

1. Details of Module and its structure

Module Detail	
Subject Name	Chemistry
Course Name	Chemistry 02 (Class XI – Semester 02)
Module Name/Title	s block elements - Part 1
Module Id	kech_21001
Pre-requisites	Periodic table and periodic properties of elements
Objectives	After going through this module, the learner will be able to: <ul style="list-style-type: none">• Describe the general characteristics of the s block elements and their compounds.• Explain preparation, properties and uses of industrially important sodium compounds• Diagnol relationship• Explain the biological significance of sodium, potasium
Keywords	Alkali and Alkaline Earth Metals, Calcium Compounds, Atomic Radii, Ionic Radii, Ionization Enthalpy, Oxidation State, Electrode Potential, Chemical Reactivity

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1. Introduction

The s-block elements are those elements in which last electron enters the outermost s-orbital. As the s-subshell has one orbital, which can accommodate only two electrons, two groups (1 & 2) constitute the s-block of the Periodic Table. Group 1 contains one electron in their outermost s orbital with general electronic configuration [noble gas] ns^1 . It consists of the elements: lithium, sodium, potassium, rubidium, caesium and francium. They are collectively known as the alkali metals. These are so called because they form hydroxides on reaction with water, which are strongly alkaline in nature. The elements of Group 2 contains two electrons in their outermost s orbital general electronic configuration [noble gas] ns^2 . These include beryllium, magnesium, calcium, strontium, barium and radium. These elements with the exception of beryllium are commonly known as the alkaline earth metals. These are so called because their oxides and hydroxides are alkaline in nature and these metal oxides remain unchanged by fire. Substances that remain unchanged by fire were called earths by ancient chemists.

The physical and chemical properties of s block elements can be easily understood in terms of the outer electron configuration. Some of these elements like sodium, potassium, calcium and magnesium are extensively involved in life processes. Monovalent sodium and potassium ions and divalent magnesium and calcium ions are found in large proportions in biological fluids. These ions perform important biological functions such as maintenance of ion balance and nerve impulse conduction. Many of their compounds have been used even in the early days of civilization. Some of the compounds of lithium, sodium, potassium, calcium and magnesium have wide industrial applications.

The s block elements are highly reactive, hence are normally found in nature in combined state. Alkali metals occur as their halides, oxides, borate, nitrates and silicates. Whereas alkaline earth metals exist as carbonates, silicates, phosphates and sulphates. Among the alkali metals sodium and potassium are abundant and lithium, rubidium and caesium have

much lower abundances. Francium is highly radioactive; its longest-lived isotope ^{223}Fr has a half-life of only 21 minutes. Pierre Curie discovered francium in 1939. Although sodium and potassium are almost equally abundant in the earth's crust, yet in seawater sodium salts are more abundant than potassium salts. This could be attributed to the preferential uptake of K^+ ions by plants from ground water.

Calcium and magnesium rank fifth and sixth in abundance respectively in the earth's crust. Strontium and barium have much lower abundances. Beryllium is rare and radium is the rarest of all comprising only 10^{-10} per cent of igneous rocks. Madam Curie was awarded Nobel Prize in 1911 for isolating radium and polonium.

In this module we will deal with the diagonal relationship, electronic configuration, occurrence and general characteristics of s block group 1 elements with special emphasis on preparation and properties of some important compounds of sodium.

Lithium and beryllium, the first elements of Group 1 and Group 2 respectively exhibit anomalous behaviour and are different from those of the other members of the respective group. In these anomalous properties they resemble the second element of the following group. Thus, lithium resembles to magnesium and beryllium to aluminium in many of their properties. This diagonal similarity is commonly referred to as diagonal relationship in the periodic table. Although the similarities in these pairs are quite pronounced, yet these are weaker than the similarities found within the group. The diagonal relationship exhibited is due to following reasons:

a) Similarity in atomic or ionic sizes: The atomic and ionic sizes decrease along a period and increase on moving down a group. On moving diagonally these two trends cancel each other as a result atomic sizes of diagonally related elements are similar.

b) Similarity in electropositive character: The electropositive character decreases across a period and increases down a group. Hence diagonally related elements have similar electropositive character.

c) Similarity in polarizing power: Charge to radius ratio of an ion increases along a period as charge on ion increases and size of ion decreases. This results in increase in polarizing power of the ion. As size of ion increase on moving down the group, hence its polarizing power decreases. These two effects cancel out diagonally; hence diagonally related elements have almost similar polarizing power.

