropositive character and highest ionization energy as compared to other alkali metals.

The reason for resemblance of properties of lithium with magnesium is that these two elements have almost same polarizing power.

The following points illustrate the anomalous properties of lithium and its diagonal relationship with magnesium:

- The melting point and boiling point of lithium are comparatively high.
- Lithium is much harder than the other alkali metals. Magnesium is also a hard metal.
- Lithium reacts with oxygen least readily to form normal oxide whereas other alkali metals form peroxides and super-oxides. Magnesium also forms normal oxide.
- LiOH like $Mg(OH)_2$ is a weak base. Hydroxides of other alkali metals are strong bases.
- Due to their appreciable covalent nature, the halides and alkyls of lithium and magnesium are soluble in organic solvents.
- Unlike other alkali metals but like magnesium, lithium forms nitride with nitrogen.

$$6Li + N_2 \rightarrow 2Li_3N$$

- Unlike other alkali metals lithium reacts directly with carbon to form ionic carbide. Magnesium also forms similar carbide.

$$2Li + 2C \rightarrow Li_2C_2$$

- The carbonates, hydroxides and nitrates of lithium as well as magnesium decompose on heating.

$$Li_2CO_3 \rightarrow Li_2O + CO_2$$

$$MgCO_3 \rightarrow MgO + CO_2$$

$$2LiOH \rightarrow Li_2O + H_2O$$

$$Mg(OH)_2 \rightarrow MgO + H_2O$$

$$4LiNO_3 \rightarrow 2Li_2O + 4NO_2 + O_2$$

$$2Mg(NO_3)_2 \rightarrow 2MgO + 4NO_2 + O_2$$

The corresponding salts of other alkali metals are stable towards heat.

- Li_2CO_3, LiOH, LiF and Li_3PO_4 are the only alkali metal salts which are insoluble in water. The corresponding magnesium compounds are also insoluble in water.
- The lithium ion and its compounds are more heavily hydrated than those of other alkali metals. Lithium chloride, like magnesium chloride, separates out from aqueous solutions as hydrated crystals, $LiCl.2H_2O$. Sodium and other alkali metal chlorides do not form hydrates at ordinary temperature, LiCl and $MgCl_2$ are deliquescent.

10.6 Elements of alkali metals and Their Compounds

10.6.1 Lithium (Li) and its compounds

Important ores-

Petalite $LiAl(Si_2O_5)_4 - 2.7 - 3.7\%$

Spodummene- $LiAl(SiO_3)_2 - 4 - 6\%$

Triphylite- $(Li, Na)_3PO_4 \cdot (Fe, Mn)_3\,(PO_4)_2 - 4\%$

Lepidolite or **Lithia mica**- $(Li, Na, K)_2, Al_2(SiO_3)_3, (F, OH)_2 - 1.5\%$

10.6.1.1 Physical Properties Of Lithium: Silvery white metal; harder than Na or K; Lightest metal known (density 0.634)

10.6.1.2 Chemical properties Of Lithium:

(i) With air: No effect by dry air

When heated air, $Li + O_2 \xrightarrow{200^\circ C} Li_2O$

(ii) With water $2Li + 2H_2O \xrightarrow{Cold} 2LiOH + H_2\uparrow$

(iii) Reacts **with dil. acids** to liberate hydrogen

(iv) With non-metals $(e.g., H_2, N_2, S \text{ and } X_2 \text{ etc.})$ form hydride, nitride, sulphide and halides, etc., LiH, Li_3N, Li_2S, LiX .

(v) With NH_3 -dark blue solution with liquid NH_3 however when NH_3 is passed over molten $Li \rightarrow LiNH_2$ is formed as

$2Li + 2NH_3 \longrightarrow 2LiNH_2 + H_2$

10.6.1.3 Compounds of Li

1. **Li_2O**: Prepared by burning Li in oxygen or air. However obtained in pure condition from $LiCO_3, LiOH, LiNO_3, Li_2SO_4,$ etc., heating, as:

$$Li_2CO_3 \longrightarrow Li_2O + CO_2$$

Li_2O is a white solid, that reacts slowly with water to form LiOH. It resembles other alkali hydroxides is possessing caustic properties.

2. **Li_2O_2**: It is prepared by adding H_2O_2 and alcohol to a concentrated solution of LiOH. It appears as white precipitate as $Li_2O_2 \cdot H_2O_2 \cdot 3H_2O$.

When dehydrated over P_2O_5 it loses water as well as H_2O_2 thus leaving behind Li_2O_2.

3. **LiOH**: Prepared by the action of ***Baryta water*** $[Ba(OH)_2]$ on lithium sulphate as:

$Li_2SO_4 + Ba(OH)_2 \longrightarrow 2LiOH + BaSO_4$

It is less soluble and much less thermally stable. It decomposes on heating as:

$$2LiOH \xrightarrow{\Delta} Li_2O + H_2O$$

4. **LiH**: Prepared by heating metal Li in a current of H_2.

   White solid substance, m.p. - 680°; electrovalent' reduces CO_2 to C; with water liberates H_2.

5. **$LiAlH_4$**: Prepared by reacting $AlCl_3$(anhy.) + LiH (in ethereal solution), powerful reducing agent.

   $$4LiH + AlCl_3 \xrightarrow{Ether} LiAlH_4 + 3LiCl$$

6. **LiCl**: Prepared by burning metal in Cl_2 or by adding HCl to $LiOH$ or Li_2CO_3, white crystalline solid extremely soluble in water also soluble in alcohol and pyridine.

7. **$LiNO_3$**: Prepared by dissolving Li_2CO_3 in dil. HNO_3. Highly soluble in water and alcohol, on heating gives $NO_2 + O_2$

8. **Li_2CO_3**: Prepared by $(NH_4)_2CO_3 +$ salt of $Li(soluble) \longrightarrow Li_2CO_3\ (ppt.)$

   White stable powder on heating above 600° decomposes it to give oxide.

   Sparingly soluble in water (≈ Ca and Mg).

   Carbonate + passing of CO_2 in solution $\longrightarrow LiHCO_3$

   $$(dissolved) \xrightarrow{\Delta} Li_2CO_3 + CO_2$$

Lithia water: Solution of $LiHCO_3$ used for gout and rheumatism.

10.6.1.4 Tests for Li:

1. **Flame test**: Crimson red.
2. **Phosphate test**: Sodium hydrogen phosphate +
   $$Li\ salt \longrightarrow (Na_2HPO_4) + \underset{White}{Li_3PO_4}\ \downarrow + HCl$$
3. **Carbonate test:** White ppt. wit ammonium carbonate. (Li_2CO_3) white

10.6.1.5 Anomalous behaviour of Li: Li differs in

1. Li hardest, its boiling point and melting point highest among the group
2. Li → least reactive,
3. Only alkali metals which reacts directly with N_2 to give nitride (Li_3N).
4. LiH → much more stable

   $$Li + H_2 \xrightarrow{700\text{–}800^\circ C} hydrides$$

   (Remaining hydrides are formed at 350-430°C)

5. With air or O_2 it only gives Li_2O, i.e., monoxide while remaining gives M_2O_2 (peroxide) and MO_2 (superoxide) also.

6. $2LiOH \xrightarrow{\text{Red hot}} Li_2O + H_2O$

Others hydroxides $\xrightarrow[hot]{\text{Red}}$ no reaction

$LiOH \longrightarrow lowest\ solubility$

7. $Li_2O \longrightarrow least\ basic$

8. Fluoride, phosphates, oxalate and carbonates of Na, K $\longrightarrow$ freely soluble

9. Carbonates and hydroxide of Li $\longrightarrow$ least stable (thermally)

Carbonates and hydroxide of other alkalies $\longrightarrow$ highly stable (thermally)

10. $LiNO_3 \xrightarrow{\Delta} Li_2O + NO_2 \uparrow + O_2 \uparrow$

$NaNO_3 \longrightarrow NaNO_2 + O_2 \uparrow$ (Thermally stable)

(Nitrates of other alkali metals also give same reactions)

11. Lithium sulphate $\longrightarrow$ does not form alums.

10.6.2 Sodium(Na) and its compounds

Important ores-NaCl (Rock salt); $NaNO_3$ (Chilesaltpetre); Na_3AlF_6 (Cryolite); $NaAlSi_3O_8$ (Felspar)

10.6.2.1 Properties of Sodium:

(i) Physical: Soft, silvery white metal; density 0.972 g/cc; m.p. = 97.8°C; b.p. = 883°C, malleable and ductile.

