The s-Block Elements

Alkali Metals

- 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.
- The electronic configurations of the alkali metals are given in the following table.

Element	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)	$[Xe]\ 6s^1$
Francium (Fr)	$[Rn]\ 7s^1$

Atomic and Ionic Radii

- Increase with increase in atomic number (i.e. on moving down the group)

Ionization Enthalpy

- Low

- Decreases on moving down the group, which is because of the following reasons:

- Increase in size
- Increase in screening of outermost electrons from the nuclear charge

Hydration Enthalpy

- Decreases on moving down the group

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

- They and their salts impart characteristic colour to an oxidising flame. Reason: after getting heat from the flame, the valence electrons get excited to higher energy level.

  Then they come back to ground state by emitting radiation in the visible region.

- Crimson red − Lithium (Li)
- Golden yellow − Sodium (Na)
- Pale violet − Potassium (K)

Chemical Properties

- **Highly reactive because of**

- large size
- low ionization enthalpy
- Reactivity increases down the group

- **Reaction with air:**

- Tarnish in dry air
- Reason: formation of oxides and then hydroxides with moisture

- Burns vigorously in air

$$\begin{aligned}&4Li + O_2 \longrightarrow 2Li_2O \text{ (monoxide)}\\&2Na + O_2 \longrightarrow Na_2O_2 \text{ (Peroxide)}\\&M + O_2 \longrightarrow MO_2 \quad \text{(superoxide)}\\&\qquad\qquad\qquad\quad (M = K, Rb, Cs)\end{aligned}$$

  - Kept in kerosene oil − Reason: highly reactive towards air and water
  - Li reacts directly with nitrogen of air to form lithium nitride (Li_3N).

- **Reaction with water:**

$$2M + 2H_2O \longrightarrow 2M^+ + 2OH^- + H_2$$

(M = alkali metal)

- **Reaction with acids:**

- Reacts vigrously with even dil. acids to give out hydrogen gas.

- $2M + 2HCl \rightarrow 2MOH$

- **Reaction with dihydrogen:**

$$2M + H_2 \longrightarrow 2MH \qquad \text{(hydride)}$$

  - Alkali metal hydrides are ionic solids with high melting point.
  - Reaction with acids: React voilently with dilute acids

$2\ M + 2\ HCl \rightarrow 2\ MCl + H_2$

- **Reaction with halogen:**

- Reacts vigorously to form halides which are ionic in nature (exception: LiX is some what covalent − Reason: High polarizing capability of Li^+ ion)
- LiI is the most covalent.

- **Reducing nature**

- Strong reducing agent

- Form unipositive ions easily, M(g) → M + (g) + e^-
- Li is the strongest and Na is the weakest.

- **Solutions in liquid ammonia**

- Dissolve in liquid ammonia to give conducting deep blue solution

- $M + (x + y)NH_3 \longrightarrow [M(NH_3)_x][e(NH_3)_y]^-$

- Reason for blue colour − The ammoniated electron absorbs energy in the visible region of light.

- The solution is paramagnetic in nature and liberates hydrogen to form amide.
- $M^+_{(am)} + e^- + NH_{3(l)} \longrightarrow MNH_{2(am)} + \frac{1}{2}H_{2(g)}$ (‘am’ stands for solution in ammonia)

- Its concentrated solution is bronze in colour and diamagnetic in nature.

- The metals also react with alcohol, gaseous ammonia, and alkynes

Uses

- In making useful alloys. For example, it is used with lead in making ‘white metal’ bearings for motor engines

- Li is used in making electrochemical cells.

- Na is used in making Na/Pb alloy, which are used to make organolead compounds − $PbEt_4$ and $PbMe_4$. These compounds were earlier used as anti-knock additives to petrol.

- Liquid Na is used as a coolant in fast breeder nuclear reactors.

- K plays an important role in biological systems

- KCl − used as a fertilizer

- KOH − in the manufacture of soft soap and as an absorbent of CO_2

- Cs is used in devising photoelectric cells.