Because of these similarities, these elements show similar properties.

2. Properties of GROUP 1 ELEMENTS: ALKALI METALS

The alkali metals show regular trends in their physical and chemical properties with the increasing atomic number. The atomic, physical and chemical properties of alkali metals are discussed below.

2.1 Electronic Configuration: All the alkali metals have [noble gas] ns^1 electronic configuration (Table 1) with one valence electron, ns^1 outside the noble gas core. This loosely held s-electron in the outermost valence shell of these elements makes them the most electropositive metals. They readily lose electron to give monovalent M^+ ions. Hence they are never found in free state in nature. Due to similar outer electron configuration, alkali metals have similar physical and chemical properties.

Table 1: Electronic configuration of alkali metals

Element	Symbol	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	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$ $4p^6 4d^{10} 5s^2 5p^6 6s^1$ or [Xe] $6s^1$
Francium	Fr	[Rn] $7s^1$

2.2 Atomic and Ionic Radii: The alkali metal atoms have the largest sizes in their respective period of the periodic table. With increase in atomic number, the atomic size becomes larger. The monovalent ions (M^+) are smaller than the parent atom as these have one shell less than the parent atom. The atomic and ionic radii of alkali metals increase on moving down the group i.e., they increase in size while going from Li to Cs as number of energy shells increase by one with each succeeding element.

2.3 Ionization Enthalpy: The ionization enthalpies of the alkali metals are considerably low and decrease down the group from Li to Cs. This is because the effect of increasing size outweighs the increasing nuclear charge, and the outermost electron is very well screened from the nuclear charge.

The first ionization enthalpy of the alkali metals is the lowest as compared to other elements of the respective periods, whereas the second ionization enthalpy of the alkali metals is quite high.

2.4 Hydration Enthalpy: The hydration enthalpies of alkali metal ions decrease with increase in ionic sizes. i.e., $\text{Li}^+ > \text{Na}^+ > \text{K}^+ > \text{Rb}^+ > \text{Cs}^+$. Li^+ has maximum degree of hydration and for this reason lithium salts are mostly hydrated, e.g., $\text{LiCl} \cdot 2\text{H}_2\text{O}$

2.5 Physical Properties: All the alkali metals are silvery white, soft and light metals. Because of the large size, these elements have low density, which increases down the group from Li to Cs. However, potassium is lighter than sodium. These metals can be cut with knife. The melting and boiling points of the alkali metals are low indicating weak metallic bonding due to the presence of only one valence electron in them. The melting and boiling points decrease down the group.

Table2: Atomic and Physical Properties of the Alkaline Earth Metals

Property	Lithium Li	Sodium Na	Potassium K	Rubidium Rb	Caesium Cs	Francium Fr
Atomic number	3	11	19	37	55	87
Atomic mass (g mol^{-1})	6.94	22.99	39.10	85.47	132.91	(223)
Electronic configuration	$[\text{He}] 2s^1$	$[\text{Ne}] 3s^1$	$[\text{Ar}] 4s^1$	$[\text{Kr}] 5s^1$	$[\text{Xe}] 6s^1$	$[\text{Rn}] 7s^1$
Ionization enthalpy / kJ mol^{-1}	520	496	419	403	376	~375
Hydration enthalpy/ kJ mol^{-1}	-506	-406	-330	-310	-276	-
Metallic radius / pm	152	186	227	248	265	-
Ionic radius M^+ / pm	76	102	138	152	167	(180)
m.p. / K	454	371	336	312	302	-
b.p. / K	1615	1156	1032	961	944	-
Density / g cm^{-3}	0.53	0.97	0.86	1.53	1.90	-
Standard potentials E^\ominus / V for (M^+ / M)	-3.04	-2.714	-2.925	-2.930	-2.927	-
Occurrence in lithosphere [†]	18*	2.27**	1.84**	78-12*	2-6*	~ 10^{-18} *