(ii) Chemical: (a) With air: Tarnished by moist air because firstly

$$Na + O_2 \xrightarrow{\Delta} Na_2O \xrightarrow{H_2O(air)} NaOH \xrightarrow{CO_2(air)} Na_2CO_3$$

When heated in dry air $\longrightarrow$ mixture of oxides (monoxide + peroxide)

(b) With water: Hydroxide + $H_2 \uparrow$

(c) With acids: $H_2 \uparrow$

(d) With H_2 : Electrovalent hydride

(e) With NH_3 : Deep blue solution and when heated in a current of $NH_3 \longrightarrow NaNH_2$

(f) With non-metals: With Cl_2, S and P forming NaCl, Na_2S, and Na_3P

(g) Reducing action: Strong, e.g.,

$Al_2O_3 + 6Na \longrightarrow 2Al + 3Na_2O$

$BeCl_2 + 2Na \longrightarrow Be + 2NaCl$

$AlCl_3 + 3Na \longrightarrow Al + 3NaCl$

(h) With alcohols:

$H_2 \uparrow as\ 2Na + 2C_2H_5OH \longrightarrow 2C_2H_5O^-Na^+ + H_2 \uparrow$

10.6.2.2 Compounds of Sodium(Na)

10.6.2.2.(a) Sodium Hydroxide NaOH (Caustic soda):

Preparation by electrolysis of NaCl [brine]

Until World War II, the principal method of producing sodium hydroxide, NaOH, involved conversion of $Ca(OH)_2$ to $CaCO_3$ using Na_2CO_3

$$Ca(OH)_2 + Na_2CO_3 \longrightarrow CaCO_3 + 2NaOH$$

The reaction goes to completion as is evident by measuring pH of the solution after reaction.

- Electrolysis of aqueous NaCl gives NaOH at cathode

$NaCl\ (aq) \rightleftharpoons Na^+ + Cl^-$

Anode: $2Cl^- \longrightarrow Cl_2 + 2e^-$

Cathode: $Na^+ + e^- \longrightarrow Na$

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

Side reaction may also occur if the products formed mix with one each other.

$2NaOH + Cl_2 \longrightarrow NaCl + NaClO + H_2O$

Or $2OH^- + Cl_2 \longrightarrow 2OCl^- + H_2$

And following reaction may also occur at the anode to a smaller extent:

$4OH^- \longrightarrow O_2 + 2H_2O + 4e^-$

Nelson Cell (Diaphragm cell)

If H_2 and Cl_2 gases mix, they react with explosion. A porous diaphragm of asbestos is used to keep the H_2 and Cl_2 (and thus cathode and anode) gases separated from one another. Product in this cell is a mixture of 11% NaOH and 16% NaCl. Concentration in a steam evaporator gives solution containing 50% NaOH and 1% NaCl. (Fig)

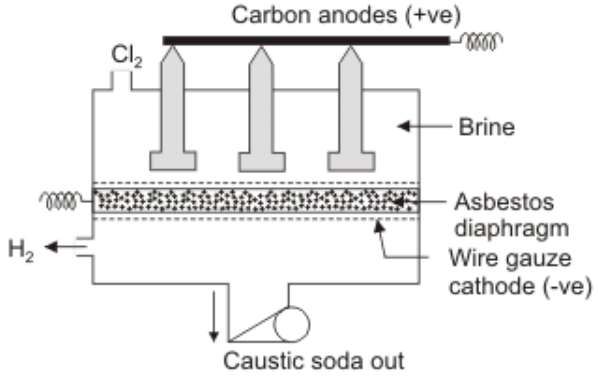


Fig. A Diaphragm Cell

Mercury Cathode Cell (Castner-Kellner Cell)

During the electrolysis of brine, Na^+ is discharged at cathode

$$Na^+ + e^- \longrightarrow Na$$

If the cathode is made of mercury, the Na atoms produced are dissolved in mercury to a different compartment and water reacts to form NaOH.

$$Na\ (ama\lg am) + H_2O \longrightarrow NaOH + \frac{1}{2}H_2 + Hg$$

Originally the anodes were made of graphite, but because traces of oxygen are produced in a side reaction, they become pitted with corrosion. The anodes are now made of steel coated with titanium (Fig).

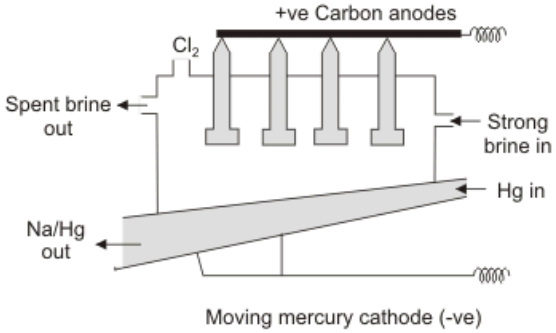


Fig. The Castner-Kellner Cell

Properties: White solid, m.p. = 318°C. Decomposes at 1300° as

$$2NaOH \longrightarrow 2Na + H_2 + O_2$$

Hydroscopic, highly soluble in water, $\longrightarrow$ gives strongly alkaline solution. Sparingly soluble in alcohol but less soluble than KOH.

- **With non-metals** (e.g., B, Si white P, S and halogens) $\longrightarrow$ sod. Salt of oxyacid of non-metal.

  i.e., $2B + 6NaOH \longrightarrow 2Na_3BO_3 + 3H_2\uparrow$

  $Si + 2NaOH + H_2O \longrightarrow Na_2SiO_3 + 2H_2\uparrow$

  $4P + 3NaOH + 3H_2O \longrightarrow 3NaH_2PO_4 + PH_3$

  $Cl_2 + 2NaOH \xrightarrow{Cold} NaClO + NaCl + H_2O$

  $3Cl_2 + 6NaOH \xrightarrow{Hot} NaClO_3 + 5NaCl + 2H_2O$

- **Weakly electropositive metals** dissolved in NaOH and evolve H_2. After forming complexes like

  $Zn \longrightarrow \underset{\text{Sod.zincate}}{Na_2ZnO_2};\ Al \longrightarrow \underset{\substack{\text{Sod.meta}\\ \text{aluminates}}}{NaAlO_2};\ Sn \longrightarrow \underset{\text{Sod.stannote}}{Na_2SnO_3}$

- Precipitates hydroxide from many salt solutions, e.g., it precipitates $Fe(OH)_2$ from $FeSO_4$ solution.
- **With CO** under pressure it gives sodium formate.

10.6.2.2.(b) Sodium chloride (Common salt), NaCl.

Preparation. The most abundant source of sodium chloride is sea water which contains 2.7 – 2.9% by mass. In tropical countries like India, common salt is generally obtained by evaporation of sea water. In India, about 50 lakh tons of salt are produced annually by solar evaporation. Crude sodium chloride obtained by crystallization of brine solution contains impurities of sodium sulphate (Na_2SO_4), calcium sulphate ($CaSO_4$), calcium chloride ($CaCl_2$) and magnesium chloride ($MgCl_2$). Since $MgCl_2$ and $CaCl_2$ are deliquescent (absorb moisture easily from the atomosphere), therefore, impure common salt gets wet in rainy season. To obtain pure sodium chloride, the crude salt is dissolved in minimum amount of water and filtered to remove insoluble impurities. The solution is then saturated with hydrogen chloride (HCl) gas when crystals of pure sodium chloride separate out. The calcium and magnesium chloride being more soluble remain in the solution.

Properties. Sodium chloride melts at 1081 K. Its solubility is 36.0 g per 100 g of H_2O at 273 K. However, the solubility does not increase appreciably with rise in temperature.

Uses.

(*i*) It is used as a common salt or table salt for domestic purposes.

(*ii*) It is used in the preparation of Na_2CO_3, NaOH and Na_2O_2.

10.6.2.2.(c) Sodium carbonate Na_2CO_3: ("Washing soda" or "soda ash")

From pre-historic times a natural deposit called Trona,$Na_2CO_3{\cdot}NaHCO_3{\cdot}2H_2O$(sesqui-carbonate)has been obtained from dried-up lake beds in Egypt.Trona is sometimes called sodium sesqui-carbonate(sesqui means one and half) and on heating it,Na_2CO_3 is obtained.

