Biological importance of sodium and potassium

Sodium ions are located in blood plasma and in the interstitial fluid surrounding the cells. These ions take part in (i) transmission of nerve signals (ii) regulation of water flow across cell membranes and (iii) transport of amino acids and sugars into the cells. Potassium ions are the most abundant cations in cell fluids. They take part in (i) oxidation of glucose to form ATP (ii) activation of enzymes and (iii) transmission of nerve signals (with Na^+ ions). The concentrations of Na^+ and K^+ ions vary considerably on the opposite sides of cell membranes. This concentration gradient of ions is due to a mechanism called sodium-potassium pump that operates across cell membranes.

Important Compounds of Sodium

1. Sodium carbonate (soda ash: Na_2CO_3 washing soda: $Na_2CO_3.10H_2O$)

Sodium carbonate is prepared by the **Solvay process** or ammonia–soda process . CO_2 gas is passed through brine solution already saturated with ammonia to form sodium hydrogen carbonate or sodium bicarbonate.

$$2NH_3 + H_2O + CO_2 \longrightarrow (NH_4)_2CO_3$$

$$(NH_4)_2CO_3 + CO_2 + H_2O \longrightarrow \underset{\text{ammonium bicarbonate}}{2NH_4HCO_3}$$

$$NH_4HCO_3 + NaCl \longrightarrow \underset{\text{sodium bicarbonate}}{NaHCO_3} + NH_4Cl$$

$NaHCO_3$ so formed precipitates out due to its low solubility. Sodium bicarbonate is filtered and heated to form sodium carbonate.

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

The filterate after the removal of $NaHCO_3$ contains NH_4Cl. It is mixed with calcium hydroxide and heated to recover ammonia.

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

This NH_3 is used again. $CaCl_2$ is a by-product in the process.

It may be noted that Solvay process cannot be used for the manufacture of K_2CO_3 because $KHCO_3$ formed in the intermediate step is highly soluble in water and not precipitaed. So it cannot be separated from solution easily.

Properties

It is a colourless solid soluble in water. The crystals obtained from solution are of the decahydrate called *washing soda* ($Na_2CO_3 \cdot 10H_2O$). On heating washing soda loses the water of crystallisation to form $Na_2CO_3 \cdot H_2O$ (monohydrate). On heating above 373K, it becomes anhydrous sodium carbonate and forms a white powder called **soda ash** (Na_2CO_3). Sodium carbonate hydrolyses in water to form an alkaline solution.

Uses

(i) It is used for washing purposes in laundary.
(ii) It is used for softening of hard water.
(iii) Na_2CO_3 is used in soap, glass and paper industries.
(iv) It is used as a lab reagent.

2. Caustic soda or sodium hydroxide (NaOH)

Sodium hydroxide is manufactured by electrolysing brine solution (NaCl) using a mercury cathode and carbon anode in Castner Kellner Cell. Sodium metal discharged at the cathode combines with mercury to form sodium amalgam. Chlorine gas is evolved at the anode.

$$Na^+ + e^- \xrightarrow{Hg} Na/Hg \text{ (sodium amalgam) (cathode)}; \qquad 2Cl^- \longrightarrow Cl_2 + 2e^- \text{ (anode)}$$

The sodium amalgam formed is brought into contact with water to give sodium hydroxide and hydrogen gas.

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

Properties

Sodium hydroxide is a white transluscent solid. It is readily soluble in water and the solution is a strong alkali. Crystals of NaOH are deliquescent.

Uses

(i) Used in the manufacture of soap, paper, artificial silk and a number of chemicals.

(ii) For mercerising cotton and manufacture of rayon.

(iii) For purification of bauxite during extraction of aluminium.

(iv) It is used as a laboratory reagent.

3. Sodium Chloride (NaCl)

Common salt is generally obtained by evaporation of sea water. Crude $NaCl$ obtained by crystallisation of brine contains impurities such as $CaCl_2$ and $MgCl_2$ which make it deliquescent. To get pure $NaCl$, a solution of crude salt is prepared in minimum amount of water. It is filtered and the solution is saturated with HCl gas when $NaCl$ separates out.

It is used as table salt for cooking and also in the manufacture of Na_2O_2, $NaOH$ and Na_2CO_3.

4. Sodium Hydrogen Carbonate (Baking soda $NaHCO_3$)

It can be prepared by saturating a solution of Na_2CO_3 with CO_2 gas. $NaHCO_3$ being less soluble, precipitates out as white solid.

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

Uses

It is known as *baking soda* because it decomposes on heating to form bubbles of CO_2 which make cakes etc. light and fluffy. It is a mild antiseptic for skin infections.

ALKALINE EARTH METALS

Beryllium, magnesium, calcium, strontium, barium and radium form group 2 of the periodic table. These elements, with the exception of Be, are known as alkaline earth metals because their oxides and hydroxides are alkaline and their oxides are found in the earth's crust.

1. Occurrence

Among alkaline earth metals, Ca and Mg are abundant. Calcium is the fifth most abundant element in earth's crust. Beryllium is not very abundant while the last element radium is radioactive.

2. Electronic configuration

Their general outer electronic configuration may be written as ns^2.

3. Atomic and ionic size

The atomic and ionic radii of alkaline earth metals are smaller than those of the alkali metals in the corresponding periods due to the increased nuclear charge in these elements. The atomic and ionic radii increase with increase in atomic number down the group.

4. Ionisation enthalpies and electropositive character

The alkaline earth metals have low ionisation enthalpies as compared to *p*–block elements and its value decreases down the group. The first ionisation enthalpies of the alkaline earth metals are higher than those of the corresponding alkali metals.