*ppm (part per million), ** percentage by weight; † Lithosphere: The Earth's outer layer: its crust and part of the upper mantle

2.6 Flame colouration: 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 of the metals to a higher energy level. When the excited electron comes back to the ground state, these emit the energy in the form of electromagnetic radiations, which fall in the visible region as given below:

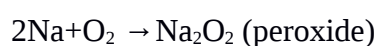
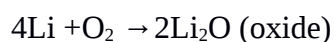
Metal	Li	Na	K	Rb	Cs
Colour	Crimson red	Yellow	Violet	Red violet	Blue
λ/nm	670.8	589.2	766.5	780.0	455.5

Alkali metals can therefore, be detected by the respective flame tests and can be quantitatively determined by flame photometry or atomic absorption spectroscopy.

These elements, especially caesium and potassium, when irradiated with light, emit electrons. Because the light energy absorbed may be sufficient to make an atom lose electron due to its low ionization enthalpy. This property makes caesium and potassium useful as electrodes in photoelectric cells.

2.7 Chemical Properties: The alkali metals are highly reactive due to their large size and low ionization enthalpy. The reactivity of these metals increases down the group.

(i) 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. Lithium forms monoxide, sodium forms peroxide, the other metals form superoxides. The superoxide O_2^- ion is stable only in the presence of large cations such as K, Rb, Cs. (M = K, Rb, Cs)



In all these oxides the oxidation state of the alkali metal is +1. Lithium shows exceptional behaviour in reacting directly with nitrogen of air to form the nitride, Li_3N as well. Because of their high reactivity towards air and water, alkali metals are normally kept in kerosene oil or inert hydrocarbon solvents to prevent their contact with air and moisture.

Problem: What is the oxidation state of K in KO_2 ?

Solution:

The superoxide species is represented as O_2^- ;

Since the compound is neutral, therefore, the oxidation state of potassium is +1.

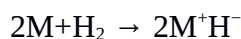
(ii) Reactivity towards water: The alkali metals react with water to form hydroxide and dihydrogen.



It may be noted that although lithium has most negative E value (Table 2), and this is attributed to its small size and very high hydration energy. But its reaction with water is less vigorous than that of sodium which has the least negative E value among the alkali metals. This can be explained on the basis of kinetics of reaction i.e. the rate at which reaction proceeds. Lithium has high melting point than sodium and heat of the reaction is not sufficient to melt lithium and hence the reaction proceeds gently. Other metals of the group react explosively with water.

They also react with proton donors such as alcohol, gaseous ammonia and alkynes.

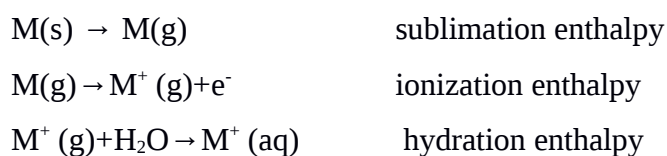
(iii) Reactivity towards dihydrogen: The alkali metals react with dihydrogen at about 673K (lithium at 1073K) to form hydrides. All the alkali metal hydrides (M^+H^-) are colourless, ionic solids with high melting points.



(iv) Reactivity towards halogens: The alkali metals readily react vigorously with halogens to form ionic halides (M^+X^-). However, lithium halides are somewhat covalent. It is because of the high polarizing power of lithium ion (The distortion of electron cloud of the anion by the cation is called polarizing power). The Li^+ ion is very small in size and has high tendency to distort electron cloud around the negative halide ion. Since anion with large size can be easily distorted, among halides, lithium iodide is the most covalent in nature.