$2(Na_2CO_3{\cdot}NaHCO_3{\cdot}2H_2O) \xrightarrow{\Delta} 3Na_2CO_3+CO_2+5H_2O$

Manufacture by solvay ammonia process. The process involving following three steps:

Step (1) CO_2 is passed through a solution of $NH_3 \longrightarrow NH_4HCO_3$ is produced

Step (2) $NH_3 \longrightarrow NH_4HCO_3$ is then reacted with

$NaCl \longrightarrow NaHCO_3 \downarrow + NH_4Cl$

Step (3) Bicarbonate $\xrightarrow{\Delta} Na_2CO_3 + H_2O + CO_2$

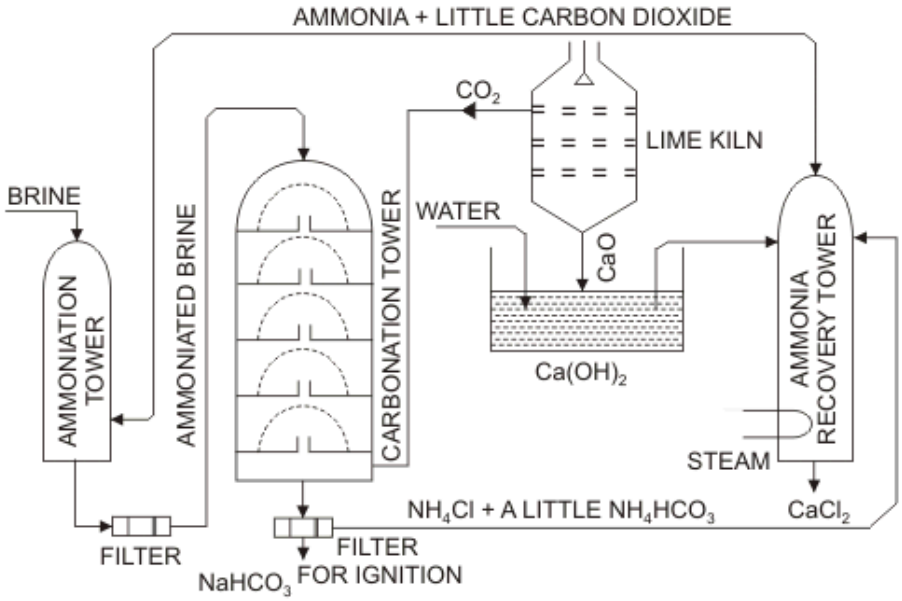
-CO_2 is obtained by the decomposition of $CaCO_3$

$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$

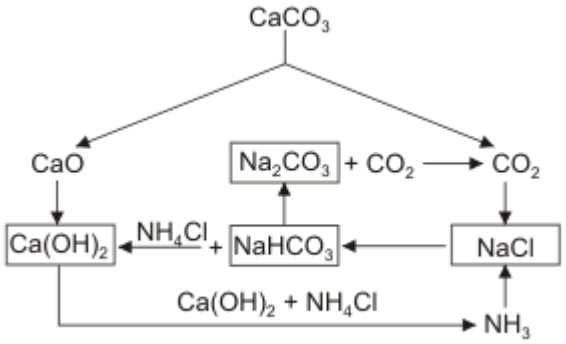
-CaO (aq) can decompose NH_4Cl to form NH_3 which can be reused

$Ca(OH)_2 + 2NH_4Cl \rightarrow 2NH_3 + 2H_2O + CaCl_2$

Thus, NH_3 is required in first instance only; it is recycled in further cycles. It is shown in the fig



Solvay-Ammonia process for the manufacture of sodium carbonate.



Flow-sheet for Solvay-Ammonia process.

Raw Materials: $NaCl, NH_3$ and limestone (for CO_2) Most of NH_3 can be recovered in the process $CaCl_2 \longrightarrow$ only by-product.

Properties: Crystallizes as decahydrate

$$Na_2CO_3 \cdot 10H_2O$$

$$\downarrow \ Gives\ on\ \text{exp}osure\ to\ air$$

$$No\ decomposition \longleftarrow Na_2CO_3\ (Thermally\ stable;) \xleftarrow{\Delta} Na_2CO_3 \cdot H_2O + 9H_2O\ (\text{Sol}uble\ in\ water;\ solution\ alkaline)$$

- **Fusion mixture** $\longrightarrow Na_2CO_3 + K_2CO_3$ mixture (as these two carbonates are thermally stable in nature)

10.6.2.2.(d) $NaHCO_3$: (Sodium bicarbonate)

Manufacture by Solvay process in lab. $NaHCO_3$ is prepared by saturating a solution of Na_3CO_3 (cold) with $CO_2 \longrightarrow NaHCO_3$ separates as white crystals (Due to its poor solubility in water).

Properties:

(a) Sparingly soluble in water; solution, mild alkaline

(b) If heated (solid or solution)

$$\xrightarrow[to]{Decomposes} Na_2CO_3 + CO_2$$

Produces CO_2 that's why used as constituent of ***"baking powder"*** and effervescent drinks.

10.6.2.2.(e) NaCN: (Sodium cyanide)

Step I

$$2NH_3 + 2Na \xrightarrow[300–400°C]{Iron\ retorts} 2NaNH_2 + H_2 \uparrow$$

Step II: $NaNH_2 + C \xrightarrow{(Metallic)} NaCN + H_2$

(ii) From calcium cynamide: Through crude calcium cyanamide

$[Ca(CN)_2] + NaCO_3 / NaCl + C$

(iii) From coal gas in its purification where HCN and NH_3 are evolved as impurities. These impurities are treated with $CuSO_4$ as,

$$CuSO_4 + NH_3 + HCN \rightarrow \underset{\text{Ammonium tricyano cuperate I}}{(NH_4)_2[Cu(CN)_3]} + (NH_4)_2SO_4 + CN_2$$

When all the copper salt has been converted into the complex cyanide, dil. H_2SO_4 is added to it. The HCN gas evolved is absorbed in NaOH solution to form NaCN as,

$$(NH_4)_2[Cu(CN)_3] + H_2SO_4 \longrightarrow (NH_4)_2SO_4 + 2HCN + CuCN$$

$$HCN + NaOH \longrightarrow NaCN + H_2O$$

Properties

- Colourless solid, readily soluble in water, solution alkaline due to hydrolysis of CN^- ion. Highly poisonous, smell of bitter almonds.
- Forms soluble complex cyanides with salts of transition elements like Cu, Au, Ag, Cd, Zn, Fe, etc., as

$$CdSO_4 + 2NaCN \longrightarrow Cd(CN)_2 + Na_2SO_4 \xrightarrow{2NaCN} Na_2[Cd(CN)_4]$$

- **Dilute solution of NaCN can dissolve Au or Ag in the presence of air or oxygen forming soluble complex cyanide as**

$$4Au + 8NaCN + 2H_2O + O_2 \longrightarrow 4Na[Au(CN)_2] + 4NaOH$$

$Na[Au(CN)_2]$ is sodium dicyano aurate (I)

- Large quantities of NaCN is used in the extraction of Au and Ag, along with, it is also employed for the **electroplating of Au and Ag** which involves electrolysis of their complex cyanides. On account of its highly poisonous character, it is also used as germicide.

10.6.2.2.(f) Sodamide $NaNH_2$:

Preparation: Steam of dry NH_3 over metal Na (iron retorts) 300-400°C.

Properties: Waxy solid, m.p. 210°C;

(a) With water $\longrightarrow NaOH + NH_3$;

(b) With dry $H_2 \longrightarrow (150-300°) \longrightarrow NaH + NH_3$

(c) With CO $\longrightarrow NaCN + NaOH + NH_3$

(d) With $CO_2 \longrightarrow Na_2CO_3 + CN \cdot NH_3 + H_2O$

(e) With $N_2O \longrightarrow NaN_3\ (sodium\ azide) + H_2O$

10.6.2.2.(g) Sodium oxide and peroxide(Na_2O and NaO)

- Sodium peroxide alongwith sodium oxide is formed when sodium burns in air or with $NaNO_2$

  $2Na + O_2 \longrightarrow Na_2O_2$

  $2Na + \frac{1}{2}O_2 \longrightarrow Na_2O$

  $Na_2O_2 + 2Na \longrightarrow 2Na_2O$

  $2NaNO_2 + 6Na \longrightarrow 4Na_2O + N_2$

- Na_2O is colourless ionic solid and aqueous solution is a strong base.

  $Na_2O + H_2O \longrightarrow 2NaOH$

- Na_2O_2 is diamagnetic (all electrons paired) and is regarded as a salt of dibasic acid H_2O_2.

  $Na_2O_2 + 2H_2O \longrightarrow 2NaOH + H_2O_2$.

- Na_2O_2 is colourless in pure form but is pale yellow in colour due to presence of sodium superoxide NaO_2 and is an oxidizing agent.

- Al is oxidised to Al_2O_3

  $3Na_2O_2 + 2Al \longrightarrow Al_2O_3 + 3Na_2O$

- Cr^{3+} salt is oxidised to CrO_4^{2-} (yellow)

  $Na_2O_2 + 2H_2O \longrightarrow 2NaOH + H_2O_2$

  $2Cr^{3+} + 3H_2O_2 + 10OH^- \longrightarrow \underset{yellow}{2CrO_4^{2-}} + 8H_2O$

  Because it reacts with CO_2 in the air, it has been used to purify air in submarines and confined spaces, as it both removes CO_2 and produces O_2.

