Chapter#2
s-Block Elements

1. What are alkali metals? Why are they called so?

Ans: The group IA elements i.e. Li, Na, K, Rb, Cs, Fr are called alkali metals. The name alkali came from Arabic which means 'The Ashes'. The Arabs used this term for these metals because they found that the ashes of plants were composed chiefly of sodium and potassium.

2. What are s-block elements?

Ans: The s-block elements are the metals in Group IA and Group IIA of the periodic table. They are called the s-block elements because s-orbitals are being filled in their outer most shells.

3. What are alkaline earth metals?

Ans: The alkaline-earth metals are beryllium, magnesium, calcium, strontium, barium and radium. They are called alkaline-earth because they produce alkalies in water and are widely distributed in earth's crust.

4. Mention the names and formulae of important minerals of alkali metals.

Ans: **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) $K_2SO_4.Al_2(SO_4)_3.4Al(OH)_3$

5. Give formulae of halite and natron.

Ans:

Halite NaCl

Natron $Na_2CO_3.H_2O$

6. Write any four points mentioning peculiar behaviour of lithium.

Ans: Following are the four points highlighting peculiar behaviour of lithium:

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

2. The lithium salts of anions with high charge density are generally less soluble in water than those of the other alkali metals, e.g. LiOH, LiF, Li_3PO_4, Li_2CO_3.

3. Lithium forms stable complex compounds, althoughcomplex 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.

7. Write any four points mentioning peculiar behaviour of beryllium.

Ans: Following are the four points highlighting peculiar behaviour of beryllium:

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 alkaline earth metals.

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

4. Beryllium in particular is quite resistant towards complete oxidation, even by acids, because of its

BeO coating.

8. Give the reaction of Be with NaOH.

Ans: Beryllium is the only member of its group which reacts with alkalies to give hydrogen. The other members do not react with alkalies.

$$Be\ (s) + 2NaOH \longrightarrow Na_2BeO_2\ (aq) + H_2\ (g)$$

Sodium beryllate

9. Why alkali metals are good reducing agents?

Ans: Since alkali metals have got low ionization energies, so they are strong reducing agents. They are highly electropositive. They react readily with halogens giving alkali metal halides.

10. How alkali metals react with oxygen?

Ans: 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 (white solid).

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

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.

$$Li_2O\ (s) + CO_2\ (g) \longrightarrow Li_2CO_3\ (s)\ \text{Lithium carbonate}$$

Sodium will undergo a similar reaction, but only if the supply of oxygen is limited. In the presence of excess of oxygen, sodium forms the pale yellow peroxide.

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

Potassium, rubidium and caesium react with oxygen to form superoxides (orange yellow). Caesium explodes spontaneously when it is in contact with air or oxygen.

$$K\ (s) + O_2\ (g) \longrightarrow KO_2\ (s)\ \text{Potassium superoxide}$$

11. How alkali metals react with water?

Ans: 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 hydrogen and produce metal hydroxide. The reaction is highly exothermic. The energy produced by the reaction may even ignite the hydrogen.

$$2Na\ (s) + 2H_2O\ (l) \longrightarrow 2NaOH\ (aq) + H_2\ (g)$$

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.

12. How alkali metals react with hydrogen?

Ans: Alkali metals form ionic hydrides with hydrogen.

$$2M\ (s) + H_2\ (g) \longrightarrow 2M^+H^-(s)$$

Rubidium and caesium react violently with hydrogen at room temperature. The other three metals require elevated temperature in order to form the hydride.

13. How alkali metal hydrides react with water?

Ans: Lithium and sodium hydrides are useful sources of hydrogen when treated with water.

$$LiH\ (s) + H_2O\ (l) \longrightarrow LiOH\ (aq) + H_2\ (g)$$

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

14. How lithium reacts with nitrogen and carbon?

Ans: Lithium is the only Group IA metal that combines with nitrogen and carbon to form nitride and carbide, respectively.

