Chapter
2

s-BLOCK ELEMENTS

Elements of group IA and IIA are called s-block elements. In electronic configuration of these elements, last electrons goes into s-orbital. The number of s-block elements are 14. (IA, IIA, H and He).

Alkali Metals:

"Elements of group IA except hydrogen are called alkali metals." Alkali metals include lithium, sodium, potassium, rubidium, cesium and francium. They are very reactive metals. When these metals are dissolved in water, they form strong alkaline solution. "The base which is water soluble is called alkali."

The name "alkali" is from Arabic language, which means "the ash". The Arabs used this term for these metals because they found ashes of some plants were composed mainly of sodium and potassium oxides.

Alkaline Earth Metals:

"The elements of IIA group are called alkaline earth metals." Groups IIA includes beryllium, magnesium, calcium strontium, barium and radium. Elements of group IIA are called alkaline earth metals because they produce alkalies in water and are widely distributed in earth's crust in the form of carbonates and silicates.

Electronic Configuration of Alkali Metals:

Alkali and alkaline earth metals are electropositive in nature. They lose electrons very easily. Properties of elements are mainly based upon the electronic configuration of elements. The electronic configuration of alkali metals (group IA) are given below:

Table – Electronic configuration of alkali metals

Elements	Symbol	At. Number	Electronic configuration
Lithium	Li	3	$1s^2, 2s^1$
Sodium	Na	11	$1s^2, 2s^2, 2p^6, 3s^1$
Potassium	K	19	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$
Rubidium	Rb	37	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^1$
Cesium	Cs	55	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^1$
Francium	Fr	87	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 6p^6, 7s^1$

Alkali metals have only one electron in 's' orbital of their valence shell. All alkali metals lose their one electron of the valence shell to form monopositive ions M^+, because their ionization energy values are very low. They form ionic compounds and show +1 oxidation state.

$$M \longrightarrow M^+ + e^-$$

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

Table – Physical constants of group IA (alkali metals)

Properties	Li	Na	K	Rb	Cs
Atomic number	3	11	19	37	55
Melting points °C	187	97.5	62.7	39.0	26.0
Boiling points (°C)	1325	889	774	688	690
Atomic radius (pm)	123	158	203	216	235
Ionic radius of +1 ion (pm)	60	95	133	148	169
Ionization energy (kJ/mole)	520	496	419	400	380
Density gm/cm^3 at (20°C)	0.53	0.97	0.86	1.53	1.9
Heat of hydration (kJ/mole)	505	475	384	345	310
Electron affinity (kJ/mole)	–60	–53	–48	–47	–46
Electronegativity	1.0	0.9	0.8	0.8	0.7

In going down the group, the number of shell increases by one at each step and equal to the number of the period to which the element belongs.

Table – Electronic configuration of alkali earth metals

Beryllium	Be	4	$1s^2, 2s^2$
Magnesium	Mg	12	$1s^2, 2s^2, 2p^6, 3s^2$
Calcium	Ca	20	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$
Strontium	Sr	38	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2$
Barium	Ba	56	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2$
Radium	Ra	88	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 60^6, 7s^2$

Alkaline earth metals have two electrons in 's' orbital of their valence shell. All alkaline earth metals lose their two electrons to form di-positive ions M^{++}, because their ionization energy values are low. They form ionic compounds and show +2 oxidation state.

$$M \longrightarrow M^{2+} + 2e^-$$

$$Ca \longrightarrow Ca^{2+} + 2e^-$$

$$Mg \longrightarrow Mg^{2+} + 2e^-$$

Table – Physical constants of group IIA (alkali earth metals)

Properties	Be	Mg	Ca	Sr	Ba
Atomic number	4	12	20	38	56
Melting points °C	1289	651	851	771	849
Boiling points (°C)	2970	1107	1448	1384	1640
Atomic radius (pm)	89	136	174	191	198
Ionic radius of +2 ion (pm)	31	65	99	113	135
2^{nd} ionization energy (kJ/mole)	1800	1450	1150	1060	970
Density gm/cm^3 at (20°C)	1.85	1.74	1.55	2.6	3.5
Heat of hydration (kJ/mole)	2337	1897	1619	1455	1250
Electron affinity (kJ/mole)	240	230	156	168	52

OCCURRENCE OF ALKALI METALS (GROUP IA)

Due to high reactivity, the alkali metals occur in nature in the combined state. None of the alkali metals is found free in nature. Sodium and potassium are abundant alkali metals and each constitute about 2.4 percent of earth's crust. Most of the earth's crust is composed of insoluble alumino silicates of alkali metals.

Table – Metals in group IA

Element	Name of mineral	Chemical formula
Lithium	Spodumene	$LiAl(SiO_3)_2$
Sodium	Rock salt (halite)	$NaCl$
	Chile salt petre	$NaNO_3$
	Natron	$Na_2CO_3.H_2O$
	Trona	$Na_2CO_3.2NaHCO_3.2H_2O$
	Borax	$Na_2B_4O_7.10H_2O$
Potassium	Carnallite	$KCl.MgCl_2.6H_2O$
	Sylvite	KCl
	Alunite (alum stone)	$K_2SO_4.Al_2(SO_4)_3.4Al(OH)_3$

Lithium deposits, usually in the form of complex minerals, are widely scattered over the earth. An important commercial source of lithium is the mineral spondumene $LiAl(SiO_3)_2$.

Small amounts of rubidium and caesium are found in potassium salts deposits. Francium has not been found in nature. It has been prepared artificially in the laboratory and is very unstable, so that a very little is known about this metal.

Occurrence of Alkaline Earth Metals (Group IIA):

Being very reactive, alkaline earth metals also do not occur in free state. The compounds of these metals occur widely in nature.

Magnesium and calcium are very abundant in earth's crust. The outer portion of the earth was originally in the form of silicates and alumino-silicates of alkaline earth metals. Magnesium and calcium, with sodium and potassium are present in the rocks as cations.

Table – Common minerals of the alkaline earth metals (Group IIA)

Element	Mineral	Formula
Beryllium	Beryl	$Be_3Al_2(SiO_3)_6$
	Chrysoberyl	Al_2BeO_4
Magnesium	Magnesite	$MgCO_3$
	Dolomite	$MgCO_3.CaCO_3$
	Carnallite	$KCl.MgCl_2.6H_2O$
	Epsom salt	$MgSO_4.7H_2O$
	Soap stone (talc)	$H_2Mg_3(SiO_3)_4$
	Asbestos	$CaMg_3(SiO_3)_4$
Calcium	Calcite (Lime stone)	$CaCO_3$
	Gypsum	$CaSO_4.2H_2O$
	Flourite	CaF_2
	Phosphorite	$Ca_3(PO_4)_2$
Strontium	Strontionite	$SrCO_3$
Barium	Barite (heavy spar)	$BaSO_4$
	Whetherite	$BaCO_3$

Magnesium halides are found in sea water. Magnesium is an essential constituent of chlorophyll.

Calcium phosphate $Ca_3(PO_4)_2$ and calcium fluoride CaF_2 are also found as minerals. Calcium is an essential constituent of many living organisms. It occurs as skeletal material in bones, teeth, sea-shells and egg shells.

Radium is a rare element. It is of great interest because of its radioactive nature.

PECULAR BEHAVIOUR OF LITHIUM FROM OTHER MEMBER OF GROUP IA

In many properties lithium has different behaviour from other members of group IA. The deviation from the regular trend is due to the following reason.

The **atomic radius of lithium is very small** as compared to the other members of group I. The ionic radius of Li^+ is 60 pm while that of Na^+ is 95 pm. Li^+ ion has greater charge density. Charge density is the ratio of charge to volume. The nuclear charge of Li^+ is screened by only two electrons of k-shell. Due to these reasons, lithium is less electropositive than other members of group IA.

Some important differences of lithium from other metals are given below:

(1) Lithium is much harder and lighter than the other alkali metals.

(2) The lithium salts of anions with high density are generally less soluble in water than those of the other alkali metals. e.g., LiOH, LiF, Li_3PO_4, Li_2CO_3 are not soluble in water.

(3) Lithium forms stable complex compounds, although complex formation generally is not a property of alkali metals. One of the stable complexes formed by lithium is $[Li(NH_3)_4]$.

(4) Lithium reacts very slowly with water, while other alkali metals react violently.

$$2Li + 2H_2O \longrightarrow 2LiOH + H_2$$

(5) Lithium salts of large polarizable anions are less stable than those of other alkali metals. Unlike other alkali metals lithium does not form bicarbonate, tri-iodide or hydrogen sulphide at room temperature.

(6) When burnt in air lithium forms only normal oxide, whereas the others form peroxides or super oxides.

$$4Li + O_2 \longrightarrow 2Li_2O \quad \text{(Lithium oxide)}$$

$$2Na + O_2 \longrightarrow Na_2O_2 \quad \text{(Sodium peroxide)}$$

(7) Lithium hydride (LiH) is more stable than the hydrides of other alkali metals.

(8) Lithium compounds are more covalent, that is why its halides are more soluble in organic solvents and the alkals and aryls of lithium are more stable than those of other alkali metals.

(9) Lithium is the latest reactive metal of all the alkali metals.

(10) When acetylene is passed over strongly heated lithium, it does not produce lithium acetylide, but other alkali metals form the corresponding metallic acetylides.

$$2Na + C_2H_2 \longrightarrow Na^+C^- \equiv C^-Na^+ + H_2$$

(Sodium Acetylide)

(11) Lithium has low electropositive character, thus its carbonate and nitrate are not so stable and therefore decompose giving lithium oxide. Carbonates of other alkali metals do not decompose. Decomposition of lithium nitrate gives different products than the nitrate of other alkali metals.

$$Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2$$

$$4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2$$

$$2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2$$

$$2KNO_3 \longrightarrow 2KNO_2 + O_2$$

(12) Lithium hydroxide when strongly heated, forms lithium oxide but the other alkali metal hydroxides do not show this behaviour.