$Na_2O_2 + CO \longrightarrow Na_2CO_3$

$2Na_2O_2 + 2CO_2 \longrightarrow 2Na_2CO_3 + O_2$

- $2Na_2O \xrightarrow{>400°C} Na_2O_2 + 2Na$ (vapour)
- Na_2O_2 gives H_2O_2 with dilute acids

  $Na_2O_2 + H_2SO_4\ (dil.) \longrightarrow H_2O_2 + Na_2SO_4$
- O_2 is formed when Na_2O_2 reacts with hot acids

  $Na_2O_2 + 4HCl(hot) \longrightarrow 4NaCl + 2H_2O + O_2$
- Na_2O_2 oxides $Mn(OH)_2$ to MnO_4^{2-} (green)

  $Mn(OH)_2 + 2Na_2O_2 \longrightarrow Na_2MnO_4 + 2NaOH$

  When mixture of $Ca(OH)_2$ and Na_2O_2 is compressed, CaO_2 is formed.

  $Na_2O_2 + Ca(OH)_2 \longrightarrow CaO_2 + 2NaOH$

10.6.2.2.(h) Microcosmic salt, $Na(NH_4)HPO_4$

Preparation of microcosmic salt can be describe as:

- It is prepared by dissolving ammonium chloride and disodium hydrogen phosphate in molecular proportion in hot water.

  $NH_4Cl + Na_2HPO_4 \longrightarrow Na(NH_4)HPO_4 + NaCl$
- The sparingly soluble microcosmic salt separates out. It is filtered and purified by recrystallisation.

Properties

(i) It is colourless crystalline solid, sparingly soluble in water.

(ii) When heated, it melts to form a clear transparent mass which has the property of combining with metallic oxides to form coloured orthophosphates.

$$Na(NH_4)HPO_4 \longrightarrow \underset{\substack{Sodium\\ metaphosphate}}{NaPO_3} + NH_3 + H_2O$$

$$NaPO_3 + CuO \longrightarrow \underset{Blue\ mass}{NaCuPO_4}$$

- On account of this property, microcosmic salt is used for the detection of coloured ions. The test is similar to borax bead test. The salt is heated on a loop of platinum wire as the result, a transparent bead is formed. When hot bead is brought in contact with coloured substance and strongly heated, a coloured bead is formed if the substance contains $Cu^{2+}, Ni^{2+}, Mn^{2+}, Co^{2+}$, etc. It is especially used for testing silica with which a cloudy bead containing floating properties of silica is obtained.

10.6.2.3 Tests of Na.

(a) Flame test : yellow colour.

(b) Pyroantimonats test :

$$\underset{\text{Solution of potassium pyroantimonate}}{K_2H_2Sb_2O_2} + \underset{\text{or any other sodium salt}}{NaOH} \rightarrow \underset{\text{Sodiumpyroantimonate}\downarrow \text{ whitecrystallineppt.}}{[Na_2H_2Sb_2O_2]}$$

(c) Sodium salt + solution of magnesium uranyl acetate

$\rightarrow$ yellow ppt. of sodium Magnesium uranylacetate $[NaMg(UO_2)_3(CH_3COO)_9 \cdot 9H_2O)]\downarrow$

10.6.3. Potassium(K):

Symbol K is derived from german word **Kali**. Important ores are

Carnallite -- $KCl \cdot MgCl_2 \cdot 6H_2O$

Polyhalite -- $K_2SO_4, MgSO_4, CaSO_4 \cdot 2H_2O$

Kainite -- $KCl, MgSO_4 \cdot 3H_2O$

Schonite -- $K_2SO_4 \cdot MgSO_4 \cdot 6H_2O$

Felspar -- $KAlSi_3O_8$;

Norwegian salt petre -- KNO_3;

Longbenite -- $K_2SO_4 \cdot 2MgSO_4$

- Resembles Na in its physical and chemical properties (softer and chemically more reactive) slightly radioactive.

10.6.3.1 Compounds of potassium:

10.6.3.1.(a) Potassium hydroxide(KOH)

Preparation :

(i) It is prepared by the electrolysis of KCl solution or from potassium carbonate by caustification process.

(ii) Pure KOH is obtained by adding potassium sulphate to a hot saturated solution of $Ba(OH)_2$.$BaSO_4$ is filtered off and filtrate is evaporated in a silver dish.

$$K_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + 2KOH$$

Properties:

- It purified like NaOH by dissolving in alcohol.
- It is white crystalline, highly delique3scent solid, extremely soluble in water; m.p. 360.4°C. Its action as an alkali, is stronger than that of NaOH.
- Absorbs CO_2 from air to form K_2CO_3.

- More soluble in alcohol than NaOH.

10.6.3.1.(b) Potash or pearl ash (K_2CO_3):

Preparation:

(i) From **Leblanc** process (It is an oldest process.)

$$KCl \xrightarrow{Hydrolysis} KOH \xrightarrow{Carbonation} K_2CO_3$$

(ii) Manufacture through Engel-Precht process:

By passing CO_2 in $MgCO_3 \cdot 3H_2O$ and KCl.

$$2KCl + CO_2 + 3[MgCO_3 \cdot 3H_2O] \longrightarrow 2[KHCO_3 \cdot MgCO_3 \cdot 4H_2O]\downarrow + MgCl_2$$

$$KHCO_3 \cdot MgCO_3 \cdot 4H_2O \xrightarrow[water]{Boiled\ with} K_2CO_3 + 2MgCO_3 + 9H_2O + CO_2$$

- **Properties:** White hygroscopic solid, m.p. -900°C very soluble in water, thermally stable.

10.6.3.1.(c) Potassium chloride(KCl): Its principal source is **starsfurt salt** where it occurs as **sylvine KCl,** and also as carnallite $KCl \cdot MgCl_2 \cdot 6H_2O$. **Dried sea weeds contain 90% KCl.**

KCl is prepared from fused carnalite-nearly pure KCl separates from the melt, leaving fused $MgCl_2$behind.

$KCl,MgCl_2.6H_2O \rightarrow KCl+MgCl_2.6H_2O$

Properties

It is colourless cubic crystal like solid soluble in water .Its solubility increases almost linearly with temperature.

10.6.3.1.(d) Potassium iodide(KI): It is prepared:

(i) From KOH

$$6KOH + 3I_2 \longrightarrow 5KI + KIO_3 + 3H_2O$$
$$\downarrow \text{Reuced by } C$$
$$KI + 3CO$$

10.6.3.1.(e)Potassium cyanide(KCN):

- KCN can be prepared by:

(i) By drying $K_4[Fe(CN)_6]$to white heat or with metallic K as:

$$K_4[Fe(CN)_6] \xrightarrow{2K} 6KCN + Fe$$

(ii) By the reaction of $K_4[Fe(CN)_6]$ with K_2CO_3 as:

$$K_4[Fe(CN)_6] + K_2CO_3 \longrightarrow 5KCN + KCNO + Fe + CO_2$$

- KCN is white crystalline solid, highly hygroscopic and very soluble in water, its aqueous solution is alkaline; **a deadly poison**.

10.6.3.1.(f) Potassium superoxide (KO_2)

Preparation: It is formed as a chrome-yellow (orange) powder by burning potassium in excess of moisture-free oxygen or air.

$$K + O_2 \longrightarrow KO_2$$

Example:

What is the most reactive alkali metal and why?

Solution:

The most reactive alkali metal is caesium due to its lowest first ionization enthalpy and lowest electronegativity.

Example:

The alkali metals have low densities. Explain.

Solution:

The alkali metals have low densities due to their large atomic sizes. In fact, Li, Na and K are even lighter than water.

Example:

The metallic lustre exhibited by sodium is explained by

(A) Diffusion of sodium ions (B) Oscillation of loose electrons

(C) Excitation of free protons (D) Existence of body–centered cubic lattice

Solution:

(B)

10.6.4 Biological Importance of Sodium and Potassium

A typical human being weighing about 70 kg contains about 90 g of Na, 170 g of K, 5 g of Fe and 0.06 g of Cu.

Although Na^+ and K^+ ions have similar chemical properties but surprisingly their biological functions are quite different. Whereas Na^+ ions are primarily found outside the cells in blood plasma and other interstitial fluids, K^+ ions are present inside the cells. These ions help in transmission of nerve signals, in regulating the flow of water across cell membranes and in the transport of sugars and amino acids into the cells.

Since K^+ ions are the most abundant cations within the cell fluids, they activate many enzymes and participate in the oxidation of glucose to produce ATP (adenosine triphosphate).

There is a very large variation in the concentration of Na^+ and K^+ ions found on the opposite sides of cell membranes. For example, in blood plasma, Na^+ ions are present to the extent of 143 mmol L^{-1} while the concentration of K^+ ions is only 5 mmol L^{-1} within the red blood cells. These concentrations change to 10 mmol L^{-1} (Na^+) and 105 mmol L^{-1} (K^+). These ionic gradients called the *sodium-potassium pump* operate across the cell membranes which consume more than one-third of the ATP used by a resting animal and about 15 kg per 24 h in a resting human being.