Oxides and Hydroxides of Alkali Metals

- Alkali metals form oxides on combustion.

- Li − Lithium oxide (Li_2O) (and some lithium peroxide Li_2O_2)
- Na − Sodium peroxide (Na_2O_2) (and some sodium superoxide NaO_2)
- K, Rb, Cs − Superoxides, MO_2
- 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.
- Oxides are hydrolysed by water to form hydroxides.

$$M_2O + H_2O \longrightarrow 2M^+ + 2OH^-$$
$$M_2O_2 + 2H_2O \longrightarrow 2M^+ + 2OH^- + H_2O_2$$
$$2MO_2 + 2H_2O \longrightarrow 2M^+ + 2OH^- + H_2O_2 + O_2$$

- Oxides and peroxides − Colourless when pure

  Superoxides − Yellow or orange

- Na_2O_2 − Oxidizing agent in inorganic chemistry
- Superoxide O_2^- is paramagnetic.
- Reason − Presence of one unpaired electron in π*2*p* molecular orbital
- Hydroxides obtained by reaction of oxides with water − White crystalline solids
- Hydroxides of alkali metals are the strongest of all the bases.
- Hydroxides dissolve freely in water with evolution of heat.

Halides

- Preparation:

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

- Physical properties:

- Colourless crystalline solids
- Have high melting points
- Trend in melting and boiling point:

  Fluoride > Chloride > Bromide > Iodide

- Soluble in water
- LiF is less soluble.

  Reason: High lattice enthalpy

- CsI is less soluble.

  Reason: Low hydration enthalpy of the ions

- High negative enthalpies of formation ($\Delta_f H^\ominus$)
- Trend:

  For fluorides − $\Delta_f H^\ominus$ is less negative down the group.

  For chlorides, bromides, and iodides − $\Delta_f H^\ominus$ is more negative down the group.

  For a given metal −

$$\text{Fluoride} \rightarrow \text{Chloride} \rightarrow \text{Bromide} \rightarrow \text{Iodide}$$

$$\xrightarrow[\Delta_f H^\ominus \text{ becomes less negative}]{}$$

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

Examples: Carbonic acid (H_2CO_3), sulphuric acid (H_2SO_4), nitric acid (HNO_3), phosphoric acid (H_3PO_4)

- Preparation:

By the reaction of alkali metals with oxo-acids

- Physical properties
- Soluble in water
- Thermally stable: Carbonates M_2CO_3; and in most cases, hydrogen carbonates ($MHCO_3$) are highly thermally stable

Exception (I) − Li_2CO_3 is less stable to heat.

Reason: Small size of Li − polarization of large CO_3^{2-} ion − formation of more stable Li_2O and CO_2

Exception (II) − $LiHCO_3$ does not exist as solid.

- Stability of carbonates and hydrogen carbonates increases down the group.

Reason: The electropositive character increases down the group.

Anomalous Properties of Lithium

- Reasons for anomalous behaviour of lithium:
- The Li atom and ion are exceptionally small in size.
- Li has high polarizing power. [charge to radius ratio (charge/ radius)]
- Lithium compounds are soluble in organic solvents due to covalent character of lithium compounds because of the above reasons.
- **Points of difference between lithium and alkali metals**
- Li is harder and has high melting and boiling points.

- Strongest reducing agent, forms Li_2O and Li_3N by reacting with air
- $LiCl$ is deliquescent and crystallises as a hydrate, $LiCl.\ 2H_2O$.
- $LiHCO_3$ does not exist as solid.
- Does not form ethynide with ethyne
- On heating, lithium nitrate decomposes to oxide while nitrates of other alkali metals decompose to nitrites.

$$4LiNO_3 \longrightarrow 2Li_2O + 4NO_2 + O_2$$
$$2NaNO_3 \longrightarrow 2NaNO_2 + O_2$$