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

(13) Lithium reacts with nitrogen to form nitride while the other members of the group do not give this reaction.

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

(14) Lithium chloride LiCl has an exothermic heat of solution, whereas chloride of sodium and potassium have endothermic heats of solution.

(*The amount of heat evolved or absorbed when one mole of a substance is dissolved in so much solvent that further dilution results in no detectable heat change is called heat of solution*).

(15) Lithium carbide Li_4C, is the only alkali metal carbide formed readily by direct reaction of carbon with lithium.

$$4Li + C \longrightarrow Li_4C$$

Peculiar behaviour of Beryllium from other member of groups IIA:

Beryllium is the lightest member of the series and differs from the other group IIA elements in many ways. This is due to its small atomic size and comparatively high electronegativity value.

The main points of difference are:

(1) Beryllium metal is almost as hard as iron and hard enough to scratch glass. The other alkaline earth metals are much softer than beryllium but still harder than the alkali metals.

(2) The melting and boiling points of beryllium are higher than other alkali earth metals.

(3) As reducing agents, the group IIA metals are all powerful enough to reduce water. However, with water, beryllium forms insoluble oxide (BeO) coating that protects it from further attack.

(4) Complete oxidation of Beryllium does not occur in the presence of an acid. It forms a protective layer which stops the further attack of an acid.

(5) Beryllium reacts with alkali NaOH to form sodium beryllate while other members do not react with alkali NaOH.

$$Be + 2NaOH \longrightarrow \underset{\text{Sodium beryllate}}{Na_2BeO_2} + H_2$$

CHEMICAL PROPERTIES OF ALKALI METALS

Good Reducing Agents:

Alkali metals are **good reducing agents**. Reducing property of an element depends upon the ionization energy value. Elements with **low ionization energy** value, are good reducing agents. Reducing agent is a substance which can lose electrons and its oxidation increases during a chemical reaction. Alkali metals have low I.E value and are good reducing agents. They are highly electropositive. They react readily with halogens and alkali metal halides.

$$Na + \frac{1}{2}Cl_2 \longrightarrow NaCl$$

Some other properties are given below:

(1) Low ionization energies make the alkali metals the most reactive family of metals.

(2) Very high second ionization energies indicate that oxidation number higher than 1 is not found in alkali metals.

(3) The cations of alkali metals e.g., Na^+, K^+, Rb^+ have low charge and large radii than the radius of any cation from the same period, so the lattice energies of their salts are realtively low. Consequently, most of the simple salts of the alkali metals are water-soluble. Most of the salts are among the strongest bases available.

(4) They react with oxygen and the surface is tarnished due to the oxides formed. Only lithium burns in air to form the normal oxide, Li_2O. In normal oxides, oxidation state of oxygen is -2.

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

White solid (Lithium oxide)

The exposed metals are oxidized almost immediately by oxygen in air and in the presence of moisture, the oxides formed react with CO_2 in the atmosphere to form carbonates.

$$Li_2O + CO_2 \longrightarrow Li_2CO_3$$

Lithium oxide Lithium carbonate

Sodium will undergo a similar reaction, but only if the supply of oxygen is limited.

$$4Na + O_2 \longrightarrow 2Na_2O$$

Limited

In the presence of excess of oxygen, sodium forms the pale yellow peroxide. In peroxide, oxidation state of oxygen is -1.

$$2Na + O_2 \longrightarrow Na_2O_2$$

Excess Sodium peroxide (Pale yellow)

The reaction of sodium metal with oxygen produces a crust of sodium peroxide that hinders further reaction even though sodium metal is still present.

Potassium, rubidium and cesium react with oxygen to form super oxides.

In super oxide, oxidation state of oxygen is $-1/2$.

$$K + O_2 \longrightarrow KO_2$$

Potassium super oxide (Orange yellow)

Caesium explodes spontaneously when it is in contact with air or oxygen.

(5) Very rapid reactions occur when alkali metals react with water. A small piece of sodium (potassium or lithium) floated on water reacts vigorously to liberate energy produced by the reaction may even ignite the hydrogen.

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

$$2K + 2H_2O \longrightarrow 2KOH + H_2$$

The reaction becomes increasingly vigorous from lithium to caesium. Potassium, rubidium and caesium are so highly reactive that they react with ice even at $-100°C$.

(6) Alkali metals form ionic hydrides with hydrogen. In ionic hydrides, oxidation state of hydrogen is -1.

$$2M + H_2 \longrightarrow 2M^+H^-$$

$$2Na + H_2 \longrightarrow 2NaH$$

Rubidium and caesium react violently with hydrogen at room temperature. The other three metals require high temperature in order to form the hydride. Lithium hydride and sodium hydrides are useful sources of hydrogen when treated with water:

$$LiH + H_2O \longrightarrow LiOH + H_2$$

Due to the presence of hydride ion (H^-), the ionic hydrides are used as powerful reducing agents.

(7) Lithium is the only group IA metal that combines with nitrogen and carbon to form nitride and carbide respectively.

$$6Li + N_2 \longrightarrow 2Li_3N \quad \text{(Lithium nitride)}$$

$$4Li + C \longrightarrow Li_4C \quad \text{(Lithium carbide)}$$

(8) Alkali metals react easily with halogens to given halides. Lithium and sodium, react slowly with chlorine at room temperature. Molten sodium burns with brilliant yellow flame in a chlorine atmosphere to form sodium chloride.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

Potassium, rubidium and caesium react vigorously with all the halogens, forming metal halides.

(9) All alkali metals form their sulphides when treated with molten sulphur.

$$2M + S \longrightarrow M_2S$$

$$2Na + S \longrightarrow Na_2S$$

$$2K + S \longrightarrow K_2S$$

CHEMICAL PROPERTIES OF ALKALINE EARTH METALS

(1) Reaction with Air:

The alkaline earth metals burn in oxygen to form normal oxides except barium which forms perioxide. Beryllium is the least reactive metal in the group. It is resistant to complete oxidation and stable in air at ordinary temperature but oxidizes rapidly at about 800°C. Therefore beryllium is not tarnished by atmospheric attack but the metal soon loses the silvery appearance.

$$2Be + O_2 \xrightarrow{800°C} 2BeO$$

When exposed to air, magnesium quickly becomes coated with the layer of MgO. This layer protects the surface from further corrosion at ordinary temperature.

$$2Mg + O_2 \longrightarrow 2MgO$$

When magnesium is burnt in air a small amount of nitride is also formed along with magnesium oxide.

When barium is heated in air or oxygen at 500–600°C, its peroxide is formed.

$$Ba + O_2 \xrightarrow{500-600°C} BaO_2$$

Barium peroxide

(2) Reaction with Hydrogen:

Hydrides are produced by treating the molten alkaline earth metals with hydrogen, usually under high pressure.

Magnesium reacts with hydrogen at high pressure and in the presence of a catalyst (MgI_2) forming magnesium hydride.

$$Mg + H_2 \xrightarrow[(MgI_2)]{\text{Pressure}} MgH_2$$

Similarly $Ca + H_2 \xrightarrow[(MgI_2)]{\text{Pressure}} CaH_2$

(3) Reaction with Nitrogen:

All group II-A elements react with nitrogen on heating giving nitrides. For example:

$$3Mg + N_2 \longrightarrow Mg_3N_2 \quad \text{(Magnesium nitride)}$$

$$3Ca + N_2 \longrightarrow Ca_3N_2 \quad \text{(Calcium nitride)}$$

The nitrides hydrolyze vigorously when treated with water, giving ammonia and the respective hydroxides.

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

$$Ca_3N_2 + 6H_2O \longrightarrow 3Ca(OH)_2 + 2NH_3$$

(4) Reaction with Sulphur:

Elements of group II-A react with sulphur, to form metal sulphides. e.g.,

$$Mg + S \longrightarrow MgS \quad \text{(Magnesium sulphide)}$$

(5) Reaction with Halogens:

All group IIA elements react directly with halogens giving halides of the type MX_2. e.g.,

$$Ca + Cl_2 \longrightarrow CaCl_2$$

(6) Reaction with Water:

Magnesium is more reactive than beryllium, even though it is not attacked by cold water. Magnesium reacts slowly with boiling water and quite rapidly with steam to liberate hydrogen.

$$Mg_{(s)} + H_2O_{(g)} \xrightarrow{100^\circ C} MgO_{(s)} + H_{2(g)}$$

Beryllium does not react with water even at red-hot temperature but remaining alkaline earth metals produce hydroxides with water.

$$M + 2H_2O \xrightarrow{100^\circ C} M(OH)_2 + H_2$$

GENERAL TREND IN PROPERTIES OF COMPOUNDS OF ALKALI AND ALKALINE EARTH METALS

(1) Oxides:

(a) Solubility of Group IA Oxides:

Normal oxide of group IA and IIA reacts with water to form hydroxides. Peroxide and superoxides on dissolving in water forms hydroxide and oxygen gas.

$$Li_2O_{(s)} + H_2O_{(l)} \longrightarrow 2LiOH_{(aq)}$$

$$2Na_2O_{2(s)} + 2H_2O_{(l)} \longrightarrow 4NaOH_{(aq)} + O_{2(g)}$$

$$2KO_2 + 2H_2O \longrightarrow 2KOH + H_2O_2 + O_2$$

The reaction of an alkali metal oxide with water is an acid-base reaction and not an oxidation reduction reaction since no element undergoes a change in its oxidation number. The reaction simply involves the decomposition of water molecule by an oxide ion.

$$O^{2-} + H_2O \longrightarrow 2OH^-$$

The basic character of alkali metal oxides increases down the group.