10.7 Group- 2; Elements-The Alkaline-Earth Metals

10.7.1 The Electronic Configurations of Alkaline-Earth Metals

Element	Symbol	Atomic number	Electronic Configuration	Abundance in earth's crust in (ppm)
Beryllium	Be	4	[He] $2s^2$	6
Magnesium	Mg	12	[Ne] $3s^2$	20,900
Calcium	Ca	20	[Ar] $4s^2$	36,300
Strontium	Sr	38	[Kr] $5s^2$	300
Barium	Ba	56	[Xe] $6s^2$	250
Radium	Ra	88	[Rn] $7s^2$	1.3×10^{-6}

All these elements are metallic in nature. These are known as **alkaline earth metals** because oxides of these metals were called **alkaline earths** as they existed in the earth and were alkaline in nature.

10.7.2 General Physical Characteristics of Alkaline–Earth Metals

Atomic Radii. The **atomic radii** of alkaline earth metals are quite larger but are smaller than those of alkali metals.

Density. The **densities** of alkaline earth metals are larger than those of alkali metals. This is due to stronger metallic bond. Density of alkaline earth metals first decreases from Be to Ca and then increases from Ca to Ra.

Cohesive Forces. Cohesive forces in alkaline earth metals are stronger than alkali metals due to smaller atomic radii and greater nuclear charge. Alkaline earth metals are harder than alkali metals due to stronger metallic bond.

Melting and Boiling Points. The melting and boiling points of alkaline earth metals are quite low but are higher than those of alkali metals.

Ionization Energy. The alkaline earth metals owing to their large size of atoms have fairly low values of **ionization energies** as compared to the *p*-block elements. However, within the group, the ionization energy, decreases as the atomic number increases.

A comparison of ionisation energies of elements of groups 1 and 2 shows that the members of group 2 have higher values. This is due to smaller size and higher nuclear charge of alkaline earth metals due to which the electrons in their outermost shell are tightly held.

It is interesting to note that although first I.E. of alkaline earth metals are higher than those of alkali metals, the second I.E. values are much lower than those of alkali metals.

- **Electropositive Character**. Alkaline earth metals are fairly **electropositive** in nature; however, due to their greater ionization energies these are less electropositive than alkali metals. In general, among the family members the electropositive character increases from Be to Ba.
- **Oxidation States**. The alkaline earth metals have two electrons in their valence shell and by losing these electrons; these atoms acquire the stable noble gas configuration. Thus, alkaline earth metals exhibit + 2 oxidation state in their compounds.

$$\underset{[Noblegas]\ ns^2}{M} \rightarrow \underset{[Noblegas]}{M^{2+}} + 2e^-$$

Reducing Properties. All the members of group 2, owing to their low ionization energies have a tendency to lose their valence electrons and thus act as strong reducing agents. Reducing character increases on moving down the group from top to bottom. However, the members of this group are weaker reducing agents than the alkali metals. *It is because of higher ionization energy of alkaline earth metals as compared to alkali metals.*

Flame colouration: Among the members of this group except Be and Mg other members give characteristics colouration to the flame.

Ca	Sr	Ba	Ra
Brick red	Crimson red	Grassy green	Crimson red

Beryllium and magnesium do not impart any colour to the flame because of their small size and higher innisation energies.Due to this, the energy of flame is insufficient to cause the excitation of their valence electrons to higher shells.

Solubility in liquid ammonia: all thses metals dissolve in liquid NH3.Dilute solutions are bright blue in colour while concentrated solutions are bronze coloured.theses solutions on evaporation yield hexammoiantes of metals which slowly decompose to give metal amides

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

10.7.3 Chemical Characteristics of Alkaline–Earth Metals

- Alkaline earth metals are **quite reactive** elements due to their *low ionization energies* and *high* electropositive character. The reactivity of these elements increases with increase in atomic number.
- Alkaline earth metals are less reactive than alkali metals.
- **Reaction with Water.** Ca, Sr and Ba react with cold water, liberating hydrogen gas

  $$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$

  $$Sr + 2H_2O \rightarrow Sr(OH)_2 + H_2$$

  Magnesium decomposes hot-water

  $$Mg + H_2O \rightarrow MgO + H_2$$

  Or $$Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$$

  Beryllium does not react with water.

  Order of Reactivity: Ba > Sr > Ca > Mg
- **Reaction with Oxygen.** All the alkaline earth metals burn in oxygen to form oxides Be, Mg and Ca form oxides whereas Ba and Sr form peroxides.

  $$2Mg + O_2 \rightarrow 2MgO$$

  $$Ba + O_2 \rightarrow BaO_2$$
- **Hydrides.** All the members of the group, except beryllium, combine with hydrogen on heating to form metal hydrides, MH_2.

  $$M + H_2 \rightarrow MH_2$$

  BeH_2 can, however, be produced by reduction of $BeCl_2$ with $LiAlH_4$.

  $$2BeCl_2 + LiAlH_4 \rightarrow 2BeH_2 + LiCl + AlCl_3$$

  BeH_2 and MgH_2 are covalent in nature and are polymeric, while CaH_2, SrH_2 and BaH_2 are ionic in nature.

The hydrides are good reducing agents and are hydrolyzed by water with evolution of H_2 gas

$$CaH_2 + 2H_2O \rightarrow Ca(OH)_2 + 2H_2$$

- **Hydroxides.** The hydroxides of these elements can be formed either by dissolving metal oxides in water or by reaction of these elements with water.

- $Be(OH)_2$ is amphoteric. The **hydroxides** of magnesium, calcium, strontium and barium are **bases** and their strength increases from magnesium to barium.
- These hydroxides are less soluble in water as compared to the alkali metal hydroxides.
- The solubility of the hydroxides in water increases with the increase in atomic number. $Be(OH)_2$ and $Mg(OH)_2$ are almost insoluble, $Ca(OH)_2$, is sparingly soluble while $Sr(OH)_2$ and $Ba(OH)_2$ are increasingly more soluble.

- **Halides.** Alkaline earth metals react with halogens directly to form halides having general formula, MX_2.

$$M + X_2 \rightarrow MX_2$$

These halides can also be obtained by the action of halogen acids on metals, their oxides, carbonates and hydroxides.

$$M(OH)_2 + 2HX \rightarrow MX_2 + 2H_2O$$

Beryllium chloride is prepared indirectly, from its oxide as follows:

$$BeO + C + Cl_2 \xrightarrow{870-1070K} BeCl_2 + CO$$

Beryllium chloride exists as monomer or dimer in vapour state but exists as polymer in solid state.

- **Carbonates.** Carbonates of alkaline earth metals are insoluble in water. These can be precipitated by addition of sodium or ammonium carbonate solution to the solution of salts of these metals. For example,

$$CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 \text{ (ppt)} + 2NaCl$$

- All the carbonates decompose on heating to give carbon dioxide and metal oxide.

$$MCO_3 \xrightarrow{\Delta} MO + CO_2$$

- However, as the atomic number increases, the stability of the carbonates towards heat **increases**. Beryllium carbonate is unstable and can be kept only in an atmosphere of CO_2.

- **Sulphates.** Sulphates of alkaline earth metals can be prepared by the reaction of H_2SO_4 with the metals, their oxides, hydroxides or carbonates.

$$Ca + H_2SO_4 \rightarrow CaSO_4 + H_2$$

$$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + CO_2 + H_2O$$

- The sulphates of alkaline earth metals are less soluble than the corresponding salts of alkali metals. Their solubilities decreases on going down the group.
- All these sulphates decompose on heating

$$MgSO_4 \xrightarrow{\Delta} MgO + SO_3$$

The thermal stability of sulphates increases on moving down the group from top to bottom.