- Fluoride and oxide are less soluble in water.
- Li shows diagonal relationship to Mg.
- **Points of similarities between Li and Mg**
- Both are harder and lighter than the other elements in the respective groups.
- Both react slowly with water; oxides of both are less soluble in water; hydroxides of both decompose on heating.
- Both form nitrides (Li_3N and Mg_3N_2) by directly reacting with nitrogen.
- Both do not form superoxide.
- On heating, carbonates of both give oxides and CO_2; hydrogen carbonates of both do not exist as solid.
- Chlorides of both ($LiCl$, $MgCl_2$) are soluble in ethanol, deliquescent, and crystallize as hydrates ($LiCl.2H_2O$, $MgCl_2.8H_2O$).
- Reason for similarities between Li and Mg − Similar size

Some Important Compounds of Sodium

Washing Soda or Sodium Carbonate ($Na_2CO_3.10H_2O$)

- **Preparation**
- Solvay process

$$2NH_3 + H_2O + CO_2 \longrightarrow (NH_4)_2CO_3$$
$$(NH_4)_2CO_3 + H_2O + CO_2 \longrightarrow 2NH_4HCO_3$$
$$NH_4HCO_3 + NaCl \longrightarrow NH_4Cl + NaHCO_3$$
$$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + CO_2 + H_2O$$

Recovering of NH_3

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

Solvay process is not applied to manufacture K_2CO_3.

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

- **Properties**
- White crystalline solid
- Easily soluble in water
- Loses water of crystallization on heating

$$Na_2CO_3.10H_2O \xrightarrow{373\text{ K}} Na_2CO_3.H_2O + 9H_2O$$
$$Na_2CO_3.H_2O \xrightarrow{>373\text{ K}} Na_2CO_3 + H_2O$$

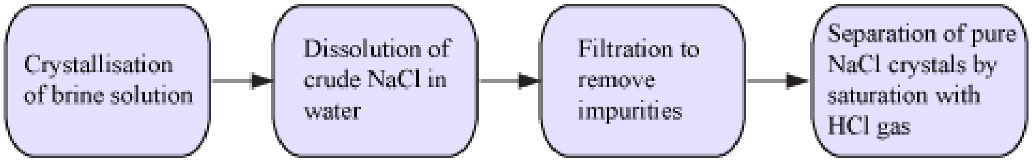
- Carbonate part is hydrolysed by water, resulting in alkaline solution.

$$CO_3^{2-} + H_2O \longrightarrow HCO_3^{-} + OH^{-}$$

- **Uses**
- In water softening
- In the manufacture of glass, soap, borax, and caustic soda
- In industries in the manufacture of paper, paints, and textile
- In laboratory as a reagent

Sodium Chloride (NaCl)

- **Preparation**



Impurities − Sodium sulphate, calcium sulphate, calcium chloride, magnesium chloride

- **Physical Properties**
- Melts at 1081 K
- Solubility is 36 g in 100 g of water at 273 K.
- With the increase in temperature, solubility does not increase.
- **Uses**
- For domestic purpose as common salt or table salt
- In the preparation of Na_2O_2, NaOH, and Na_2CO_3

Caustic Soda or Sodium Hydroxide (NaOH)

- **Preparation**
- By electrolysis of NaCl in Castner − Kellner cell

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

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

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

- **Properties**
- White translucent solid
- Gives alkaline solution by dissolving in water
- Crystals are deliquescent.

- At the surface of NaOH solution, Na_2CO_3 is formed by the reaction with CO_2 in the atmosphere.
- **Uses**
- In petroleum refining process
- In the manufacture of soap paper, artificial silk, and a number of chemicals
- In the process of purifying bauxite
- In textile industries to mercerise cotton fabrics
- In the preparation of pure fats and oils
- As a reagent in laboratory

Baking Soda or Sodium Hydrogen Carbonates ($NaHCO_3$)

- **Called Baking Soda**

Reason − It decomposes, producing bubbles of CO_2 on heating. This is the reason why holes are formed in cakes or pastries and as a result, they are light and fluffy.