(v) Reducing nature: The alkali metals are strong reducing agents, lithium being the most and sodium the least powerful (Table 2).

The standard electrode potential (E^\ominus) which, measures the reducing power represents the overall change:



With the small size of its ion, lithium has the highest hydration enthalpy, which accounts for its high negative E value and its high reducing power.

Problem

The E for Cl^2/Cl is +1.36, for I_2/I is +0.53, for Ag^+/Ag is +0.79, Na^+/Na is -2.71 and for Li^+/Li is -3.04. Arrange the following ionic species in decreasing order of reducing strength:

I^- , Ag , Cl^- , Li , Na

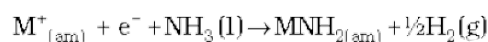
Solution

The order is $\text{Li} > \text{Na} > \text{I}^- > \text{Ag} > \text{Cl}^-$

(vi) Solutions in liquid ammonia: The alkali metals dissolve in liquid ammonia giving deep blue solutions, which are conducting in nature.



The blue colour of the solution is due to the ammoniated electron, which absorbs energy in the visible region (red) of light and thus imparts blue colour to the solution. The solutions are paramagnetic and on standing slowly liberate hydrogen resulting in the formation of amide.



(where 'am' denotes solution in ammonia.)

In concentrated solution, the blue colour changes to bronze colour and becomes diamagnetic due to formation of metallic ion clusters.

2.8 Uses

Lithium: Lithium metal is used to make useful alloys, for example lithium-lead alloy or 'white metal' used for making bearings for motor engines and sheets for cables, lithium-aluminium alloy used to make aircraft parts, and lithium-magnesium alloy to make armour plates and aerospace components.

It is used in thermonuclear reactions to provide energy for propelling rockets and guided missiles.

Lithium is also used to make electrochemical cells.

Sodium: Sodium is used to make a Na/Pb alloy needed to make PbEt_4 and PbMe_4 . These organolead compounds were earlier used as anti-knock additives to petrol, but nowadays vehicles use lead-free petrol.

Liquid sodium metal is used as a coolant in fast breeder nuclear reactors.

Sodium is used in sodium vapour lamps and as reducing agent in the synthesis of many organic reactions.

Potassium: Potassium has a vital role in biological systems. Potassium chloride is used as a fertilizer. Potassium hydroxide is used in the manufacture of soft soap. It is also used as an excellent absorbent of carbon dioxide.

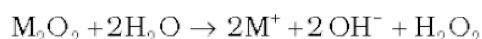
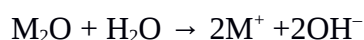
Caesium is used in devising photoelectric cells (solar cells).

3. General Characteristics of the Compounds of the Alkali Metals

All the common compounds of the alkali metals are generally ionic in nature. General characteristics of some of their compounds are discussed here.

3.1 Oxides and Hydroxides

On combustion in excess of air, lithium forms mainly the oxide, Li_2O (plus some peroxide Li_2O_2), sodium forms the peroxide, Na_2O_2 (and some superoxide NaO_2) whilst potassium, rubidium and caesium form the superoxides, MO_2 . Under appropriate conditions pure compounds M_2O , M_2O_2 and MO_2 may be prepared. The increasing stability of the peroxide or superoxide, as the size of the metal ion increases, is due to the stabilisation of large anions by larger cations through lattice energy effects. These oxides, peroxides and superoxides are easily hydrolysed by water to form the hydroxides according to the following reactions:



The peroxides and superoxides act as oxidizing agents as these produce H_2O_2 and O_2 . The oxides and the peroxides are colourless when pure, but the superoxides are yellow or orange in colour. The superoxides are also paramagnetic. Sodium peroxide is widely used as an oxidising agent in inorganic chemistry.

