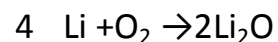


s-block elements

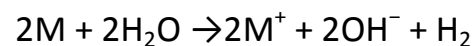
- All the alkali metals have one valence electron.
- They readily lose electron to give monovalent M^+ ions.
- The alkali metal atoms have the largest sizes in a particular period of the periodic table.
- The monovalent ions (M^+) are smaller than the parent atom.
- The atomic and ionic radii of alkali metals increase on moving down the group.
- The ionization enthalpies of the alkali metals are considerably low and decrease down the group because the effect of increasing size outweighs the increasing nuclear charge, and the outermost electron is very well screened from the nuclear charge.
- The hydration enthalpies of alkali metal ions decrease with increase in ionic sizes.
- **Physical Properties**
 - silvery white, soft and light metals
 - Low density which increases down the group from Li to Cs.
 - Melting and boiling points of the alkali metals are low indicating weak metallic bonding.
 - The alkali metals and their salts impart characteristic colour to an oxidizing flame. This is because the heat from the flame excites the outermost orbital electron to a higher energy level. When the excited electron comes back to the ground state, there is emission of radiation.
 - Detected by the respective flame tests and can be determined by flame photometry or atomic absorption spectroscopy.

- **Chemical Properties**

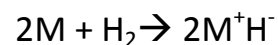
- **Reactivity towards air:** The alkali metals tarnish in dry air due to the formation of their oxides which in turn react with moisture to form hydroxides. They burn vigorously in oxygen forming oxides.



- **Reactivity towards water:** The alkali metals react with water to form hydroxide and dihydrogen.



- **Reactivity towards dihydrogen:** The alkali metals react with dihydrogen at about 673K (lithium at 1073K) to form hydrides.



- **Reactivity towards halogens :** The alkali metals readily react vigorously with halogens to form ionic halides, M^+X^- .

- **Reducing nature:** The alkali metals are strong reducing agents, lithium being the most and sodium the least powerful.

- **Solutions in liquid ammonia:** The alkali metals dissolve in liquid ammonia giving deep blue solutions which are conducting in nature.

- **ANOMALOUS PROPERTIES OF LITHIUM**

The anomalous behaviour of lithium is due to the : (i) exceptionally small size of its atom and ion, and (ii) high polarising power.

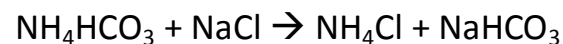
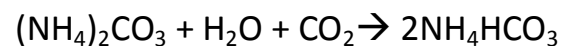
- **Points of Difference between Lithium and other Alkali Metals**

- Lithium is much harder. Its m.p. and b.p. are higher than the other alkali metals.
- Lithium is least reactive but the strongest reducing agent among all the alkali metals. On combustion in air it forms mainly monoxide, Li_2O and the nitride, Li_3N unlike other alkali metals.
- LiCl is deliquescent and crystallises as a hydrate, $\text{LiCl}\cdot 2\text{H}_2\text{O}$ whereas other alkali metal chlorides do not form hydrates.
- Lithium hydrogencarbonate is not obtained in the solid form while all other elements form solid hydrogencarbonates.
- Lithium unlike other alkali metals forms no ethynide on reaction with ethyne.

- Lithium nitrate when heated gives lithium oxide, Li_2O , whereas other alkali metal nitrates decompose to give the corresponding nitrite.
- LiF and Li_2O are comparatively much less soluble in water than the corresponding compounds of other alkali metals.

Sodium Carbonate (Washing Soda), $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$

- Sodium carbonate is generally prepared by Solvay Process. In this process, advantage is taken of the low solubility of sodium hydrogencarbonate whereby it gets precipitated in the reaction of sodium chloride with ammonium hydrogencarbonate. The latter is prepared by passing CO_2 to a concentrated solution of sodium chloride saturated with ammonia, where ammonium carbonate followed by ammonium hydrogencarbonate are formed.



Properties

- Sodium carbonate is a white crystalline solid which exists as a decahydrate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$. This is also called washing soda. It is readily soluble in water. On heating, the decahydrate loses its water of crystallisation to form monohydrate. Above 373K, the monohydrate becomes completely anhydrous and changes to a white powder called soda ash.

Uses

- (i) It is used in water softening, laundering and cleaning.
- (ii) It is used in the manufacture of glass, soap, borax and caustic soda.
- (iii) It is used in paper, paints and textile industries.
- (iv) It is an important laboratory reagent both in qualitative and quantitative analysis.