5. Hydration enthalpies

The hydration enthalpies of alkaline earth metal ions are larger than those of alkali metal ions. Hence compounds of alkaline earth metals are more hydrated than those of alkali metals. For example, crystalline calcium chloride and magnesium chloride have the formulae $CaCl_2 . 6H_2O$ and $MgCl_2 . 6H_2O$ respectively while $NaCl$ and KCl do not form such hydrates.

The hydration enthalpies of alkaline earth metal ions decrease with increase in size down the group. It follows the order: $Be^{2+} > Mg^{2+} > Ca^{2+} > Sr^{2+} > Ba^{2+}$

6. Flame colouration

Calcium imparts brick red colour, strontium gives crimson red colour and barium imparts apple green colour to the bunsen flame. Be and Mg do not give any colour to the flame. This is because the electrons in *Be* and *Mg* are strongly bound to the nucleus and hence are not easily excited by the flame.

7. Melting and boiling point

The melting and boiling points of alkaline earth metals are higher than the corresponding alkali metals due to smaller sizes.

Chemical properties

Generally, reactivity of *s*-block elements increases down the group with increase in electropositive character.

(i) Reactivity towards air and water

Beryllium and magnesium do not react with air and water at ordinary temperature because of the formation of an oxide film on their surface. But powdered *Be* burns in air to form beryllium oxide (BeO) and beryllium nitride (Be_3N_2). Similarly, magnesium burns in air to form MgO and

Mg_3N_2. Ca, Ba and Sr are readily attacked by air to form the corresponding oxide and nitride.

Beryllium is unreactive with water, but Mg reacts with boiling water or steam to form the oxide. Calcium, strontium and barium react with cold water forming their hydroxides.

$$Ca + 2H_2O \longrightarrow Ca\,(OH)_2 + H_2$$

(ii) Reactivity towards hydrogen

Except Be, all the other metals of this group react with hydrogen on heating to form metal hydrides having the general formula MH_2 where M is the group 2 metal. $Be\,H_2$ can be prepared by the action of $BeCl_2$ with $LiAlH_4$.

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

(iii) Action of acids

The alkaline earth metals readily react with dilute acids liberating H_2.

(iv) Reactivity towards halogens

Group 2 metals directly combine with halogen to form divalent halides of the formula MX_2 where X is the halogen. $BeCl_2$ is however, prepared by passing Cl_2 over a hot mixture of BeO and coke.

$$BeO + C + Cl_2 \xrightarrow{600-800\,K} BeCl_2 + CO$$

(v) Reaction with liquid ammonia

Like alkali metals, alkaline earth metals dissolve in liquid ammonia to form blue solutions containing ammoniated electrons.

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

(vi) Reducing nature

Alkaline earth metals are good reducing agents. However, their reducing power is less than those of the corresponding alkali metals. Be has less negative reduction potential value compared to other metals of the group. Still it is a good reducing agent due to the large hydration energy of Be^{2+} ion.

Uses

1. Metallic beryllium is used for making windows of X-ray tubes.
2. Magnesium – aluminium alloys are used in air-craft construction.
3. A suspension of magnesium hydroxide in water (milk of magnesia) is used as an antacid in medicine.

4. Ca and Ba are used to remove air from vacuum tubes

5. Radium salts are used in radiotherapy for the treatment of cancer.

General characteristics of compounds

1. Oxides and hydroxides

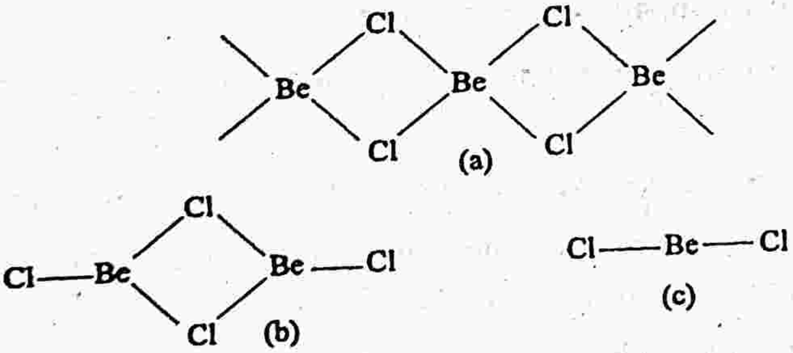
(i) ***Oxides and hydroxides:*** Alkaline earth metals burn in air or oxygen to form their oxides. BeO is covalent while all other oxides are ionic.

BeO is amphoteric in character, while the oxides of the rest of the elements in group 2 are basic. The oxides of these metals (except BeO) react with water to form their corresponding hydroxides.

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

The hydroxides of alkaline earth metals are basic except $Be(OH)_2$ which is amphoteric. The basic strength increases from $Mg(OH)_2$ to $Ba(OH)_2$. *The solubility and thermal stability of hydroxides also increase downward in the group.* $Be(OH)_2$ 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.

(ii) ***Halides:*** Beryllium halides are covalent. Other halides of these metals are ionic. Beryllium chloride is soluble in organic solvents indicating its covalent nature. In the solid state $BeCl_2$ has a polymeric chain structure involving $Be–Cl–Be$ bridge



(a) structure of solid $BeCl_2$ (b) $BeCl_2$ dimer (c) monomer

In the vapour phase $BeCl_2$ tends to form a chloro-bridged dimer which dissociates into the linear monomer at elevated temperatures of the order of 1200 K.

Halides of the other group 2 metals (except fluorides) are soluble in water. The tendency to form halide hydrates decreases down the group. e.g., $MgCl_2 . 8H_2O$, $CaCl_2 . 6H_2O$, $SrCl_2 . 6H_2O$, $BaCl_2 . 2H_2O$.