- **Nitrides.** Alkaline earth metals burn in nitrogen to form nitrides, M_3N_2 which are ionic in nature.

$$3M + N_2 \rightarrow M_3N_2$$

Nitrides on reactions with water are hydrolysed and ammonia is released.

$$M_3N_2 + 6H_2O \rightarrow 3M(OH)_2 + 2NH_3$$

- **Carbides.** BeO when heated with C at about 2000^0C Be_2C is formed. The metals Mg, Ca, Sr and Ba form carbides of the formula MC_2. These carbides are formed when the metal or the metal oxide is heated with carbon in an electric furnace.
- These carbides are ionic in nature.
- Be_2C yields methane on hydrolysis whereas carbides of other metals yield acetylene

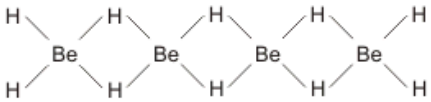
$$Be_2C + 4H_2O \rightarrow 2Be(OH)_2 + CH_4$$

$$CaC_2 + 2H_2O \rightarrow Ca(OH)_2 + C_2H_2$$

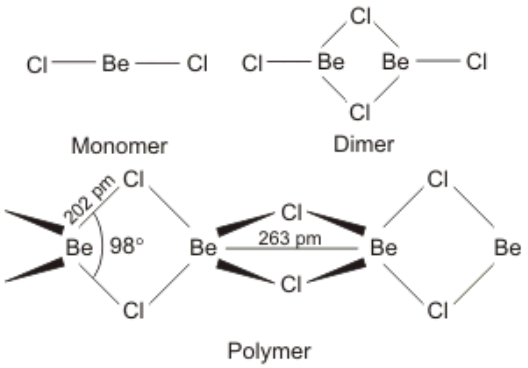
- **Complex Compounds.** Alkaline earth metals have a tendency to form stable complexes. Their ability to form complexes is more as compared to alkali metals. This is because of the smaller size and higher charge of the alkaline earth metal ions as compared to alkali metal ions. Among the elements of group 2, beryllium and magnesium, have greater tendency of complexation.
- Carbides of both Al and Be give methane on hydrolysis.
- $BeCl_2$ like Al_2Cl_6 has a bridged polymeric structure.
- The salts of beryllium as well as aluminium are extensively hydrolysed.
- Beryllium hydride is electron deficient and polymeric with multicentre bonding, just like aluminium hydride but unlike hydrides of other alkaline earth metals.

10.7.4 Some more important points about alkaline earth metal compounds are:

1. **BeO :** Very hard solid (4 : 4 coordination); rest oxides are with 6 : 6 co-ordination. These oxides are highly stable due to their large ionic crystal lattice energies. BeO and MgO are almost insoluble in water. **Due to their high melting points, MgO and CaO are used as basic lining of furnaces.**
2. The basic character and solubility of hydroxides increases down the group. **The aqueous suspension of $Mg(OH)_2$ is called milk of magnesis and used as anti acid drug (antacid).** $Ba(OH)_2$ is almost as strong as the alkali hydroxides and most soluble among all hydroxides.
3. BeH_2 and MgH_2 are known to have polymeric bridge structure with hydrogen present as bridge between two metal atoms.



4. $BeCl_2$ shows polymeric bridge structure in its solid state which it is monomeric in vapour phase with zero dipole moment.



5. Sulphate, carbonates, nitrates and hydroxides of these metals decompose on heating to yield oxides.

10.7.5 Main ores of alkaline earth metals are:

1. Be : In complex silicates *e.g.,* **felspar** and **mica.**

beryl – $Be_3Al_2\ (SiO_3)_6$

2. Mg: Magnesite – $MgCO_3$; **dolomite** – $MgCO_3.CaCO_3$; **carnallite** – $KCl\ .\ MgCl_2\ .\ 6H_2O$.

Asbestos – $CaMg\ (SiO_3)_4$

3. Ca: Largely as carbonates *e.g.,* limestone, chalk, calcite, marble, dolomite etc. and sulphates *e.g.,* gypsum, anhydride ($CaSO_4$) alongwith fluorspar (CaF_2), calcium silicate ($CaSiO_3$) etc.

4. Sr: **Celestine** ($SrSO_4$); **strontianite** $SrCO_3$.

5. Ba : $BaCO_3$ (witherite) ; $BaSO_4$ (barytes)

6. Ra: Pitchblende and carnotite.

Outlines of extraction of various alkaline earth metals are:

1. Be ores are first oxidized to give BeO which is then reduced with the help of Cu and coke to get Be.

2. Mg is usually extracted by the electrolysis of fused $MgCl_2$ or MgO. Carnallite is used as the source of $MgCl_2$ and magnesite for MgO. From MgO the metal can also be obtained through its reduction by C, Si or Al at about 2000° C in vacuum. In the modern process **(pidgeon process)** mixture of ferrosilicon and Al is used as reducing agent.

3. Ca: By the electrolysis of fused $CaCl_2$ mixed with about 12% CaF_2 in graphite vessel.

4. Sr and Ba – By reduction of their oxides on heating with Al (alumino-thermic process) at 1200° C.

$$3SrO + 2Al \rightarrow Al_2O_3 + 3Sr$$
$$3BaO + 2Al \rightarrow Al_2O_3 + 3Ba.$$

10.7.5.1 Some important compounds of Be and their important characteristic include:

1. BeO : Prepared by burning Be in air or by heating hydroxide, carbonate or nitrate of Be.

White powder, insoluble in water, partly ionic, amphoteric, unchanged by heating in air, On heating with carbon, silicon or boron it yields : Carbide (Be_2C) ; silicide (BeSi) or boride (Be_3B_2) ; with high m.p. (2570^0) $\rightarrow$ used as refractory material.

2. $BeCl_2$: Prepared by burning metal in Cl_2 or by HCl + metal, white crystalline solid, m. p. 44°, fumes in air like PCl_5 and very readily hydrolysed in the presence of water.

$$BeCl_2 + H_2O \rightarrow BeO + 2HCl$$

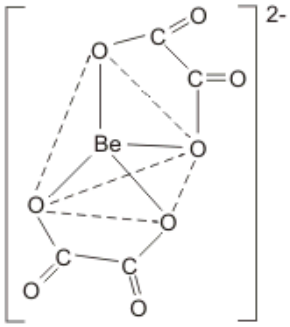
$$BeCl_2 + 2H_2O \rightarrow Be(OH)_2 + 2HCl$$

$BeCl_2 \rightarrow$ Dissolves freely in water (exothermic) $\rightarrow [Be(H_2O)_4]^{2+}$ on crystallization $[Be(H_2O)_4Cl_2]$ is deposited

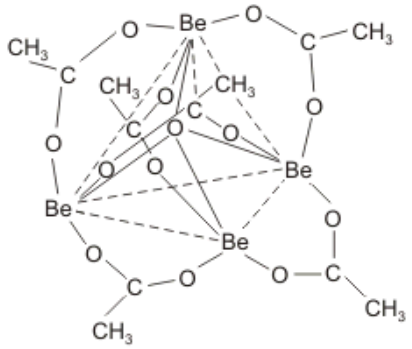
3. **$BeSO_4.4H_2O$:** Prepared by dissolving BeO in hot conc. H_2SO_4. Crystallites as tetrahydrate.

Tetrahydrate loses water at 400°C; anhyd. $BeSO_4$ decomposes above 500°C.

$BeSO_4 \xrightarrow{\Delta} BeO + SO_3$. Tetrahydrate $\rightarrow$ Very soluble; Anhydr. $\rightarrow$ almost (insoluble in water (cold); with hot water insoluble anhyd is converted into tetrahydrate (soluble).
Be gives Be oxalate ion with oxalic acid.
Be gives Be acetate with acetic acid $[Be_4O\ (CH_3COO)_6]$.



Structure of beryllium oxalate ion $[Be(OX)_2]^{2-}$



Structure of basic beryllium acetate, $Be_4O(CH_3COO)_6$

10.7.5.2 Some important compounds of Mg and their important characteristics

1. **MgH_2 :** Preparation by passing H_2 under pressure on heated Mg powder

$$(Mg + H_2 \rightarrow MgH_2)$$

Light grey solid, gives out H_2 on treatment with H_2O. Fairly stable and decomposes only when strongly heated.

2. **Mg_3N_2 :** Preparation by heating Mg to redness in NH_3 or N_2

$$3Mg + N_2 \rightarrow Mg_3N_2;$$
$$3Mg + 2NH_3 \rightarrow Mg_3N_2 + 3H_2$$

Light yellow solid; reacts with water as

$$Mg_3N_2 + 6H_2O \rightarrow 2NH_3 + 3Mg(OH)_2$$

3. MgO : Preparation when magnesite is heated to high temperature as

$$MgCO_3 \rightarrow MgO + CO_2$$

$$or \qquad 2Mg + O_2 \rightarrow 2MgO$$

Also preparation from

$$\left[KCl\right].MgCl_2.6H_2O \xrightarrow{Calcination} Mg(OH)_2 \xrightarrow{\Delta} MgO + H_2O.$$

White powder melts at 2300° C; basic; changes gradually to Mg $(OH)_2$ in water; **used as refractory material in furnaces.**

4. $Mg(OH)_2$: Preparation by the action of an alkali with Mg salt.

$$MgSO_4 + 2NaOH \rightarrow Mg(OH)_2 + Na_2SO_4$$

Sparingly soluble in water; ppt. by NH_4OH in the presence of NH_4Cl. Weak base when heated at high temperature changes to MgO.