The hydroxides, which are obtained by the reaction of the oxides with water, are all white crystalline solids. The alkali metal hydroxides are the strongest of all bases and dissolve freely in water with evolution of much heat on account of intense hydration. The basic strength of these hydroxides increases on moving down the group.

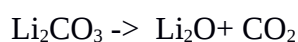
Problem 3: Why is KO_2 paramagnetic?

Solution: The superoxide O_2^- is paramagnetic because of one unpaired electron in π^*2p molecular orbital.

3.2 Halides: The alkali metal halides, MX, (X=F,Cl,Br,I) are all high melting, colourless crystalline solids. They can be prepared by the reaction of the appropriate oxide, hydroxide or carbonate with aqueous hydrohalic acid (HX). All of these halides have high negative enthalpies of formation; the $\Delta_f H$ values for fluorides become less negative as we go down the group, whilst the reverse is true for $\Delta_f H$ for chlorides, bromides and iodides. For a given metal $\Delta_f H$ always becomes less negative from fluoride to iodide. It means alkali metal fluorides are the most stable whereas the iodides are the least stable

The melting and boiling points always follow the trend: fluoride > chloride > bromide > iodide. All these halides are soluble in water. The low solubility of LiF in water is due to its high lattice enthalpy whereas the low solubility of CsI is due to smaller hydration enthalpy of its two ions. Other halides of lithium being covalent in nature, are soluble in ethanol, acetone and ethylacetate; LiCl is soluble in pyridine also and in contrast NaCl being ionic is insoluble in organic solvents.

3.3 Salts of Oxo-Acids: Oxo-acids are those in which the acidic proton is on a hydroxyl group with an oxo group attached to the same atom e.g., carbonic acid, $H_2CO_3(OC(OH)_2)$; sulphuric acid, $H_2SO_4(O_2S(OH)_2)$. The alkali metals form salts with all the oxo-acids. They are generally soluble in water and thermally stable. Their carbonates (M_2CO_3) and in most cases the hydrogencarbonates ($MHCO_3$) also are highly stable to heat. As the electropositive character increases down the group, the stability of the carbonates and hydrogencarbonates increases. Lithium carbonate is not so stable to heat; lithium being very small in size polarises a large CO_3^{2-} ion leading to the formation of more stable Li_2O and CO_2 .



Except lithium all other alkali metals form solid hydrogencarbonate. An interesting fact is that no other metal forms solid hydrogencarbonate.

4. 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 polarizing power (i.e., charge/radius ratio). As a result, there is increased covalent character of lithium compounds, which is responsible for their solubility in organic solvents.
- iii. High ionization enthalpy and non-availability of d orbitals.

Further, lithium shows diagonal relationship to magnesium, which has been discussed subsequently.

4.1 Difference between Lithium and other Alkali Metals

- i. Lithium is much harder.
- ii. Its melting point and boiling point are higher than the other alkali metals.
- iii. 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 which form peroxide and superoxides.
- iv. 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.
- v. Lithium hydrogencarbonate is not obtained in the solid form while all other elements form solid hydrogencarbonates.
- vi. Lithium unlike other alkali metals forms no ethynide on reaction with ethyne.
- vii. Lithium nitrate when heated gives lithium oxide, Li_2O , whereas other alkali metal nitrates decompose to give the corresponding nitrite.



(viii) LiF and Li_2O are comparatively much less soluble in water than the corresponding compounds of other alkali metals. Because of covalent character lithium halides are soluble in organic solvents, while other alkali metal halides are insoluble.