- **Sodium Chloride, NaCl**

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 gas. Crystals of pure sodium chloride separate out.

- **Properties**

Sodium chloride melts at 1081K. It has a solubility of 36.0 g in 100 g of water at 273 K. The solubility does not increase appreciably with increase in temperature.

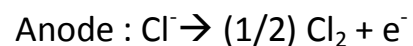
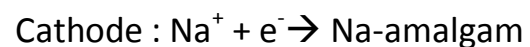
- **Uses**

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

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

- **Sodium Hydroxide (Caustic Soda), NaOH**

- Sodium hydroxide is generally prepared commercially by the electrolysis of sodium chloride in Castner-Kellner cell. A brine solution is electrolysed using a mercury cathode and a carbon anode. Sodium metal discharged at the cathode combines with mercury to form sodium amalgam. Chlorine gas is evolved at the anode.



The amalgam is treated with water to give sodium hydroxide and hydrogen gas.

- Properties

Sodium hydroxide is a white, translucent solid. It melts at 591 K. It is readily soluble in water to give a strong alkaline solution. Crystals of sodium hydroxide are deliquescent. The sodium hydroxide solution at the surface reacts with the CO₂ in the atmosphere to form Na₂CO₃.

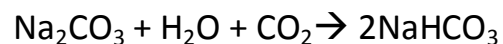
- Uses

It is used in (i) the manufacture of soap, paper, artificial silk and a number of chemicals, (ii) in petroleum refining, (iii) in the purification of bauxite, (iv) in the textile industries for mercerising cotton fabrics, (v) for the preparation of pure fats and oils, and (vi) as a laboratory reagent.

- **Sodium hydrogencarbonate or Baking soda**

Sodium hydrogencarbonate is known as baking soda because it decomposes on heating to generate bubbles of carbon dioxide.

Sodium hydrogencarbonate is made by saturating a solution of sodium carbonate with carbon dioxide. The white crystalline powder of sodium hydrogencarbonate, being less soluble, gets separated out.



Sodium hydrogencarbonate is a mild antiseptic for skin infections. It is used in fire extinguishers.

- **GROUP 2 ELEMENTS : ALKALINE EARTH METALS**

- **Electronic Configuration:** These elements have two electrons in the s -orbital of the valence shell. Their general electronic configuration may be represented as [noble gas] ns^2 . Like alkali metals, the compounds of these elements are also predominantly ionic.
- **Atomic and Ionic Radii:** The atomic and ionic radii of the alkaline earth metals are smaller than those of the corresponding alkali metals in the same periods. This is due to the increased nuclear charge in these elements. Within the group, the atomic and ionic radii increase with increase in atomic number.
- **Ionization Enthalpies:** The alkaline earth metals have low ionization enthalpies due to fairly large size of the atoms. Since the atomic size increases down the group, their ionization enthalpy decreases. The first ionisation enthalpies of the alkaline earth metals are higher than those of the corresponding Group 1 metals. This is due to their small size as compared to the corresponding alkali metals.
- **Hydration Enthalpies:** Like alkali metal ions, the hydration enthalpies of alkaline earth metal ions decrease with increase in ionic size down the group. The hydration enthalpies of alkaline earth metal ions are larger than those of alkali metal ions.

- **Physical Properties**

- The alkaline earth metals, in general, are silvery white, lustrous and relatively soft but harder than the alkali metals.
- The melting and boiling points of these metals are higher than the corresponding alkali metals due to smaller sizes. The trend is, however, not systematic. Because of the low ionization enthalpies, they are strongly electropositive in nature. The electropositive character increases down the group from Be to Ba.
- The alkaline earth metals like those of alkali metals have high electrical and thermal conductivities which are typical characteristics of metals.

- **Chemical Properties**

- **Reactivity towards the halogens:** All the alkaline earth metals combine with halogen at elevated temperatures forming their halides.
- **Reactivity towards hydrogen:** All the elements except beryllium combine with hydrogen upon heating to form their hydrides, MH_2 . BeH_2 , however, can be prepared by the reaction of $BeCl_2$ with $LiAlH_4$.
- **Reactivity towards acids:** The alkaline earth metals readily react with acids liberating dihydrogen.

- Reducing nature: Like alkali metals, the alkaline earth metals are strong reducing agents. This is indicated by large negative values of their reduction potentials. However their reducing power is less than those of their corresponding alkali metals.
- **Solutions in liquid ammonia:** Like alkali metals, the alkaline earth metals dissolve in liquid ammonia to give deep blue black solutions forming ammoniated ions.