5. $MgCl_2$: Obtained in laboratory in its hydrate form by dissolving $MgO(OH)_2$ or $MgCO_3$ in dilute HCl as :

$$MgO + 2HCl \rightarrow MgCl_2 + H_2O$$

$$MgCO_3 + 2HCl \rightarrow MgCl_2 + CO_2 + H_2O$$

Anhydrous $MgCl_2$ cannot be prepared by heating the hydrated salt because on heating it undergoes partial hydrolysis. Thus it is obtained.

(i) By heating Mg in current of dry chlorine or HCl gas as :

$$Mg + \underset{(dry)}{Cl_2} \rightarrow MgCl_2$$

$$Mg + \underset{(dry)}{2HCl} \rightarrow MgCl_2 + H_2$$

(ii) By heating $MgCl_2 . 6H_2O$ in vacuum as

$$MgCl_2.6H_2O \xrightarrow{\Delta} MgCl_2 + 6H_2O$$

(iii) By igniting the double salts of Mg and NH_4Cl. NH_4Cl sublimes leaving the anhydrous salt as

$$MgCl_2.\ NH_4Cl.6H_2O \xrightarrow{\Delta} MgCl_2.NH_4Cl \xrightarrow{\Delta} MgCl_2 + NH_4Cl(Sublimes)$$

(iv) By heating a mixture of MgO and C in a current of dry Cl_2

$$MgO + C + Cl_2 \rightarrow MgCl_2 + CO\uparrow$$

Properties of $MgCl_2$ include colourless, crystalline solid, highly deliquescent, exceedingly soluble in water, crystallises as $MgCl_2 .6H_2O$. On heating basic magnesium chlorides of variable composition are obtained as

$$MgCl_2.\ 6H_2O \xrightarrow{\Delta} Mg(OH)Cl + HCl + 5H_2O$$

On strong heating MgO is obtained as,

$$Mg(OH)Cl \xrightarrow{\Delta} MgO + HCl$$

When a saturated solution of $MgCl_2$ is mixed with MgO the resulting paste sets to a hard marble like mass having the composition $MgCl_2$. 5MgO. xH_2O. It is called as **magnesia cement** or **sorel cement**. This cement is **uses** as

(a) Dental filling

(b) A finish for plaster due to its high polish

(c) For cementing glass and porcelain

(d) For making artificial stones

NOTE:

When sorel cement is mixed with saw dust, cork waste etc, a fairly weather proof wood like material called xylotite is obtained which is used as covering for floor, laboratory tables etc.

6. $MgSO_4$: Occur as $MgSO_4$. $7H_2O$, commonly called Epsom salt (Epsom = Name of place where huge amounts of $MgSO_4$ seen). It is prepared

(i) In lab by the action of dil. H_2SO_4 with MgO or Mg metal as

$$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2\uparrow$$

$$MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$$

The solution on concentration gives crystals of $MgSO_4$.$7H_2O$.

(ii) By dissolving dolomite in boiling dilute H_2SO_4

$$MgCO_3 \ .CaCO_3 + 2H_2SO_4 \rightarrow MgSO_4 + CaSO_4 \ ppt. + 2H_2O + 2CO_2$$

Insoluble $CaSO_4$ is removed by filtration.

(iii) Commercially it is prepared by dissolving mineral **kiesserite** in boiling water and then crystallising the resulting solution.

Colourless, efflorescent, soluble in water, bitter in taste, isomorphous with $ZnSO_4$. $7H_2O$, can easily form double salts like $MgSO_4$. K_2SO_4 . $6H_2O$. On heating to 200° C all water of crystallisation is lost. On further heating MgO and SO_2 are obtained as

$$MgSO_4.7H_2O \xrightarrow{200°C} MgSO_4 + 7H_2O$$

$$MgSO_4 \xrightarrow[above\ 200°C]{\Delta} MgO + SO_2 + O_2$$

Test of Mg:

(a)Charcoal cavity test: Here, the Mg salt is heated till it is converted into white residue in charcoal cavity. This white residue on moistening with cobalt nitrate solution is converted into pale pink mass of MgO.CaO.

(b) Mg salt + α-nitrosobenzene-azo-β-naphthol+ NaOH$\rightarrow$ Blue coloured solution is obtained.

10.7.5.3 Some important compounds of Ca and their important characteristics

1. CaO (quick lime)

Preparation by heating limestone to $1000°C$ in lime lines as

$$CaCO_3 \rightleftharpoons CaO + CO_2$$

reversible reaction, proceeds efficiently in forward direction when CO_2 is allowed to escape from the system.

Properties : Amorphous white solid, with high melting point (2273K), usually obtained in the form of hard lumps and when heated in oxyhydrogen flame emits a bright light called **lime light.**

Easily absorbs CO_2 and H_2O as

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

Exothermic $\Delta H = -64.5kJ / mol$

$$CaO + CO_2 \rightarrow CaCO_3$$

The addition of limited quantity of water converts limp of lime to slaked lime (as the process is called **slaking)** Similarly, if we use soda in the place of water the product is called **soda lime (NaOH/CaO)**

Combines with solid acidic oxides at high temperature to give their salts as :

$$CaO + SiO_2 \rightarrow CaSiO_3$$

$$6CaO + P_4O_{10} \rightarrow 2Ca_3(PO_4)_2$$

Very useful as drying agent, in bleaching powder production (as used for the preparation of slaked lime) and as a constituent of **mortar.**

2. $Ca(OH)_2$ (slaked lime)

Preparation by adding water to quick lime as

$$CaO + 2H_2O \rightarrow Ca(OH)_2 + H_2\uparrow$$

White powder, sparingly soluble in water, **aqueous solution called lime water and aqueous suspension is called milk of lime.** CO_2 has the ability to turn lime water milky as

$$Ca(OH)_2 + CO_2 \rightarrow \underset{(Milypp.t)}{CaCO_3} + H_2O$$

This milkiness disappears with excess of CO_2 as soluble calcium hydrogen carbonate is formed:

$$CaCO_3 + H_2O + CO_2 \rightarrow \underset{Soluble}{Ca(HCO_3)_2}$$

Being basic in nature it reacts with acids and acidic gases forming their salts.

On reaction with chlorine, it forms hypochlorite a constituent of bleaching powder as :

$$2Ca(OH)_2 + 2Cl_2 \rightarrow CaCl_2 + CaOCl_2 + 2H_2O$$

Slaked lime is used in mortar production, lime water production, disinfectant, bleaching powder production etc.

3. $CaCO_3$ (Limestone)

Commonest form of $CaCO_3$ is lime stone and it is also found as chalk, marble, coral, calcite, aragonite etc. A precursor of quick lime and slaked lime. It is prepared as:

$$Ca(HCO_3)_2 \xrightarrow{\Delta} CaCO_3 + H_2O + CO_2\uparrow$$

Easily decomposed on heating as:

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2\uparrow$$

4. $CaSO_4$: $CaSO_4$ usually occur as $CaSO_4 . 2H_2O$ **(Gypsum** or **alabaster** or **selenite)**

It forms compounds like $CaSiO_3$ and $Ca(AiO_2)_2$ with SiO_2 and Al_2O_3 respectively.

5. Portland cement

Mixture of Ca silicates and Ca aluminates with small amount of gypsum.

Composition: CaO = 50-60%, SiO_2 = 20-25%,

Al_2O_3 = 5-10%; MgO = 2-3%, Fe_2O_3 = 1-2%; SO_3 = 1-2%

Essential raw materials-limestone for CaO, etc., **clay** for SiO_2, Al_2O_3, Fe_2O_3, etc.

6. Plaster of Pairs, $CaSO_4 .½ H_2O$ or $(CaSO_4)_2H_2O$

It occurs in nature as gypsum and the anhydrous salt as anhydride. It is prepared by precipitating a solution of calcium or nitrate with dilute sulphuric acid.

The effect of heat on gypsum or the dehydrate presents a review of interesting changes. On heating the monoclinic gypsum is first converted into orthorhombic form without loss of water. When the temperature reaches 120°C, the hemihydrate or plaster of paris is the product. The latter losses water becomes anhydrates above 200°C and finally above 400°C, it decomposes into calcium oxide.

$$2CaSO_4 \rightarrow 2CaO + 2SO_2 \uparrow + O_2 \uparrow$$

$$\underset{\text{Gypsum}\ (Monoclinic)}{CaSO_4 \cdot 2H_2O} \xrightarrow[H_2O]{Heating} \underset{\text{Hardening}}{CaSO_4 \cdot H_2O} \xrightarrow[Setting]{120^\circ C} \underset{H_2O}{CaSO_4 \cdot \frac{1}{2}H_2O} \xrightarrow{200^\circ C} \underset{\text{Plaster of Paris}}{CaSO_4}$$

The following conditions are necessary

(i) The temperature should not be allowed to rise above 393 K because above this temperature the whole water of crystallization is lost. The resulting anhydrous $CaSO_4$ is called dead burnt plaster because it loses the properties of setting with water.