Potassium super oxide (KO_2) has a very interesting use in breathing equipments for **mountaineers** and in space craft. It has the ability to absorb carbon dioxide while giving out oxygen at the same time.

$$4KO_{2(s)} + 2CO_{2(g)} \longrightarrow 2K_2CO_{3(s)} + 3O_{2(g)}$$

(b) Solubility of Group IIA:

The basic characters of alkali metal oxides and alkaline earth metal oxides increase downward in a group. Due to smaller ionic radii, the elements of group IIA are less basic as compared to elements of group IA.

Group IA	Group IIA
Li_2O (Basic)	BeO (Amphoteric)
Na_2O (Basic)	MgO (Basic)
K_2O (Basic)	CaO (Basic)
Rb_2O (Basic)	SrO (Basic)
CsO_2 (Basic)	BaO (Basic)

Along the group from top to bottom, basic characters of oxides increase. Along the period from left to right, basic characters decrease.

All basic oxides react with acids to form salt and water.

$$Na_2O + 2HCl \longrightarrow 2NaCl + H_2O$$

$$CaO + 2HCl \longrightarrow CaCl_2 + H_2O$$

BeO is amphoteric oxides, it reacts with acids as well as bases to form salt and water.

$$BeO + H_2SO_4 \longrightarrow BeSO_4 + H_2O$$

Base Acid Salt Water

$$BeO + 2NaOH \longrightarrow Na_2BeO_2 + H_2O$$

Acid Base Sodium beryllate (Salt)

(2) Hydroxides:

(a) Solubility of Hydroxides:

Alkali metal hydroxides are crystalline solids and highly soluble in water except LiOH. LiOH is slightly soluble in water. Hydroxides of group IA are strong bases except LiOH and hygroscopic in nature.

(The substance which **absorbs moisture** from the atmosphere is called **hygroscopic** substance). The solubility of alkaline earth metal increases downward in a group $Be(OH)_2$ is insoluble in water. $Mg(OH)_2$ is sparingly or slightly soluble while $Ba(OH)_2$ is highly soluble.

Solubility of alkaline earth metals **increases downward** due to increase in the size of cations. Due to greater ionic radii of group IIA cations, the lattice energy of hydroxide decreases. *"Lattice energy is the amount of energy required to break ionic solid into its gaseous ions."*

(b) Thermal Stability:

Alkali metal hydroxides are stable to heat except LiOH, while alkaline earth metal hydroxides like $Mg(OH)_2$ and $Ca(OH)_2$ decompose on heating.

$$2LiOH_{(s)} \longrightarrow Li_2O_{(s)} + H_2O_{(l)}$$

$$Mg(OH)_{2(s)} \longrightarrow MgO_{(s)} + H_2O_{(l)}$$

$Be(OH)_2$ is amphoteric in character.

$$Be(OH)_2 + H_2SO_4 \longrightarrow BeSO_4 + 2H_2O$$

$$Be(OH)_2 + 2NaOH \longrightarrow Na_2BeO_2 + 2H_2O$$

Note: NaOH is called caustic soda and KOH is called caustic potash. A saturated solution of $Ca(OH)_2$ is called **milk of lime** and is used to identify CO_2. A saturated solution of $Mg(OH)_2$ is called **milk of magnesia**. Milk of magnesia is used to reduce the acidity of stomach.

(3) (i) Carbonates of Alkali Metals:

(a) Effect of Heat:

The carbonates of alkali metals are all **soluble in water** and are stable towards heat except Li_2CO_3 which is not only insoluble but also decompose on heating to lithium oxide.

$$Li_2CO_3 \longrightarrow Li_2O + CO_2$$

The decomposition is made easy because the electrostatic attraction in converting from carbonate to oxide is considerable. In case of large cation like K^+ ion K_2CO_3 (pearl ash), the gain in electrostatic attraction is relatively much less and the decomposition is difficult.

(b) Sodium carbonate is very important industrial chemical. At temperature below 35.2°C, Na_2CO_3 crystalizes out from water as $Na_2CO_3.10H_2O$ which is called washing soda. Above this temperature it crystallizes out from water as

$Na_2CO_3.H_2O$. On standing in air, $Na_2CO_3.10H_2O$ slowly loses water and converted to a white powder $Na_2CO_3.H_2O$. *"The substance which loses water of crystallization in atmosphere at room temperature is called efflorescent."*

(c) The aqueous solution of Na_2CO_3 is base due to hydrolysis of the salt. Na_2CO_3 is salt of strong base and weak acid.

$$Na_2CO_3 + 2H_2O \longrightarrow 2NaOH + H_2CO_3$$

(ii) Carbonates of Alkaline Earth Metals:

Unlike the alkali metal carbonates, the alkaline earth metal carbonates are slightly soluble in water, with the **solubility decreasing down the group**. They also decompose on heating and the ease of decomposition decreases down the group.

$$MgCO_3 \xrightarrow{450^\circ} MgO + CO_2$$

$$CaCO_3 \xrightarrow{900^\circ} CaO + CO_2$$

$$SrCO_3 \longrightarrow SrO + CO_2$$

$$BaCO_3 \longrightarrow BaO + CO_2$$

The ease of decomposition can be related to the size of the metal ion, the smaller the ion, the more is the lattice energy of the resulting oxide and hence higher the stability of the product.

(4) Nitrates of Group IA and IIA:

Nitrates of both alkali and alkaline earth metals are soluble in water.

Nitrate of Li, and alkaline earth metals decompose on heating to give O_2, NO_2 and the metallic oxide whereas nitrates of alkali metals decomposes to give alkali metal nitrite and oxygen except lithium.

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

$$2NaNO_3 \longrightarrow 2NaNO_2 + O_2$$

$$2KNO_3 \longrightarrow 2KNO_2 + O_2$$

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

$$2Ca(NO_3)_2 \longrightarrow 2CaO + 4NO_2 + O_2$$

(5) Sulphates of Group IA and IIA:

(a) Al the alkali metals sulphates are soluble in water.

(b) The **solubility of sulphates** of alkaline earth metals gradually **decreases** down the group. $BeSO_4$ and $MgSO_4$ are fairly soluble in water. $CaSO_4$ is slightly soluble and $SrSO_4$ and $BaSO_4$ are almost insoluble.

Calcium sulphate occurs in nature as gypsum $CaSO_4.2H_2O$. When it is heated above 100°C, it loses three quarters of its water of crystallization, giving a white powder called Plaster of Paris.

$$\underset{\text{Gypsum}}{2CaSO_4.2H_2O} \longrightarrow \underset{\text{Plaster of Paris}}{(CaSO_4)_2.H_2O} + 3H_2O$$

MANUFACTURE OF SODIUM METAL BY DOWN'S CELL

Most of sodium metal is produced by the electrolysis of fused sodium chloride. Since the melting point of sodium chloride is 801°C, some amount of **calcium chloride** is added to lower its melting point and to permit the furnace to operate at about 600°C.

Construction of Cell (J.C. Down designed this cell):

In the electrolytic cell, the large block of **graphite** at the centre acts as anode above which there is a dome for the collection of chlorine. The **cathode** is a circular bar of **copper** or iron which surrounds the anode but it separated from it by an iron screen, which terminates in gauze. The arrangement permits the electric current to pass freely but prevents sodium and chlorine from mixing after they have been set free at the electrodes.

Sodium metal rises in a special compartment from which it is taken out of intervals. The cell produces dry chlorine and 99.9 g percent pure sodium. The process is carried out at 600°C and it has the following reactions.

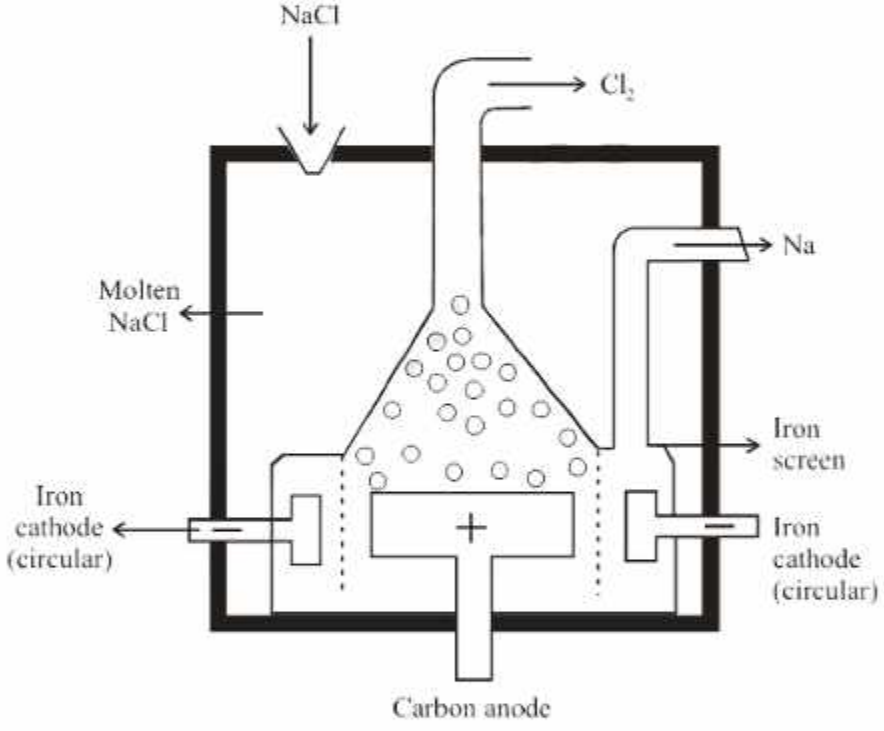


Fig. Down's cell

Reactions:

When sodium chloride is heated at high temperature, it ionizes as follows:

$$NaCl_{(s)} \rightleftharpoons Na^{+}_{(l)} + Cl^{-}_{(l)}$$

Na^{+} ions move towards cathode and Cl^{-} ions move towards anode.