- **ANOMALOUS BEHAVIOUR OF BERYLLIUM**

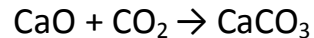
- Beryllium has exceptionally small atomic and ionic sizes and thus does not compare well with other members of the group. Because of high ionisation enthalpy and small size it forms compounds which are largely covalent and get easily hydrolysed.
- Beryllium does not exhibit coordination number more than four as in its valence shell there are only four orbitals. The remaining members of the group can have a coordination number of six by making use of d-orbitals.
- The oxide and hydroxide of beryllium, unlike the hydroxides of other elements in the group, are amphoteric in nature.
-

- **Calcium Oxide or Quick Lime, CaO**

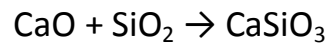
- It is prepared on a commercial scale by heating limestone (CaCO₃) in a rotary kiln at 1070-1270 K.



- The carbon dioxide is removed as soon as it is produced to enable the reaction to proceed to completion.
- Calcium oxide is a white amorphous solid. It has a melting point of 2870 K. On exposure to atmosphere, it absorbs moisture and carbon dioxide.



- The addition of limited amount of water breaks the lump of lime. This process is called slaking of lime. Quick lime slaked with soda gives solid sodalime. Being a basic oxide, it combines with acidic oxides at high temperature.



- Uses:

- (i) It is an important primary material for manufacturing cement and is the cheapest form of alkali.

- (ii) It is used in the manufacture of sodium carbonate from caustic soda.

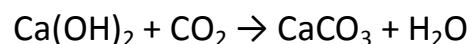
- (iii) It is employed in the purification of sugar and in the manufacture of dye stuffs.

- **Calcium hydroxide**

- Calcium hydroxide is prepared by adding water to quick lime, CaO.

- It is a white amorphous powder. It is sparingly soluble in water. The aqueous solution is known as lime water and a suspension of slaked lime in water is known as milk of lime.

- When carbon dioxide is passed through lime water it turns milky due to the formation of calcium carbonate.



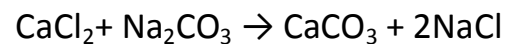
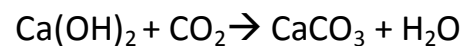
- On passing excess of carbon dioxide, the precipitate dissolves to form calcium hydrogencarbonate.



- **Uses:** (i) It is used in the preparation of mortar, a building material. It is used in white wash due to its disinfectant nature. (iii) It is used in glass making, in tanning industry, for the preparation of bleaching powder and for purification of sugar.

- **Calcium Carbonate, CaCO₃**

- Calcium carbonate occurs in nature in several forms like limestone, chalk, marble etc. It can be prepared by passing carbon dioxide through slaked lime or by the addition of sodium carbonate to calcium chloride.



Excess of carbon dioxide should be avoided since this leads to the formation of water soluble calcium hydrogencarbonate. Calcium carbonate is a white fluffy powder.

It is almost insoluble in water. When heated to 1200 K, it decomposes to evolve carbon dioxide.

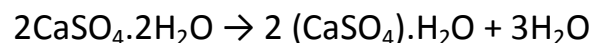
Uses:

It is used as a building material in the form of marble and in the manufacture of quick lime. Calcium carbonate along with magnesium carbonate is used as a flux in the extraction of metals such as iron.

Specially precipitated CaCO₃ is extensively used in the manufacture of high quality paper. It is also used as an antacid, mild abrasive in tooth paste, a constituent of chewing gum, and a filler in cosmetics.

- **Calcium Sulphate (Plaster of Paris), CaSO₄·½ H₂O**

- It is a hemihydrate of calcium sulphate. It is obtained when gypsum, CaSO₄·2H₂O, is heated to 393 K.



- Above 393 K, no water of crystallisation is left and anhydrous calcium sulphate, CaSO₄ is formed. This is known as 'dead burnt plaster'.
- It has a remarkable property of setting with water. On mixing with an adequate quantity of water it forms a plastic mass that gets into a hard solid in 5 to 15 minutes.

- **Uses:**The largest use of Plaster of Paris is in the building industry as well as plasters. It is used for immobilising the affected part of organ where there is a bone fracture or sprain. It is also employed in dentistry, in ornamental work and for making casts of statues and busts.