(ii) The gypsum should not be allowed to come in contact with carbon containing fuel otherwise some of it will be reduced to calcium sulphite.

Properties

It is a white powder. On mixing with 1/3rd its weight, it forms a plastic mass which sets into a hard mass of interlocking crystals of gypsum within 5 to 15 minutes. It is due to this reason that it is called plaster. The addition of common salt accelerates the rate of setting, while a little borax or alum reduces it. The setting of plaster of paris is believed to be due to rehydration and its reconversion into gypsum.

$$\underset{\text{plater of paris}}{2CaSO_4 \cdot \frac{1}{2}H_2O} + 3H_2O \longrightarrow \underset{\text{gypsum}}{2CaSO_4 \cdot 2H_2O}$$

Uses:

(i) Plaster of paris is used for producing moulds for pottery and ceramics & casts of statues & busts.

(**ii**) It is used in surgical bandages used for plastering broken or fractured bones.

(iii) It is also used in dentistry.

10.7.5.3.1Industrial Uses of Lime, Slaked Lime and Limestone

Uses of lime

Calcium oxide is called lime or quick lime. Its main industrial uses are.

(i) It is used in steel industry to remove phosphates and silicates as slag.

(ii) It is used to make cement by mixing it with silica, alumina or clay.

(iii) It is used in making glass.

(iv) It is used in lime soda process for the conversation of Na_2CO_3 *to* $NaOH$ & Vice versa.

(v) It is used for softening water, for making slaked lime $Ca(OH)_2$ by treatment with water and calcium carbide CaC_2.

Uses of slaked lime $[Ca(OH)_2]$

(i) Slaked lime is used as a building material in form of mortar. It is prepared by mixing 3-4 times its weight of sand and by gradual addition of water. It sets into a hard mass by loss of H_2O and gradual absorption of CO_2 from air.

(ii) In manufacture of bleaching powder by passing Cl_2 gas.

(iii) In making glass and in the purification of sugar and coal gas.

(iv) It is used in softening of had water.

Uses of lime stone $(CaCO_3)$

(i) It is used as building material in form of marble.

(ii) In manufacture of quick lime.

(iii) It is used as a raw material for the manufacture of Na_2CO_3 in solvay- ammonia process.

(iv) Commercial limestone contains iron oxide, alumina,magnesia, silica & sulphur with a CaO content of 22-56% MgO, MgO content upto 21%. It is used as such as a fertilizer.

10.7.6 Bleaching Powder

The exact chemical composition of bleaching powder is not yet known but it behaves as if it contains calcium hypochlorite $Ca(OCl)_2$ and basic calcium chloride,

$$Cl_2.Ca(OH)_2.H_2O.$$

Preparation: It is prepared by passing chorine over slaked lime

$$3Ca(OH)_2 + 2Cl_2 \rightarrow \underbrace{Ca(OCl)_2 + CaCl_2.Ca(OH)_2.H_2O + H_2O}_{\text{Bleaching Powder}}$$

Properties:

(i) Reaction with dilute acids: with dilute acids, it gives chlorine which is known as available chlorine.

$$CaOCl_2 + 2HCl \longrightarrow CaCl_2 + H_2O + Cl_2 \uparrow$$

$$CaOCl_2 + H_2SO_4 \longrightarrow CaSO_4 + H_2O + Cl_2 \uparrow$$

(ii) When treated with water it decomposes into calcium chloride and calcium hypochlorite.

$$2CaOCl_2 + H_2O \longrightarrow CaCl_2 + Ca(OCl)_2 + H_2O$$

(iii) Bleaching powder reacts with CO_2 (atmospheric) and gives chlorine which accounts for its oxidizing and bleaching actions.

$$CaOCl_2 + CO_2 \longrightarrow CaCO_3 + Cl_2 \uparrow$$

(iv) Action of heat: On heating bleaching powder gives a mixture of chlorate and chloride.

$$6CaOCl_2 \xrightarrow{\Delta} Ca(ClO_3)_2 + 5CaCl_2$$

10.7.7 Biological Importance of Magnesium and Calcium

Magnesium and calcium ions play important role in biological processes. An adult body contains about 25 g of Mg and 1200 g of Ca compared with only 5 g of Fe and 0.06 g of Cu. the daily requirement of Ca^{2+} and Mg^{2+} ions is 200-300 mg. Mg^{2+} ions are concentrated animal cells and Ca^{2+} ions are concentrated in the body fluids outside the cell, in much the same way as K^+ concentrates inside the cell and Na^+ outside.

Mg^{2+} ions catalyse a number of enzymatic reactions. We know that energy is stored in form of phosphate linkages in ATP. The formation of these linkages, i.e., storage of energy is catalysed by Mg^{2+} ions. Conversely, hydrolysis of phosphate linkages is accompanied by release of energy. This process is also catalysed by Mg^{2+} ions.

Mg^{2+} ions are present is chlorophyll –a green colouring pigment in plants which absorbs light and is essential for photosynthesis.

About 99% of body calcium is present in bones and teeth as apatite, $Ca_3(PO_4)_2$ and the enamel on teeth as fluoropatite $[3\,Ca_3(PO_4)_2.CaF_2]$ Ca^{2+} ions are also important in blood clotting and are required to trigger the contraction of muscles and to maintain the regular beating of the heart. The concentration of Ca^{2+} ions in blood plasma is about 100 mg L^{-1}. This concentration is maintained by two hormones called **calcitonin** and **parathyroid.** Do you know that bone is not an inert and unchanging substance? The calcium ions in bones are continuously dissolving into and redepositing from the blood plasma to the extent of 400 mg per day in man. In normal adult this exchange is in balance, but in elderly people, especially women, there is sometimes net loss of bone calcium, leading to the diseases called **osteoporosis.**

Quick Chapter Recap

SOME KEY FACTS TO REMEMBER

- The general electronic configuration of s-block elements is (Noble gas) ns^2.
- Alkali metals are soft having low melting and boiling points. This is due to weak intermetallic bonding.
- All alkali metals are paramagnetic but their salts are diamagnetic.
- Alkali metals give characteristic flame colourations due to their low ionization enthalpies.
- Alkali metals (except Li) exhibit photoelectric effect.
- Alkali metals are strong reducing agents and reducing character increases down the group (except Li, which is strongest reducing agent due to its high oxidation potential).
- Alkali metal cations (Li^+, Na^+, K^+…) are diamagnetic because they have no unpaired electron but alkali metals are paramagnetic because they have one unpaired electron.
- Basic character : LiOH<NaOH<KOH<RbOH<CaOH
- Solution of alkali metals in liquid NH_3 are highly conducting and deep blue in colour. The blue colour is due to the presence of ammoniated electrons.
- Li forms oxide (Li_2O), Na forms peroxide (Na_2O_2) and other alkali metals form superoxides also (MO_2 : M = K, Rb, Ca).
- Alkali metals are dept under kerosene oil because they readily combine with oxygen and moisture of the air.
- Ionic mobility in water : $Li^+<Na^+<K^+<Rb^+<Ca^+$ because of degree of hydration.
- The elements showing diagonal relationship, Li-Mg, Be-Al, B-Si.
- Alkaline earth metals also (except Be and Mg) give characteristic colour to flame.
- Basic character : $Be(OH)_2<Mg(OH)_2<Ca(OH)_2<Sr(OH)_2<Ba(OH)_2$.
- $Mg(OH)_2$ is called milk of magnesia and is used as antacid.
- $Ca(OH)_2$ is called milk of lime.
- Solubility decreases : $BeSO_4>MgSO_4>CaSO_4>SrSO_4>CsSO_4$;
  $BaCO_3>MgCO_3>CaCO_3>SrCO_3>BaCO_3$
- $BeCl_2$ is covalent while $MgCl_2$ and $CaCl_2$ are ionic.
- $BeCl_2$ and $AlCl_3$ behave as strong Lewis acids and are soluble in organic solvents.
- Both Be and Al form fluoro complex ions i.e. BeF_4^{2-} and AIF_4 in solution.
- Quicklime is Cao. On adding water to it, calcium hydroxide is formed. This process is called slaking of lime.
- When CO_2 is passed through lime water, milkiness is formed (due to $CaCO_3$) and on further passing CO_2 gas, milkiness disappears [due to $Ca(HCO_3)_2$].
- Plaster of Paris is $CaSO_4$. $1/2H_2O$. On mixing with water it forms a plastic mass which sets into a hard solid mass (gypsum). This is called setting of Plaster of Paris.
- $2CaSO_4.\ H_2O + 3H_2O \rightarrow \underset{\text{Gypsum}}{2CaSO_4.2H_2O}$
- Na^+, K^+, Mg^{2+} and Ca^{2+} are found in large proportions in biological fluids. These ions perform important biological functions such as maintenance of ion balance and nerve impulse conduction.