- **Preparation**

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

Uses

- As an antiseptic for skin infection
- For extinguishing fire

Biological Importance of Na and K

- Sodium ions take part
- in the transmission of nerve signals
- in regulating the water flow across cell membranes
- in the transportation of sugars and amino acids into cells

- Potassium ions participate in
- activating enzymes in cell fluids
- the oxidation of glucose to produce ATP
- the transmission of nerve signals

Alkaline Earth Metals

- Group 2 elements: Beryllium, magnesium, calcium, strontium, barium and radium.
- They are called alkaline earth metals because their oxides and hydroxides are alkaline in nature, and their oxides are found in earth's crust.

Electronic Configuration

- General outer electronic configuration is ns^2
- The electronic configurations of the alkaline earth metals are given in the following table:

Element	Electronic Configuration
Beryllium (Be)	$1s^2 2s^2$
Magnesium (Mg)	$1s^2 2s^2 2p^6 3s^2$
Calcium (Ca)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Strontium (Sr)	$[Kr]\ 5s^2$
Barium (Ba)	$[Xe]\ 6s^2$

Radium (Ra)	[Rn] $7s^2$

Atomic and ionic radii

- Smaller than the corresponding alkali metals

  Reason: Greater nuclear charge

- Increase with the increase in atomic number (i.e., on moving down the group)

Ionisation Enthalpy

- Low
- Decreases on moving down the group. This is because of:
- Increase in size
- Increase in the screening of valence electrons from the nuclear charge
- Higher first ionisation enthalpy than that of the corresponding alkali metals. This is due to smaller size and higher nuclear charge as compared to the corresponding group 1 alkali metals.
- Second ionisation enthalpy is lower than that of the corresponding alkali metals

Hydration enthalpy

- Decreases on moving down the group (i.e. with increase in ionic size)
- Larger hydration enthalpy than that of the corresponding alkali metal ions. This is why compounds of alkaline earth metals are more extensively hydrated than those of alkali metals.
- Example**:** NaCl is not hydrated whereas $MgCl_2$ exists as $MgCl_2.6H_2O$. KCl does not form hydrate whereas $CaCl_2$ exists as $CaCl_2 \cdot 6H_2O$.

Physical properties

- Silvery white, lustrous and soft (harder than alkali metals) (Exception − Beryllium and Magnesium are greyish)

- Melting and boiling points of these metals are higher than those of the corresponding alkali metals.

  Reason: Smaller size than alkali metals

- Strongly electropositive in nature. Reason: Low ionisation enthalpies

- Flame test:

- Calcium − Brick red
- Strontium − Crimson
- Barium − Apple green
- Mg and Be do not impart colour because the electrons are too strongly bound to be excited.

- Good conductors of heat and electricity

Chemical Properties

- Less reactive than alkali metals

- Reactivity increases down the group

- Reaction with air and water:

- Be and Mg are kinetically inert to oxygen and water. Reason − formation of oxide film on their surface

- However, when Mg and powdered Be burn, they give MgO and Mg_3N_2, and BeO and Be_3N_2 respectively.

- Ca, Sr and Ba react readily with air to form oxides and nitrides; also react with water (even cold), to form hydroxides

- Reaction with halogens:

  - At high temperature

$M + X_2 \rightarrow MX_2$ (Halides)

(X = F, Cl, Br, I)

- BeF_2 is prepared by the thermal decomposition of $(NH_4)_2BeF_4$
- $BeCl_2$ is prepared from BeO by heating it with carbon and chlorine to about 600-900 K

$BeO + C + Cl_2 \rightleftharpoons BeCl_2 + CO$

- Reaction with hydrogen:

- React with hydrogen to form hydrides (Exception − Beryllium)
- Preparation of BeH_2 from $BeCl_2$ and $LiAlH_4$ −

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

- Reaction with acids:

$M + 2HCl \rightarrow MCl_2 + H_2$

(M = Alkali earth metal)

- Reducing nature:

- Strong reducing agent (weaker than corresponding alkali metals)
- Be is a weaker reducing agent than other alkaline earth metals.