At cathode: $Na^+ + e^- \longrightarrow Na$ (Reduction)

At anode: $Cl^- \longrightarrow \frac{1}{2}Cl_2 + e^-$ (Oxidation)

Not cell reaction: $NaCl \longrightarrow Na + \frac{1}{2}Cl_2$

Advantages of Process:

(i) The metallic fog is not produced.

(ii) Liquid sodium can easily be collected at 600°C.

(iii) Material of the cell is not attacked by the products formed during the electrolysis.

MANUFACTURE OF SODIUM HYDROXIDE

Sodium hydroxide NaOH, is also called caustic soda or lye.

"Caustic" means "to corrode the skin."

On large scale, sodium hydroxide is prepared by the Nelson cell method (diaphragm cell) or Castner-Kellner's cell (Hg-cell).

Construction of Nelson Cell:

The cell is made of steel tank provided with a constant level device to keep the inner part of the tank filled the specified level with brine. Saturated solution of NaCl is called brine.

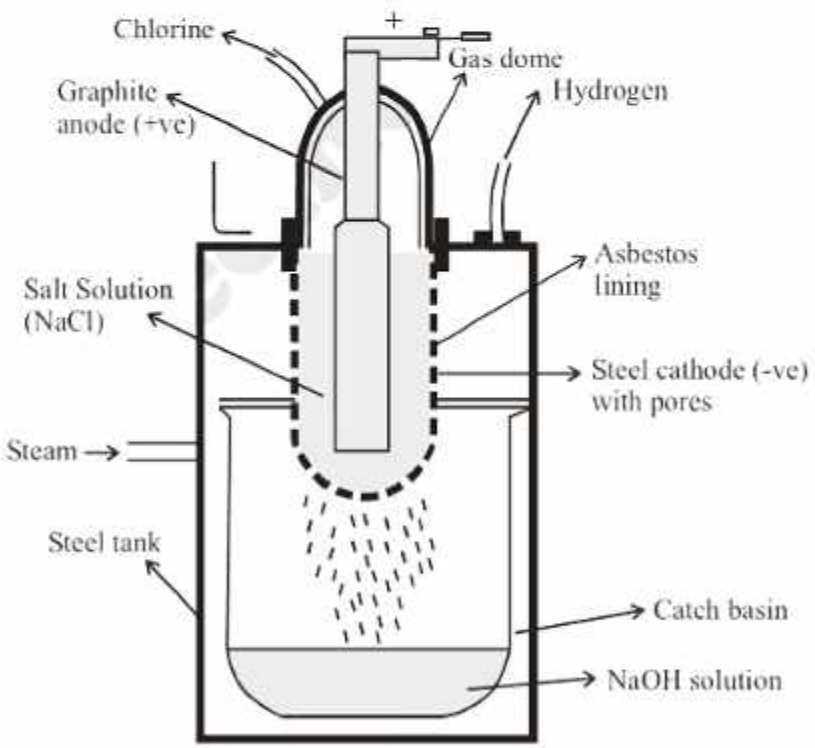
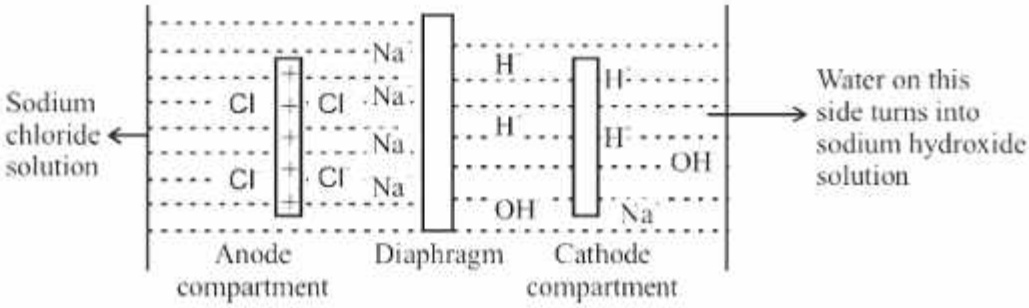


Fig. Nelson cell for production of NaOH

An oblong perforated steel vessel linked inside with asbestos diaphragm serves as a cathode. A graphite anode is held within the U shape diaphragm and it projects into the salt solution. The steam is blown during the process which keeps the electrolyte warm and helps to keep the perforations clear.

The chlorine released at the anode, rises into the dome at the top while hydrogen released at the cathode, escapes through a pipe. The sodium hydroxide solution slowly percolates into a catch basin.



The figure shows a simplified version of the cell in order to understand the purpose of diaphragm. When the electrolysis takes place, chlorine is given off at the anode according to the following reaction.

Reaction at Anode: $2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e^-$

Reaction at Cathode: At the cathode, hydrogen is discharged by the reduction of water. Reduction potential of H_2O is greater than Na^+ ions.

$$2H_2O + 2e^- \longrightarrow 2OH^-_{(aq)} + H_{2(g)}$$

The overall result of the above reaction is that the brine loses its chloride ions and the solution turns increasingly alkaline in cathode compartment.

$$2Cl^- + 2H_2O \longrightarrow 2OH^- + H_2 + Cl_2$$

$$2NaCl + 2H_2O \longrightarrow 2NaOH + H_2 + Cl_2$$

Na^+ ions are not oxidized or reduced in this cell. The ions which do not undergo oxidation or reduction are called spectator ions.

We can face two major problems during the working of the cell:

(1) Chlorine produced can react with hydroxide ions in cold giving hypo-chlorite ions.

$$Cl_{2(g)} + 2OH^-_{(aq)} \longrightarrow OCl^-_{(aq)} + Cl^-_{(aq)} + H_2O$$

(2) Hydroxide ions may be attracted towards anode, where they can be discharged releasing oxygen gas. This oxygen gas may contaminate the chlorine and renders it impure.

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

Solution of the two Problems:

(1) The first problem is solved by using asbestos diaphragm. This keeps the two solutions separate while allowing sodium ions to move towards the cathode. This movement of ions keep the current flowing through the external circuit.

(2) The second problem is solved keeping the level of brine in anode compartment slightly higher, this keeps the direction of flow of liquid toward the cathode and thus preventing the possibility of hydroxide ions to reach the anode.

The solution that flows out of the cathode compartment contains 11% NaOH and 16% NaCl. Evaporation of this solution crystallizes the less soluble NaCl which is filtered off, the liquid contains about 50% NaOH and only 1% NaCl as an impurity. For commercial purposes this small impurity is not important.

GYPSUM

Its chemical name is calcium sulphate di-hydrate and chemical formula $CaSO_4.2H_2O$. It is a mineral that occurs in large deposits throughout the world.

Uses of Gypsum:

(1) Role in Agriculture:

Gypsum is applied to the soil as a **source of calcium and sulphur**. The calcium supplied by gypsum in fertilizers is of importance in crop production in area where soils are subject to extensive leaching.

Sulphur has been recognized as an essential constituent of plants. For centuries, sulphur compounds had been applied to soils because of their observed beneficial effect on plant growth. Aside from serving as a **constituent of protein** and various other compounds in plants, sulphur has an influence on **chlorophyll development** in plant leaves. Although it is not a constituent of chlorophyll but plants deficient in sulphur exhibits a pale green colour.

The **root system** of several plants has been observed to be greatly enlarged by the application of sulphur. It has been reported that good crops are produced by the application of sulphur containing materials such as gypsum.

(2) Role in Industry:

Portland cement is made by strongly heating a finely powdered mixture of clay and limestone. The final product, known as clinker, is cooled and then ground into a very fine powder. During the grinding about 2–5% of gypsum is added which prevents the cement from hardening too rapidly. The addition of gypsum increases the setting time of cement and increases interlocking property.

PLASTER OF PARIS

Plaster of Paris is also called calcium sulphate hemihydrates or half hydrate. Its chemical formula is $CaSO_4 . \frac{1}{2}H_2O$. Plaster of Paris can be prepared by careful heating of gypsum at 100°C.

$$2CaSO_4.2H_2O \longrightarrow (CaSO_4)_2.H_2O + 3H_2O$$

Plaster of Paris

When gypsum is strongly heated at 400°C, it changes to a anhydrous calcium sulphate. Anhydrous salt absorb water very slowly and called "dead burnt gypsum".

$$CaSO_4.2H_2O \xrightarrow{\text{Heat}} CaSO_4 + 2H_2O$$ (Dead burnt gypsum)

Plaster of Paris when mixed with half of its weight of water, it forms a plastic type viscous mass and then sets to a hard porous mass. This process is completed within 10 to 15 minutes. During the process, expansion about 1% in volume also occurs, which fills the moulds completely and thus a sharp casting is obtained.

Uses of Plaster of Paris:

(i) It is used for making plaster walls and roots.

(ii) Plaster of Paris bandages are used for holding in place fractured bones after they have been set.

(iii) It is used to make casts for statuary and coins etc.

Special plasters contain plaster of Paris and other ingredients which vary with the demands of the use to which they are to be put.

Two varieties of plasters are made.

(1) Cement Plaster:

It is plaster of Paris to which usually glue or other oils have been added to prolong the time of setting.

(2) Hard Finish Plaster:

These are made by the calcinations of the anhydrous sulphate with alum or borax. These plasters are set very slowly but give a hard finish.

When mixed with wood pulp and allowed to set in the form of boards, it forms a material, much used in the construction of buildings as wall boards and partitions.

LIME

Lime, (CaO) is a soft, white compound which is obtained by the thermal decomposition of $CaCO_3$.