4.2 Similarities between Lithium and Magnesium

The similarity between lithium and magnesium is particularly striking and arises because of their similar sizes: atomic radii, $\text{Li} = 152 \text{ pm}$, $\text{Mg} = 160 \text{ pm}$; ionic radii: $\text{Li}^+ = 76 \text{ pm}$, $\text{Mg}^{2+} = 72 \text{ pm}$. The main points of similarity are:

- i. Both lithium and magnesium are harder and lighter than other elements in the respective groups.
- ii. Lithium and magnesium react slowly with water. Their oxides and hydroxides are much less soluble and their hydroxides decompose on heating.
- iii. The oxides, Li_2O and MgO do not combine with excess oxygen to give any superoxide.
- iv. (Both form a nitride, Li_3N and Mg_3N_2 , by direct combination with nitrogen.
- v. The carbonates of lithium and magnesium decompose easily on heating to form the oxides and CO_2 . Solid hydrogencarbonates are not formed by lithium and magnesium.
- vi. Both LiCl and MgCl_2 are soluble in ethanol.

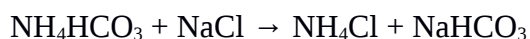
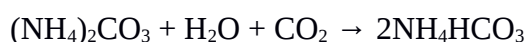
vii. Both LiCl and MgCl₂ are deliquescent and crystallise from aqueous solution as hydrates, LiCl·2H₂O and MgCl₂·8H₂O.

5. Some Important Compounds of Sodium

Industrially important compounds of sodium include sodium carbonate, sodium hydroxide, sodium chloride and sodium hydrogen bicarbonate. The large scale production of these compounds and their uses are described below:

5.1 Sodium Carbonate (Washing Soda), Na₂CO₃·10H₂O

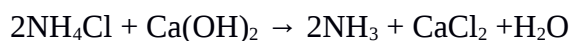
Preparation: Sodium carbonate commonly known as washing soda is generally prepared by Solvay Process. In this process, advantage is taken of the low solubility of sodium hydrogencarbonate whereby it gets precipitated in the reaction of sodium chloride with ammonium hydrogencarbonate. The latter is prepared by passing CO₂ to a concentrated solution of sodium chloride saturated with ammonia, where ammonium carbonate followed by ammonium hydrogencarbonate are formed. The equations for the complete process may be written as:



Sodium hydrogencarbonate crystal separates because of common ion effect due to Na⁺ ions of NaCl. These are heated to give sodium carbonate.



In this process NH₃ is recovered when the solution containing NH₄Cl is treated with Ca(OH)₂ and calcium chloride is obtained as a by-product.



It may be mentioned here that Solvay process cannot be extended to the manufacture of potassium carbonate because potassium hydrogencarbonate is too soluble to be precipitated by the addition of ammonium hydrogencarbonate to a saturated solution of potassium chloride.

Properties:

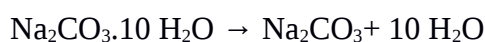
- i. Sodium carbonate is a white crystalline solid, which exists as a decahydrate Na₂CO₃·10H₂O. It is readily soluble in water.

- ii. 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.

373K



above 373K



- iii. Sodium carbonate, being a salt of strong base NaOH and weak acid H_2CO_3 , gets hydrolysed by water to form an alkaline solution.

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.

5.2 Sodium Chloride, NaCl (Common salt)

The most abundant source of sodium chloride is seawater, which contains 2.7 to 2.9% by mass of the salt. In tropical countries like India, common salt is generally obtained by evaporation of seawater. Approximately 50 lakh tons of salt are produced annually in India by solar evaporation. Crude sodium chloride, generally obtained by crystallisation of brine solution, contains sodium sulphate, calcium sulphate, calcium chloride and magnesium chloride as impurities. Calcium chloride, CaCl_2 , and magnesium chloride, MgCl_2 are deliquescent (absorb moisture easily from the atmosphere), hence 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 gas. Crystals of pure sodium chloride separate out. Calcium and magnesium chloride, being more soluble than sodium chloride, remain in solution.

To prevent salt from getting wet in rainy season calcium phosphate or calcium hydrogenphosphate are added. This makes the salt dry and free flowing.