- Solutions in liquid ammonia:

Dissolve in liquid ammonia to give deep, blue-black solution

$M + (x + y)NH_3 \rightarrow [M(NH_3)_x]^{2+} + 2[e(NH_3)_y]^-$

Uses

- **Beryllium:**

- For making alloys such as Cu−Be alloys which are used for making high strength springs
- Metallic Be, for making windows of X-ray tubes

- **Magnesium:**

- For making alloys with Al, Zn, Mn and Sn

- Mg−Al alloys are used in the construction of aircraft

- Magnesium is used in flash powders and bulbs, incendiary bombs and signals

- Milk of magnesia (suspension of $(MgOH)_2$ in water) is used as an antacid

- $MgCO_3$ is a component in toothpaste

- **Calcium:**
- In the process of extraction of metals from oxides which are difficult to reduce with carbon
- For removing air from vacuum tubes. Reason − reacts with oxygen and nitrogen at high temperature
- **Barium:**
- For removing air from vacuum tubes. Reason − reacts with oxygen and nitrogen at high temperature
- **Radium:**
- In radiotherapy (e.g., in the treatment of cancer)

General characteristics of compounds of the alkaline earth metals

- Compounds of alkaline earth metals are ionic in nature (but less ionic than compounds of corresponding alkali metals).
- Reason − Greater nuclear charge, smaller size
- Compounds of lighter elements (Be, Mg) are more covalent than those of heavier and large elements (Ca, Sr, Ba).

Oxides and hydroxides

- **Oxides**
- Burn in oxygen forming metal monoxide (MO)
- MO has rock salt structure. (Exception − BeO)
- Thermally stable − Reason: High enthalpies of formation
- Ionic in nature (Exception − BeO is amphoteric)
- React with water to form hydroxides, which are sparingly soluble

$$MO + H_2O \longrightarrow M(OH)_2$$

- **Hydroxide**

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

- Basic in nature (less basic than alkali metal hydroxide)

  Exception − $Be(OH)_2$ is amphoteric.

  $$\mathrm{Be(OH)_2} + \underset{\text{Base}}{\mathrm{2OH^-}} \longrightarrow \underset{\text{Beryllate ion}}{\mathrm{[Be(OH)_4]^{2-}}}$$

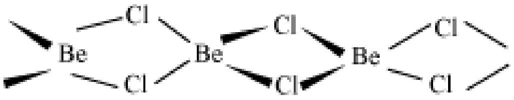
  $$\mathrm{Be(OH)_2} + \underset{\text{Acid}}{\mathrm{2HCl}} + \mathrm{2H_2O} \longrightarrow \mathrm{[Be(OH)_4]Cl_2}$$

- Less stable than alkali metal hydroxide

  Halides

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

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

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

- Tendency to form halide hydrates decreases down the group.

  Example − $MgCl_2 \cdot 8H_2O$, $CaCl_2 \cdot 6H_2O$, $SrCl_2 \cdot 6H_2O$, $BaCl_2 \cdot 2H_2O$

  Salts of oxo-acids

- Carbonates

- Insoluble in water

- Precipitated by adding sodium or ammonium carbonate

- Solubility decreases down the group.

- When heated, decompose to given oxide and CO_2

- Thermal stability increases down the group.
- $BeCO_3$ is unstable and hence, kept in the atmosphere of CO_2.
- Sulphate
- White solids
- Thermally stable
- Solubility decreases down the group.
- $BeSO_4$ and $MgSO_4$ are readily soluble in water. Reason − High hydration enthalpy of Be^{2+} and Mg^{2+}
- Nitrates
- Preparation: By dissolving carbonates in dilute nitric acid
- Tendency to form hydrates decreases down the group. Reason − Increase in size, decrease in hydration enthalpy

  Example − $Mg(NO_3)_2.6H_2O$, $Ba(NO_3)_2$
- When heated, decompose to give oxide

$$2M(NO_3)_2 \longrightarrow 2MO + 4NO_2 + O_2$$

Anomalous Properties of Beryllium

- Exceptionally small size of the atom and the cation
- Forms covalent compounds (due to small size and high ionisation enthalpy)
- Highest coordination number is four while other members exhibit six.
- Reason: Be has four orbitals in its valence shell while other members can use *d*-orbitals
- Unlike other elements, the oxides and hydroxides of Be are amphoteric in nature.