(1) Role of Lime in Agriculture:

Large quantities of calcium oxide are used in agriculture for **neutralizing acidic soil and** for preparing sprays. It increase the pH and increases the amount of readily soluble phosphorus in soil.

Calcium oxide is used in large amounts for making **lime-sulphur sprays** which have a strong fungicidal action. The hydroxide of calcium is obtained when the oxide of the calcium is allowed to react with water. The process is called **slaking of lime** and it is an exothermic reaction.

$$CaO + H_2O \longrightarrow Ca(OH)_2 \qquad \text{(Slaked lime)}$$

It has been found that application of lime to acidic soils increases the amount of readily soluble phosphorus.

(2) Function of Calcium in Plant Growth:

The presence of calcium is essential for the normal development of plants. The quantity of calcium required by different plants varies considerably. An adequate supply of calcium appears to stimulate the **development of root hairs** and, in fact, the entire root system.

Calcium is also necessary for normal leave development and tends to accumulate in leaves as well as in bark. An adequate supply of calcium is also essential for the optimum activity of micro-organisms that produce nitrates.

The effect of calcium on the supply of available phosphorus in the soil is a special significance. Soils containing sufficient calcium are slightly alkaline in nature.

When a deficiency of calcium exists various substances such as aluminium and manganese may accumulate in plants in harmful concentrations.

Uses of Lime:

(1) Large quantities of lime are used in the extraction and refining of metals.

(2) Lime is also used in paper industries.

(3) The ability of lime to react with sand at high temperature forming calcium silicate $CaSiO_3$ serves as an important basis for glass manufacture.

$$CaO + SiO_2 \longrightarrow CaSiO_3$$

(4) Lime is used ceramic industry for producing different types of sanitary materials.

(5) Ordinary mortar, also called lime mortar, is prepared by mixing freshly slaked lime with sand and water to form a thick paste. Mortar is made by mixing slaked lime (one volume) with sand (three or four volumes) and water to make a thick paste. This material when placed between the stones and bricks hardens or sets, thus binding the block firmly together. The equations for the chemical reactions that take place when mortar hardens are:

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

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

$$Ca(OH)_2 + SiO_2 \longrightarrow CaSiO_3 + H_2O$$

(6) Lime is also used in refining of sugar and other food products.

(7) Lime is sued in the manufacturing of bleaching powder, which used for the bleaching of the fabric and paper pulp.

$$Ca(OH)_2 + Cl_2 \xrightarrow{40^\circ C} Ca(OCl)Cl + H_2O$$

(8) A suspension of the calcium hydroxide $Ca(OH)_2$ is called milk of lime and is used as a white-wash.

(9) When lime is heated with coke at about 2800°C in an electric furnace, calcium carbide is produced.

$$CaO + 3C \longrightarrow CaC_2 + CO$$

Calcium carbide on hydrolysis produces acetylene.

$$CaC_2 + 2H_2O \longrightarrow Ca(OH)_2 + C_2H_2 \quad \text{(Acetylene)}$$

(10) Lime is often employed as a dehydrating agent, for example, in the preparation of absolute alcohol and the drying of ammonia gas. A mixture of sodium hydroxide and calcium hydroxide (soda lime) is often employed to remove both water and carbon dioxide form certain gases.

EXERCISE

Q.1 Fill in the blanks:

(i) Alkali metals are ———— reactive than alkaline-earth metals.

(ii) Alkali metals decompose water vigorously producing ———— and hydrogen.

(iii) When heated in a current of dry hydrogen, alkaline earth metals form white crystalline ———— of the type MH_2.

(iv) The beryllium hydroxide, like the hydroxide of aluminium is amphoteric, while the hydroxides of the other members of the group IIA are ————.

(v) The elements of the group IA are termed as alkali metals, because their ———— are alkaline.

(vi) Spodumene is an ore of ———— metal.

(vii) Alkali metal nitrates on heating give the corresponding ———— and oxygen (except Li).

(viii) Na_2CO_3, H_2O is the chemical formula of a mineral of sodium which is known as ————.

(ix) Metallic bicarbonates are decomposed on heating into their carbonates, alongwith ———— and ————.

(x) Metal nitrates other than the alkali metals on heating decompose into the corresponding metal ———— alongwith the evolution of nitrogen peroxide and oxygen.

ANSWERS

(i)	more	**(ii)**	hydroxide	**(iii)**	hydrides
(iv)	basic	**(v)**	hydroxides, oxides	**(vi)**	lithium
(vii)	nitrite	**(viii)**	natron	**(ix)**	CO_2, H_2O
(x)	oxides				

Q.2 Indicate True or False:

(i) Group IA elements are called alkali metals because their chlorides are alkaline in nature.

(ii) Alkali metals are very good conductor of electricity.

(iii) The hydroxides of alkali metals and alkaline-earth metals are soluble in water.

(iv) Plaster of Paris is a hemihydrate.

(v) Alkali metals have low melting and boiling points as compared to those of alkaline-earth metals.

(vi) Lithium carbonate is decomposed to its oxide, but the carbonates of the other alkali metals are stable towards heat.

(vii) All alkali metal sulphates are insoluble in water.

(viii) Lithium combines with nitrogen to form lithium nitride but other alkali metals do not react with nitrogen.

(ix) Trona is a mineral of lithium.

(x) Alkaline-earth metals are stronger reducing agents than alkali metals.

ANSWERS

(i)	False	**(ii)**	True	**(iii)**	True	**(iv)**	True	**(v)**	True
(vi)	True	**(vii)**	False	**(viii)**	True	**(ix)**	False	**(x)**	False

Q.3 Multiple choice questions. Encircle the correct answer:

(i) Which one of the following does not belong to alkaline-earth metals:

(a) Be **(b)** Ra

(c) Ba **(d)** Rn

(ii) The oxides of beryllium are:

(a) Acidic **(b)** Basic

(c) Amphoteric **(d)** None of these

(iii) Which ion will have the maximum value of heat of hydration:

(a) Na^{+} **(b)** Ca^{2+}

(c) Ba^{2+} **(d)** Mg^{2+}

(iv) Which one of the following is not an alkali metals:

(a) Francium **(b)** Caesium

(c) Rubidium **(d)** Radium

(v) Which of the following sulphates is not soluble in water:

(a) Sodium sulphate **(b)** Potassium sulphate

(c) Zinc sulphate **(d)** Barium sulphate

(vi) The element cesium bears resemblance with:

(a) Ca **(b)** Cr

(c) Both of these metals **(d)** None of these metals

(vii) Chile saltpetre has the chemical formula:

(a) $NaNO_3$ **(b)** KNO_2

(c) $Na_2B_4O_7$ **(d)** $Na_2CO_3.H_2O$

(viii) The mineral $CaSO_4.2H_2O$ has the general name:

(a) Gypsum **(b)** Dolomite

(c) Calcite **(d)** Epsom salt

(ix) Down's cell is used to prepare:

(a) Sodium carbonate **(b)** Sodium bicarbonate

(c) Sodium metal **(d)** Sodium hydroxide

(x) Which element is deposited at the cathode during the electrolysis of brine in diaphragm cell:

(a) H_2 **(b)** Na

(c) Cl_2 **(d)** O_2

ANSWERS

(i)	(d)	**(ii)**	(c)	**(iii)**	(d)	**(iv)**	(d)	**(v)**	(d)
(vi)	(d)	**(vii)**	(a)	**(viii)**	(a)	**(ix)**	(c)	**(x)**	(a)

Q.4 (a) Give the names, electronic configurations and occurrence of s-block elements.

(b) Discuss the peculiar behaviour of lithium with respect to the other members of alkali metals.

Ans.

(a) Descriptive question. For details see text book.

(b) Descriptive question. For details see text book.

Q.5 Discuss the trends in chemical properties of compounds like oxides, hydroxides carbonates, nitrates and sulphates of IA and IIA group elements.

Ans. Detailed question. For answer see text book.

Q.6 Compare the chemical behaviour of lithium with magnesium.

Ans. As Li and Mg have diagonal relationship so they resemble in chemical properties which are given below:

(i) On heating in air, both Li and Mg form normal oxides.

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

$$2Mg + O_2 \longrightarrow 2MgO$$

(ii) Carbonates and phosphates of both Li and Mg are insoluble in water.

(iii) Carbonates of both Li and Mg are unstable and decompose to form oxides.

$$Li_2CO_3 \longrightarrow Li_2O + CO_2$$

$$MgCO_3 \longrightarrow MgO + CO_2$$

(iv) The nitrates of both Li and Mg decompose to form NO_2 and O_2 leaving a residue of oxides:

$$4LiNO_3 \longrightarrow 2LiO_2 + 4NO_2 + O_2$$

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

(v) Both Li and Mg react with N_2 to form nitrides:

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

$$3Mg + N_2 \longrightarrow Mg_3N_2$$

(vi) Both Li and Mg give only monoxides Li_2O and MgO.

Q.7 (a) Mention the properties of beryllium in which it does not resemble with its own family.

(b) Why the aqueous solution of Na_2CO_3 is alkaline in nature?

Ans.

(a) Detailed question. Consult text book.

(b) When Na_2CO_3 is dissolved in water

$$Na_2CO_3 + 2H_2O \longrightarrow 2NaOH + H_2CO_3$$

A strong alkali NaOH and a weak acid H_2CO_3 are formed. Due to formation of NaOH, solution of Na_2CO_3 is alkaline in nature.

Q.8 (a) Describe with diagram the manufacture of sodium by Down's cell.

(b) Point out the three advantages of this process.

Ans. Detailed questions. See text book.

Q.9 (a) Compare the physical and chemical properties of alkali metals with those of alkaline-earth metals.