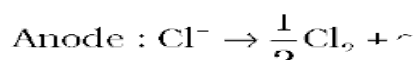
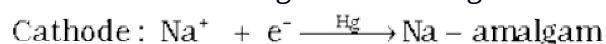
Properties: Sodium chloride melts at 1081K. It has a solubility of 36.0 g in 100 g of water at 273 K. The solubility does not increase appreciably with increase in temperature.

Uses:

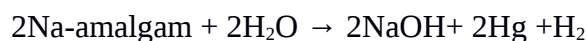
- (i) It is used as a common salt or table salt for domestic purpose.
- (ii) It is used for the preparation of Na_2O_2 , NaOH and Na_2CO_3 .

5.3 Sodium Hydroxide (Caustic Soda), NaOH

Preparation: 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 CO_2 in the atmosphere to form Na_2CO_3 .



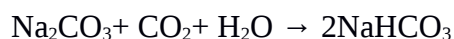
Uses: It is used in

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

5.4 Sodium Hydrogencarbonate (Baking Soda), NaHCO_3

Sodium hydrogencarbonate is known as baking soda because it decomposes on heating to generate bubbles of carbon dioxide (leaving holes in cakes or pastries and making them light and fluffy).

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



Uses:

- (i) Sodium hydrogencarbonate is a mild antiseptic for skin infections.
- (ii) It is used in fire extinguishers.
- (iii) It is used in the preparation of baking powder
- (iv) It is used as an antacid in digestive powders.

7. Biological Importance of Sodium and Potassium

A typical 70 kg man contains about 90 g of Na and 170 g of K compared with only 5 g of iron and 0.06 g of copper.

Sodium ions are found primarily on the outside of cells, being located in blood plasma and in the interstitial fluid, which surrounds the cells. These ions participate in the transmission of nerve signals, in regulating the flow of water across cell membranes and in the transport of sugars and amino acids into cells. Sodium and potassium, although so similar chemically, differ quantitatively in their ability to penetrate cell membranes, in their transport mechanisms and in their efficiency to activate enzymes. Thus, potassium ions are the most abundant cations within cell fluids, where they activate many enzymes, participate in the oxidation of glucose to produce ATP and, with sodium, are responsible for the transmission of nerve signals.

There is a very considerable variation in the concentration of sodium and potassium ions found on the opposite sides of cell membranes. As a typical example, in blood plasma, sodium is present to the extent of 143 mmolL^{-1} , whereas the potassium level is only 5 mmolL^{-1} . Within the red blood cells these concentrations change to $10 \text{ mmolL}^{-1}(\text{Na}^+)$ and $105 \text{ mmolL}^{-1} (\text{K}^+)$. These ionic gradients demonstrate that a discriminatory mechanism, called the sodium-potassium pump, operates across the cell membranes which consumes more than one-third of the ATP used by a resting animal and about 15 kg per 24 h in a resting human.

7. Summary

In this module we have discussed that the s-Block of the periodic table constitutes Group 1 (alkali metals) and Group 2 (alkaline earth metals). They are so called because their oxides and hydroxides are alkaline in nature. The alkali metals are characterised

by one electron in the valence shell of their atoms. These are highly reactive metals forming monovalent (M^+) ions. There is a regular trend in the physical and chemical properties of the alkali metals with increasing atomic numbers. The atomic and ionic sizes increase and the ionization enthalpies decrease systematically down the group. Lithium in Group 1 shows similarities in properties to the second member of the next group. Such similarities are termed as the 'diagonal relationship' in the periodic table. As such this element is anomalous as far as its group characteristics are concerned.

The alkali metals are silvery white, soft and low melting. They are highly reactive. The compounds of alkali metals are predominantly ionic. Their oxides and hydroxides are soluble in water forming strong alkalies. Important compounds of sodium include sodium carbonate, sodium chloride, sodium hydroxide and sodium hydrogencarbonate. Sodium hydroxide is manufactured by the Castner-Kellner process and sodium carbonate by the Solvay process. Monovalent sodium and potassium ions are found in large proportions in biological fluids. These ions perform important biological functions such as maintenance of ion balance.