Diagonal Relationship Between Be and Al

- Reason: Same charge/radius ratio

- Similarities:
- Like Al, it is not readily attached by acids; (Reason − presence of an oxide film on the surface of the metal)
- Both dissolve in excess of alkali, giving beryllate ion $[Be(OH)_4]^{2-}$ and aluminate ion $[Al(OH)_4]^-$.
- Like Al_2Cl_6, $BeCl_2$ has a Cl^- bridged polymeric structure. Both these compounds are Lewis acids, and are soluble in organic solvents.
- Both form fluoro complex ions, $[BeF_4]^{2-}$, $[AlF_6]^{3-}$

Some Important Compounds of Calcium

Quick Lime or Calcium Oxide (CaO)

- **Preparation:**

  In rotary kiln,

  $$\underset{\text{lime stone}}{CaCO_3} \xrightarrow{1070-1270K} CaO + CO_2$$

  For the completion of the reaction, CO_2 is removed as soon as it is produced.

- **Properties:**
- White amorphous solid
- Melts at 2870 K
- Absorbs moisture and CO_2 when exposed to atmosphere

  $$CaO + H_2O \longrightarrow Ca(OH)_2$$
  $$CaO + CO_2 \longrightarrow CaCO_3$$

- Basic oxide − combines with acidic oxides on heating

  $$CaO + SiO_2 \longrightarrow CaSiO_3$$
  $$6CaO + P_4O_{10} \longrightarrow 2Ca_3(PO_4)_2$$

- Slaking of lime − the process of breaking lump of lime by adding a limited amount of water
- When slaked with soda, it gives solid soda lime
- **Uses:**
- For manufacturing cement
- For manufacturing sodium carbonate from caustic soda
- For manufacturing dye stuffs
- For purifying sugar
- It is the cheapest form of alkali

Slaked Lime or Calcium Hydroxide [$Ca(OH)_2$]

- **Preparation:**

By adding water to quick lime

$$\underset{\text{Quick lime}}{CaO} + H_2O \longrightarrow Ca(OH)_2$$

- **Properties:**
- White amorphous powder
- Sparingly soluble in water
- Lime water − clear aqueous solution of $Ca(OH)_2$
- Milk of lime − suspension of slaked lime in water
- With chlorine, slaked lime gives bleaching powder

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

- On passing carbon dioxide through lime water, it turns milky due to the formation of calcium carbonate ($CaCO_3$)

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

- On passing excess of CO_2, calcium carbonate ($CaCO_3$) dissolves, forming calcium hydrogen carbonate

$$\underset{\text{Calcium hydrogencarbonate}}{CaCO_3 + CO_2 + H_2O \longrightarrow Ca(HCO_3)_2}$$

- **Uses:**
- For the preparation of mortar, a building material
- Disinfectant in nature, and hence, used in white washing
- For the manufacture of glass
- For the preparation bleaching powder
- For purifying sugar
- In the tanning industry

Calcium Carbonate ($CaCO_3$)

- **Naturally Occurs As Limestone, Chalk, Marble**
- **Preparation:**
- By passing CO_2 through slaked lime

On passing excess of CO_2, $CaCO_3$ dissolves to form water-soluble calcium hydrogen carbonate.