(b) What happens when:

(i) Lithium carbonate is heated.

(ii) Lithium hydroxide is heated to red-hot.

(iii) Beryllium is treated with sodium hydroxide.

(iv) Lithium hydride is treated with water.

Ans.

(a) Comparison of properties of alkali and alkaline earth metals:

(i) Similarities of group IA with IIA:

(1) Both group IA and IIA elements are s-block elements and last electron goes into the s-orbital.

(2) Both alkali and alkaline earth metals are electropositive and good reducing agents.

(3) Both alkali and alkaline earth metals are very reactive and do not exist in free state.

(4) Ionization energy of both alkali and alkaline earth metal decreases downward in group.

(5) Both alkali and alkaline earth metal elements react with water to form hydrogen gas. e.g.,

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

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

(6) Hydroxide of both families are basic in nature.

(7) Both alkali and alkaline earth metal elements can be prepared by the electrolysis of their fused halides. e.g.,

$$2NaCl \longrightarrow 2Na + Cl_2$$

$$CaCl_2 \longrightarrow Ca + Cl_2$$

(8) Melting points and boiling points of both groups IA and IIA decrease downward in a group.

(9) Elements of both alkali and alkaline earth metals reach with oxygen to form oxides.

(10) The elements of both families on heating on Bunsen burner flame impart characteristic flame. e.g.,

Na – yellow, K – violet, Cs – blue, Ca – brick red, Sr – crimson red, Ba – grassy green

(ii) Dissimilarities of group IA and IIA:

(1) Alkali metals have one electron in their outermost shell while alkaline earth metals have two electrons in their outermost shell.

(2) Alkali metals have +1 oxidation state while alkaline earth metals have +2 oxidation state in their compounds.

(3) Hydroxides of alkaline earth metals are less basic than alkali metals.

(4) Atomic radii and ionic radii of alkaline earth metals is less than alkali metals.

(5) Carbonates of alkaline earth metals are water insoluble while carbonates of alkali metals are soluble except Li_2CO_3.

(6) Alkaline earth metals are harder and heavier than alkali metals.

(7) Heat of hydration of alkaline earth metals is greater than alkali metals.

(8) Alkaline earth metals on heating with oxygen form normal oxides while alkali metals form normal oxide, peroxide and super oxides.

$$2Be + O_2 \longrightarrow 2BeO \quad \text{(Normal oxide)}$$

$$2Mg + O_2 \longrightarrow 2MgO \quad \text{(Normal oxide)}$$

$$4Li + O_2 \longrightarrow 2Li_2O \quad \text{(Normal oxide)}$$

$$2Na + O_2 \longrightarrow Na_2O_2 \quad \text{(Peroxide)}$$

$$K + O_2 \longrightarrow KO_2 \quad \text{(Super oxide)}$$

(9) Carbonates of alkaline earth metals on decomposition evolve CO_2 and normal oxides. Carbonates of alkali metals are thermally stable except lithium carbonate which decomposes on heating.

$$Li_2CO_3 \longrightarrow Li_2O + CO_2$$

$$MgCO_3 \longrightarrow MgO + CO_2$$

$$CaCO_3 \longrightarrow CaO + CO_2$$

(10) Due to smaller size ionization energy of IIA group is greater than IA group.

(b) Li and Mg have diagonal relationship so they resemble in chemical.

(i) Lithium carbonate is heated:

$$Li_2CO_3 \longrightarrow \underset{\text{Lithium oxide}}{Li_2O} + CO_2$$

(ii) Lithium hydroxide is heated to red-hot:

$$2LiOH \longrightarrow \underset{\text{Lithium oxide}}{Li_2O} + H_2O$$

(iii) Beryllium is treated with sodium hydroxide:

$$Be + 2NaOH \longrightarrow Na_2BeO_2 + H_2$$

Sodium beryllate

(iv) Lithium hydride is treated with water:

$$LiH + H_2O \longrightarrow LiOH + H_2$$

Lithium hydroxide

Q.10 Give formulas of the following minerals:

Ans.

(a)	Dolomite	$MgCO_3 . CaCO_3$
(b)	Asbestos	$CaMg_3(SiO_3)_4$
(c)	Halite	$NaCl$
(d)	Natron	$Na_2CO_3 . H_2O$
(e)	Beryl	$Be_3Al_2(SiO_3)_6$
(f)	Sylvite	KCl
(g)	Phosphorite	$Ca_3(PO_4)_2$
(h)	Chile saltpeter	$NaNO_3$

Q.11 Answer the following question briefly:

(a) Why alkali and alkaline earth metals are among the reactive elements of the periodic table?

(b) Why lime-water turns milky with CO_2 but becomes clear with excess CO_2?

(c) How gypsum is converted into plaster of pairs?

(d) Why 2% gypsum is added in the cement?

(e) Why lime is added to an acidic soil?

(f) How lime and sand are used to make glass?

(g) How lime mortar is prepared?

Ans.

(a) Alkali and alkaline earth metals more electropositive because they can easily loose one and two electrons from their valence shell, respectively. Their electropositive character and low ionization energy make them reactive elements of periodic table.

(b) Lime water turns milky due to the formation of $CaCO_3$

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

Lime water

But in the excess of CO_2 solution becomes clear due to the formation of $Ca(HCO_3)_2$ which is soluble in water

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

in excess

(c) When gypsum is heated at 100°C it looses three quarters (3/4) of its water and resulting products is called as Plaster of Paris.

$$2CaSO_4.2H_2O \xrightarrow{100°C} (CaSO_4)_2.H_2O + 3H_2O$$

Gypsum Plaster of Paris

(d) During the formation of cement when it is in the form of **clinker** then 2% gypsum is added. The addition of gypsum decreases the setting time of cement.

(e) Lime is added to acidic soils because it neutralizes the soil and increase the amount of readily soluble phosphorus.

(f) Lime has ability to react with sand at high temperature and form calcium silicate ($CaSiO_3$) which serves as an important basis for manufacturing of glass.

(g) Lime mortar is prepared by mixing freshly prepared slaked lime (one volume) and (2 or 4 volumes) and water. Lime mortar binds the stones and bricks firmly.

The equations when mortar hardens are:

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

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

$$Ca(OH)_2 + SiO_2 \longrightarrow CaSiO_3 + H_2O$$

DO YOU KNOW?

(1) IA group metals are called alkali metals and IIA are called alkline earth metals.

(2) Group I-B metal are called coinage metals.

(3) Francium is not found is nature. It is prepared artificially in the laboratory.

(4) The name alkali came from Arabic which means "The ash".

(5) Li shows crimson red flame.

(6) Rb and Cs show violet flame.

(7) NaOH is called caustic soda. The word caustic means "which corrodes".

(8) Milk of magnesia is used for the treatment of acidity in the stomach.

(9) Chile saltpeter is found in South Africa.

(10) $CaSO_4 . \frac{1}{2}H_2O$ is called "Plaster of Paris".

(11) Down's cell was designed by J.C. Down.

(12) KO_2 is used in breathing equipments for mountaineers. It has the ability to absorb CO_2 and evolve O_2.

(13) Gypsum is added in cement (2-5%) and it decreases the setting time of cement.

(14) Li_2O is white in colour.

(15) Na_2O_2 is pale yellow.

(16) KO_2 is orange yellow.

(17) Lithium is least reactive of all alkali metals.

CHAPTER 2

s-BLOCK ELEMENTS

MULTIPLE CHOICE QUESTIONS

1. Which one of the following is not an alkali metals:

(a) Fr **(b)** Cs

(c) Rb **(d)** Ra

2. Down's cell is used to prepare:

(a) Na_2CO_3 **(b)** $NaHCO_3$

(c) Na **(d)** NaOH

3. Chile saltpeter has the chemical formula:

(a) $NaNO_3$ **(b)** KNO_3

(c) $Na_2B_4O_7$ **(d)** $Na_2B_4O_7 . 10H_2O$

4. Which ion will have the maximum value of heat of hydration:

(a) Na^+ **(b)** Cs^+

(c) Ba^{+2} **(d)** Mg^{+2}

5. Beryllium metal is as hard as:

(a) Fe **(b)** Cu

(c) Zn **(d)** Diamond

6. One of the following alkali metals is the most reactive which is that:

(a) Cs **(b)** K

(c) Na **(d)** Li

7. Which of the following does not give flame test:

(a) Zn **(b)** Ba

(c) Sr **(d)** Ca

8. Sodium metal cannot be stored in:

(a) Toluene **(b)** Alcohol

(c) Benzene **(d)** Kerosene oil

9. Dolomite is a carbonate of:

(a) Ca and Ba **(b)** Mg and Al

(c) Ca and Mg **(d)** Na

10. Which one of the following configurations corresponds to an alkaline earth metals:

(a) $[Ar]\ 3d^{10}\ 4s^2$ **(b)** $[Ne]\ 3d^2\ 3p^2$

(c) $[Ar]\ 4s^2$ **(d)** $[Ar]\ 3d^{10}\ 4s^1$

11. Dead burnt gypsum is:

(a) $CaSO_4 . 2H_2O$ **(b)** $CaSO_4 . \frac{1}{2} H_2O$

(c) $CaSO_4$ **(d)** $MgSO_4 . 7H_2O$

12. When some water is added to plaster of Paris, it becomes hard and expansion in volume how much:

(a) 1% **(b)** 10%

(c) 4% **(d)** 2%

13. Which of the following compound in formed when Na burn in excess of air:

(a) NaO_2 **(b)** Na_2O_2

(c) Na_2O **(d)** Na_2O_3

14. Chemical formula of slaked lime is:

(a) $CaCO_3$ **(b)** $Ca(OH)_2 . H_2O$

(c) $Ca(OH)_2$ **(d)** $CaSO_4$

15. Which of the following is not a function of sulphur:

(a) Enlarged root system in plants **(b)** Chlorophyll development in leaves

(c) Good yield of crops **(d)** Control of pH of soil

16. Which is produced at the cathode during the electrolysis of brine in Nelson's cell:

(a) H_2 **(b)** Na

(c) Cl_2 **(d)** O_2

17. Which of the following elements can produce H_2 gas when treated with sodium hydroxide:

(a) Be **(b)** Mg

(c) Ca **(d)** Sr

18. The operating temp of Down's cell is:

(a) 800°C **(b)** 600°C

(c) 1800°C **(d)** 900°C

19. $CaMg_3(SiO_3)_4$ is the composition of:

(a) Dolomite **(b)** Gypsum

(c) Calcite **(d)** Asbestos

20. Halite is chemical name of:

(a) KCl **(b)** NaCl

(c) $MgCl_2$ **(d)** $SrCl_2$

answers

1.	(d)	**2.**	(c)	**3.**	(a)	**4.**	(d)	**5.**	(a)
6.	(a)	**7.**	(a)	**8.**	(b)	**9.**	(c)	**10.**	(c)
11.	(c)	**12.**	(a)	**13.**	(b)	**14.**	(c)	**15.**	(d)
16.	(a)	**17.**	(a)	**18.**	(b)	**19.**	(d)	**20.**	(b)

SOLVED EXERCISE

Q.1 Fill in the blanks:

(i) Alkali metals are ———— reactive than alkaline-earth metals.

(ii) Alkali metals decompose water vigorously producing ———— and hydrogen.

(iii) When heated in a current of dry hydrogen, alkaline earth metals form white crystalline ———— of the type MH_2.

(iv) The beryllium hydroxide, like the hydroxide of aluminium is amphoteric, while the hydroxides of the other members of the group IIA are ————.

(v) The elements of the group IA are termed as alkali metals, because their ———— are alkaline.

(vi) Spodumene is an ore of ———— metal.

(vii) Alkali metal nitrates on heating give the corresponding ———— and oxygen (except Li).

(viii) Na_2CO_3, H_2O is the chemical formula of a mineral of sodium which is known as ————.

(ix) Metallic bicarbonates are decomposed on heating into their carbonates, alongwith ———— and ————.

(x) Metal nitrates other than the alkali metals on heating decompose into the corresponding metal ———— alongwith the evolution of nitrogen peroxide and oxygen.

answers

(i)	more	**(ii)**	hydroxide	**(iii)**	hydrides
(iv)	basic	**(v)**	hydroxides, oxides	**(vi)**	lithium
(vii)	nitrite	**(viii)**	natron	**(ix)**	CO_2, H_2O
(x)	oxides				

Q.2 Indicate True or False:

(i) Group IA elements are called alkali metals because their chlorides are alkaline in nature.

(ii) Alkali metals are very good conductor of electricity.

(iii) The hydroxides of alkali metals and alkaline-earth metals are soluble in water.

(iv) Plaster of Paris is a hemihydrate.

(v) Alkali metals have low melting and boiling points as compared to those of alkaline-earth metals.

(vi) Lithium carbonate is decomposed to its oxide, but the carbonates of the other alkali metals are stable towards heat.

(vii) All alkali metal sulphates are insoluble in water.

(viii) Lithium combines with nitrogen to form lithium nitride but other alkali metals do not react with nitrogen.

(ix) Trona is a mineral of lithium.

(x) Alkaline-earth metals are stronger reducing agents than alkali metals.

answers

(i)	False	**(ii)**	True	**(iii)**	True	**(iv)**	True	**(v)**	True
(vi)	True	**(vii)**	False	**(viii)**	True	**(ix)**	False	**(x)**	False

Q.3 Multiple choice questions. Encircle the correct answer:

(i) Which one of the following does not belong to alkaline-earth metals:

(a) Be **(b)** Ra

(c) Ba **(d)** Rn

(ii) The oxides of beryllium are:

(a) Acidic **(b)** Basic

(c) Amphoteric **(d)** None of these

(iii) Which ion will have the maximum value of heat of hydration:

(a) Na^{+} **(b)** Cs^{2+}

(c) Ba^{2+} **(d)** Mg^{2+}

(iv) Which one of the following is not an alkali metals:

(a) Francium **(b)** Caesium

(c) Rubidium **(d)** Radium

(v) Which of the following sulphates is not soluble in water:

(a) Sodium sulphate **(b)** Potassium sulphate

(c) Zinc sulphate **(d)** Barium sulphate

(vi) The element caesium bears resemblance with:

(a) Ca **(b)** Cr

(c) Both of these metals **(d)** None of these metals

(vii) Saltpetre has the chemical formula:

(a) KNO_3 **(b)** KNO_2

(c) $Na_2B_4O_7$ **(d)** $Na_2CO_3 . H_2O$

(viii) The mineral $CaSO_4 . 2H_2O$ has the general name:

(a) Gypsum **(b)** Dolomite

(c) Calcite **(d)** Epsom salt

(ix) Down's cell is used to prepare:

(a) Sodium carbonate **(b)** Sodium bicarbonate

(c) Sodium metal **(d)** Sodium hydroxide

(x) Which element is deposited at the cathode during the electrolysis of brine in diaphragm cell:

(a) H_2 **(b)** Na

(c) Cl_2 **(d)** O_2

answers

(i)	(d)	**(ii)**	(c)	**(iii)**	(d)	**(iv)**	(d)	**(v)**	(d)
(vi)	(d)	**(vii)**	(a)	**(viii)**	(a)	**(ix)**	(c)	**(x)**	(a)

Q.4 (a) Give the names, electronic configurations and occurrence of s-block elements.

(b) Discuss the peculiar behaviour of lithium with respect to the other members of alkali metals.

Ans.

(a) Descriptive question. For details see text book.

(b) Descriptive question. For details see text book.

Q.5 Discuss the trends in chemical properties of compounds like oxides, hydroxides carbonates, nitrates and sulphates of IA and IIA group elements.

Ans. Detailed question. For answer see text book.

Q.6 Compare the chemical behaviour of lithium with magnesium.

Ans. As Li and Mg have diagonal relationship so they resemble in chemical properties which are given below:

(i) On heating in air, both Li and Mg form normal oxides.

(ii) Carbonates and phosphates of both Li and Mg are insoluble in water.

(iii) Carbonates of both Li and Mg are unstable and decompose to form oxides.

$$Li_2CO_3 \longrightarrow Li_2O + CO_2$$

$$MgCO_3 \longrightarrow MgO + CO_2$$

(iv) The nitrates of both Li and Mg decompose to form NO_2 and O_2 leaving a residue of oxides:

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

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

(v) Both Li and Mg react with N_2 to form nitrides:

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

$$3Mg + N_2 \longrightarrow Mg_3N_2$$

(vi) Both Li and Mg give only monoxides Li_2O and MgO.

Q.7 (a) Mention the properties of beryllium in which it does not resemble with its own family.

(b) Why the aqueous solution of Na_2CO_3 is alkaline in nature?

Ans.

(a) Detailed question. Consult text book.

(b) When Na_2CO_3 is dissolved in water

$$Na_2CO_3 + 2H_2O \longrightarrow 2NaOH + H_2CO_3$$

A strong alkali NaOH and a weak acid H_2CO_3 are formed. Due to formation of NaOH, solution of Na_2CO_3 is alkaline in nature.

Q.8 (a) Describe with diagram the manufacture of sodium by Down's cell.

(b) Point out the three advantages of this process.

Ans. Detailed questions. See text book.

Q.9 (a) Compare the physical and chemical properties of alkali metals with those of alkaline-earth metals.

(b) What happens when:

(i) Lithium carbonate is heated.

(ii) Lithium hydroxide is heated to red-hot.

(iii) Beryllium is treated with sodium hydroxide.

(iv) Lithium hydride is treated with water.

Ans.

(a) Detailed questions. See text book.

(b) (i) Lithium carbonate is heated:

$$Li_2CO_3 \xrightarrow{\Delta} \underset{\text{Lithium oxide}}{Li_2O} + CO_2$$

(ii) Lithium hydroxide is heated to red-hot:

$$2LiOH \xrightarrow{\Delta} \underset{\text{Lithium oxide}}{Li_2O} + H_2O$$

(iii) Beryllium is treated with sodium hydroxide:

$$Be + 2NaOH \xrightarrow{\Delta} \underset{\text{Sodium beryllate}}{Na_2BeO_2} + H_2$$

(iv) Lithium hydride is treated with water:

$$LiH + H_2O \xrightarrow{\Delta} \underset{\text{Lithium hydroxide}}{LiOH} + H_2$$

Q.10 Give formulas of the following minerals:

Ans.

(a)	Dolomite	$MgCO_3 . CaCO_3$
(b)	Asbestos	$CaMg_3(SiO_3)_4$
(c)	Halite	$NaCl$
(d)	Natron	$Na_2CO_3 . H_2O$
(e)	Beryl	$Be_3Al_2(SiO_3)_6$
(f)	Sylvite	KCl
(g)	Phosphorite	$Ca_3(PO_4)_2$
(h)	Chile saltpeter	$NaNO_3$

Q.11 Answer the following question briefly:

(a) Why alkali and alkaline earth metals are among the reactive elements of the periodic table?

(b) Why lime-water turns milky with CO_2 but becomes clear with excess CO_2?

(c) How gypsum is converted into plaster of pairs?

(d) Why 2% gypsum is added in the cement?

(e) Why lime is added to an acidic soil?

(f) How lime and sand are used to make glass?

(g) How lime mortar is prepared?

Ans.

(a) Alkali and alkaline earth metals more electropositive because they can easily loose one and two electrons from their valence shell, respectively. Their electropositive character and low ionization energy make them reactive elements of periodic table.

(b) Lime water turns milky due to the formation of $CaCO_3$

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

Lime water

But in the excess of CO_2 solution becomes clear due to the formation of $Ca(HCO_3)_2$ which is soluble in water

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

in excess

(c) When gypsum is heated at 100°C it looses three quarters (3/4) of its water and resulting product is called as Plaster of Paris.

$$2CaSO_4 . 2H_2O \xrightarrow{100^\circ C} (CaSO_4)_2 . H_2O + 3H_2O$$

Gypsum Plaster of Paris

(d) During the formation of cement when it is in the form of **clinker** then 2% gypsum is added. The addition of gypsum increases the setting time of cement.

(e) Lime is added to acidic soils because it neutralizes the soil and increase the amount of readily soluble phosphorus.

(f) Lime has ability to react with sand at high temperature and form calcium silicate ($CaSiO_3$) which serves as an important basis for manufacturing of glass.

(g) Lime mortar is prepared by mixing freshly prepared slaked lime (one volume) and (2 or 4 volumes) of water. Lime mortar binds the stones and bricks firmly.

The equations when mortar hardens are:

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

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

$$Ca(OH)_2 + SiO_2 \longrightarrow CaSiO_3 + H_2O$$

DO YOU KNOW?

1. IA group metals are called alkali metals and IIA are called alkaline earth metals.
2. Group I-B metal are called coinage metals.
3. Francium is not found in nature. It is prepared artificially in the laboratory.
4. The name alkali came from Arabic which means "ash".
5. Li shows crimson red flame.
6. Rb and Cs show violet flame.
7. NaOH is called caustic soda. The word caustic means "which corrodes".
8. Milk of magnesia is used for the treatment of acidity in the stomach.
9. Chile saltpeter is found in South Africa.
10. $CaSO_4 . \frac{1}{2} H_2O$ is called "Plaster of Paris".
11. Down's cell was designed by J.C. Down.
12. KO_2 is used in breathing equipments for mountaineers. It has the ability to absorb CO_2 and evolve O_2.
13. Gypsum is added in cement (2-5%) and it decreases the setting time of cement.

SHORT QUESTIONS

Q.1 What is the difference between slaked lime and un-slaked lime?

Ans. $Ca(OH)_2$ is called slaked lime and CaO is called un-slaked lime. It is also called quick lime.

Q.2 What is the effect of heat on gypsum?

Ans. When gypsum is heated at 100°C, it changes into plaster of Paris.

$$2CaSO_4 . 2H_2O \xrightarrow{100°C} (CaSO_4)_2H_2O + 3H_2O$$

On heating at 400°C, it changes into dead burnt gypsum (anydrous $CaSO_4$).

$$CaSO_4 . 2H_2O \xrightarrow{400°C} CaSO_4 + 2H_2O$$

Q.3 Why lithium shows peculiar behaviour from other alkali metals?

Ans. Atomic number of Li is very small as compared to the other members of IA. The ionic radius of Li^+ is 60 Pm and that of Na^+ is 95 Pm. Li^+ has greater charge density. The nuclear charge of lithium atom is screened by only two electrons of K-shell. Therefore, the properties of Li differ from the other members of the group.

Q.4 Shows that BeO is amphoteric in nature (acidic and basic).

Ans. It reacts with acids and bases to form salt and H_2O.

$$BeO + H_2SO_4 \longrightarrow BeSO_4 + H_2O$$

$$BeO + 2NaOH \longrightarrow Na_2BeO_2 + H_2O$$

Q.5 What is the difference between peroxides and superoxides?

Ans. In the peroxides (H_2O_2, Na_2O_2), the oxidation state of O is -1 and in superoxides (KO_2, RbO_2), the oxidation state of O is $-\frac{1}{2}$.

Per oxides dissolve in H_2O to form NaOH and O_2

$$2Na_2O_2 + 2H_2O \longrightarrow 4NaOH + O_2$$

Superoxides dissolve in H_2O to form hydroxides, hydrogen peroxide and O_2.

$$2KO_2 + 2H_2O \longrightarrow 2KOH + H_2O_2 + O_2$$

Q.6 What is the formula of milk of magnesia and milk of lime, washing soda, soda ash and pearl ash, baking soda?

Ans. Milk of magnesia $Mg(OH)_2$

Milk of lime $Ca(OH)_2$

Washing soda	$Na_2CO_3 . 10H_2O$
Soda ash	Na_2CO_3
Pearl ash	K_2CO_3
Baking soda	$NaHCO_3$

Q.7 What is the action of heat on the following KNO_3, $NaNO_3$, $LiNO_3$, $Ca(NO_3)_2$, $Mg(NO_3)_2$, $CaCO_3$, $Ca(HCO_3)_2$?

Ans. $2KNO_3 \xrightarrow{\Delta} 2KNO_2 + O_2$

$2NaNO_3 \xrightarrow{\Delta} 2NaNO_2 + O_2$

$4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2$

$2Ca(NO_3)_2 \xrightarrow{\Delta} 2CaO + 4NO_2 + O_2$

$2Mg(NO_3)_2 \xrightarrow{\Delta} 2MgO + 4NO_2 + O_2$

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

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

Q.8 How the melting point of NaCl is decreased?

Ans. The melting point of NaCl is 801°C. It is decreased by adding $CaCl_2$ as impurity the melting point of NaCl is decreased to 600°C by adding $CaCl_2$.

Q.9 Which is spectator ion in the Nelson's cell?

Ans. The ion present in the cell and is not oxidised or reduced is called spectator ion. In the Nelson's cell, Na^+ ion is spectator ion.

In the Nelson's cell, H_2O is reduced instead of Na^+. The over all reaction is

$$2NaCl + 2H_2O \longrightarrow 2NaOH + H_2 + Cl_2$$

Q.10 What are two problems faced during the working of Nelson's cell?

Ans. (i) Cl_2 produced can react with OH^- ions in the cold state giving hypochlorite ions.

$$Cl_{2(g)} + 2OH^-_{(aq)} \longrightarrow OCl^-_{(aq)} + Cl^-_{(aq)} + H_2O$$

(ii) OH^- ions may be attracted towards anode where they can be discharged releasing oxygen which can contaminate chlorine and render it impure.

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

Q.11 What is the solution of the two problems faced in the Nelson's cell?

Ans. (i) The first problem is solved by using asbestos diaphragm. This keeps the two solutions separate while allowing Na^+ ions to move towards the cathode which keeps the current flowing through the external circuit.

(ii) The second problem is solved by keeping the level of brine towards anode compartment slightly higher which prevents the possibility of OH^- ions to reach the anode.

Q.12 What are uses of plaster of Paris?

Ans. (i) It is used for making plaster walls and roofs.

(ii) It is used for making plaster bandages.

(iii) It is used to make casts for statuary.

Q.13 What is the function of calcium in plants?

Ans. (i) Calcium present in plants stimulates the development of roots, hairs and entire root system of plants.

(ii) It is also required for development of leaves.

(iii) It is also essential for the optimum activity of microorganisms that produce nitrates.

Q.14 Give three uses of lime?

Ans. (i) It is used to prepare CaC_2 which is used to prepare acetylene gas for welding.

(ii) It is used in the extraction and refining of metals.

(iii) Lime is used for the manufacture of bleaching powder.

Q.15 Why sodium metal is stored in kerosene oil?

Ans. Sodium is highly reactive with air oxygen and moisture.

$$2Na + O_2 \longrightarrow Na_2O_2$$

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

It is stored in kerosene oil, which is mixture of hydrocarbons. Sodium metal does not react with hydrocarbons and does not contact with air oxygen or moisture.

Q.16 What are two similar behaviours of Li and Mg?

Ans. (i) Li and Mg both react with O_2 to form normal oxides:

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

$$2Mg + O_2 \longrightarrow 2MgO$$

(ii) Both Li and Mg react with N_2 to form nitrides:

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

$$3Mg + N_2 \longrightarrow Mg_3N_2$$

Q.17 Write the formula of the ores of alkali metals and alkaline earth metals.

Ans.

Spodumene	$LiAl(SiO_3)_2$
Halite (rock salt)	$NaCl$
Chile saltpeter	$NaNO_3$
Natron	$Na_2CO_3 . H_2O$
Trona	$Na_2CO_3 . 2NaHCO_3 . 2H_2O$
Borax	$Na_2B_4O_7 . 10H_2O$
Carnallite	$KCl . MgCl_2 . 6H_2O$
Sylvite	KCl

Alunite	$K_2SO_4 . Al_2(SO_4)_3 . 4Al(OH)_3$
Beryl	$Be_3Al_2(SiO_3)_6$ / $3BeO . Al_2O_3 . 6SiO_2$ (Oxide ore)
Chrysoberyl	Al_2BeO_4
Magnesite	$MgCO_3$
Dolmite	$MgCO_3 . CaCO_3$
Epsom salt	$MgSO_4 . 7H_2O$ (Epsomite)
Soap stone (talc)	$H_2Mg_3(SiO_3)_4$
Asbestos	$CaMg_3(SiO_3)_4$
Calcite / Aragonite	$CaCO_3$
Gypsum	$CaSO_4 . 2H_2O$
Strontionite	$SrCO_3$
Barite	$BaSO_4$
Whetherite	$BaCO_3$
Phosphrite	$Ca_3(PO_4)_2$ (Phosphate rock)

Q.18 What is lime mortar? Give its function.

Ans. It is made by mixing slaked lime with sand (three volumes) and water to make a thick paste. This material when placed between the stones and bricks hardens or sets, thus building the blocks firmly together.

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

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

$$Ca(OH)_2 + SiO_2 \longrightarrow CaSiO_3 + H_2O$$