$$CaCO_3 + CO_2 + H_2O \longrightarrow Ca(HCO_3)_2$$

- By adding aqueous solution of sodium carbonate to calcium chloride

$$Na_2CO_3 + CaCl_2 \longrightarrow CaCO_3 + 2NaCl$$

- **Properties:**

- White fluffy powder
- Insoluble in water
- Decomposes to CO_2 on heating

$$CaCO_3 \xrightarrow{1200K} CaO + CO_2$$

- Reacts with dilute acids to liberate CO_2

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$
$$CaCO_3 + H_2SO_4 \longrightarrow CaSO_4 + H_2O + CO_2$$

- **Uses:**
- As a building material in the form of marble
- For manufacturing quick lime
- As a flux with $MgCO_3$ in the extraction of metals such as iron
- Precipitated $CaCO_3$ for the manufacture of high quality paper
- As antacid, mild abrasive in toothpaste, in chewing gums, and as filler in cosmetics

Plaster Of Paris or Calcium Sulphate $\left(CaSO_4 \cdot \frac{1}{2}H_2O\right)$

- **Hemihydrate of Calcium Sulphate**
- **Preparation:**

By heating gypsum to 393 K

$$\underset{\text{Gypsum}}{2(CaSO_4 \cdot 2H_2O)} \longrightarrow 2(CaSO_4) \cdot \frac{1}{2}H_2O + 3H_2O$$

Above 393 K, water is liberated to form anhydrous calcium sulphate ($CaSO_4$), known as ‘dead burnt plaster’.

- **Property:**

On mixing with water, it forms a plastic mass which sets into a hard solid in 5 − 15 minutes.

- **Uses:**
- As plaster and as building material
- In surgical bandages, for setting broken and fractured bones
- In dentistry
- For making statues, busts, and in ornamental work

Cement

- Also called Portland cement
- Reason: Resembles the natural limestone extracted in the Isle of Portland, England
- Composition of cement:

$CaO = 50 - 60\%$

$SiO_2 = 20 - 25\%$

$Al_2O_3 = 5 - 10\%$

$MgO = 2 - 3\%$

$Fe_2O_3 = 1 - 2\%$

$SO_3 = 1 - 2\%$

- For a good-quality cement:

The ratio of $SiO_2: Al_2O_3 = 2.5$ to 4

The ratio of $CaO: (SiO_2 + Al_2O_3 + Fe_2O_3) = 2$

- Manufacture of cement:
- Raw materials − limestone and clay

On strong heating, clay and lime fuse and react to form ‘Cement clinker’.
↓
‘Cement clinker’ is mixed with 2 − 3% by weight of gypsum to form cement.

- Important ingredients present in Portland cement:

  Dicalcium silicate (Ca_2SiO_4) = 26%

  Tricalcium silicate (Ca_3SiO_5) = 51%

  Tricalcium aluminate ($Ca_3Al_2O_6$) = 11%

- Setting of cement:

  On mixing with water, cement sets to give a hard solid

- Reason − hydration of the molecules of the constituents and their rearrangement

- Reason for adding gypsum − to slow down the process of setting of the cement; this results in sufficient hardening of the cement

- Uses:

  As a construction material −

- In concrete and reinforced concrete

- In plastering

- In construction of buildings, bridges and dams

Biological Importance of Magnesium and Calcium

- Amount of Mg in the human body = 25 g (approx)

  Amount of Ca in the human body = 1200 g (approx)

- Daily requirement of Ca and Mg in the human body = 200 − 300 mg
- Importance of Mg:
- Required as cofactor by the enzymes that utilise ATP in phosphate transfer
- Present in chlorophyll
- Importance of Ca:
- Bones and teeth contain about 99% body calcium.
- Plays important roles in − neuromuscular function, inter-neuronal transmission, cell membrane integrity, blood coagulation
- The calcium concentration in plasma is maintained at 100 mgL^{-1} by two hormones −
- Calcitonin
- Parathyroid hormone
- Bone is not an inert and unchanging substance. It is continuously being solubilised and re-deposited to the extent of 400 mg per day in humans.