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

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

15. How alkali metals react with halogens?

Ans: Alkali metals react easily with halogens to give halides. Lithium and sodium, for example, react slowly with chlorine at room temperature. Molten sodium burns with a brilliant yellow flame

in a chlorine atmosphere to form sodium chloride.

$$2Na\ (s) + Cl_2\ (g) \longrightarrow 2NaCl\ (s)$$

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

16. How alkali metals react with sulphur?

Ans: All alkali metals form their sulphides when treated with molten sulphur. The general reaction is:

$$2M\ (s) + S\ (s) \longrightarrow M_2S\ (s)$$

17. How alkaline-earth metals react with oxygen?

Ans: The alkaline-earth metals burn in oxygen to form oxides or in the case of barium, the peroxide. 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 ºC. Therefore beryllium is not tarnished by atmospheric attack but the metal soon loses the silvery appearance.

$$2Be\ (s) + O_2\ (g) \xrightarrow{800\ ^oC} 2BeO\ (s)$$

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

$$2Mg\ (s) + O_2\ (g) \longrightarrow 2MgO\ (s)$$

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.

$$Ba\ (s) + O_2\ (g) \xrightarrow{500\text{-}600\ ^o\ C} BaO_2\ (s)\ \text{Barium peroxide}$$

18. What are alkali metals? Why are they called so?

Ans: Group IA elements are called alkali metals because of their property to form strong alkalies with water.

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

19. What are alkaline earth metals? Why are they called so?

Ans: Group IIA elements are called alkaline earth metals because of their presence in the earth's crust and alkaline character.

20. Why alkali and alkaline earth metals are the most reactive elements of the periodic table?

Ans: Since ionization energies of alkali and alkaline earth metals are very low and they have large atomic sizes, therefore, they are the most reactive elements of the periodic table.

21. How alkaline-earth metals react with hydrogen?

Ans: Hydrides are produced by treating the molten alkaline earth metals with hydrogen, usually under high pressures. Magnesium reacts with hydrogen at high pressure and in the presence of a catalyst (MgI_2) forming magnesium hydride.

$$Mg\ (s) + H_2\ (g) \xrightarrow{Pressure/MgI_2} MgH_2\ (g)$$

$$Ca\ (s) + H_2\ (g) \longrightarrow CaH_2\ (g)$$

22. How alkaline earth metals react with nitrogen? What are the products of hydrolysis of their nitrides?

Ans: All Group II-A elements react with nitrogen on heating giving nitrides. For example, magnesium reacts with nitrogen to give magnesium nitride.

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

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

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

23. How alkaline earth metals react with sulphur?

Ans: With sulphur, magnesium gives magnesium sulphide, MgS. The other Group II-A metals also react similarly.

$$Mg + S \longrightarrow MgS$$

24. How alkaline earth metals react with halogens?

Ans: All group II-A elements react directly with halogens giving halides of the type MX_2 e.g.

$$Ca\ (s) + Cl_2\ (g) \longrightarrow CaCl_2\ (g)$$

25. How alkaline earth metals react with water?

Ans: 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.

$$\underset{\text{Steam}}{Mg\ (s) + H_2O\ (g)} \xrightarrow{100\ ^oC} MgO\ (s) + H_2\ (g)$$

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

$$M\ (s) + 2H_2O\ (l) \xrightarrow{100\ ^oC} M(OH)_2\ (s) + H_2\ (s)$$

26. Why KO_2 is used in breathing equipment by mountaineers?

Ans: Potassium superoxide (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.

$$4KO_2\ (s) + 2CO_2\ (g) \longrightarrow 2K_2CO_3\ (s) + 3O_2\ (g)$$

27. What is the trend of solubility of alkaline earth metal oxides in water?

Ans: The solubility of alkaline earth metal oxides in water increases down the group. BeO and MgO are insoluble but CaO, SrO and BaO are soluble and react with water to form the corresponding hydroxides.

28. What is the trend of basic character of oxides of alkaline earth metals?

Ans: The basic character of the oxides of alkaline earth metals increases down the group. The tendency for group IIA oxides to form alkaline solution is relatively less than that of alkali metals.

29. The reaction of an alkali metal oxide with water is an acid-base reaction and not an oxidation reduction reaction. Justify.

Ans: 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.

$$Li_2O\ (s) + H_2O\ (l) \longrightarrow 2LiOH\ (aq)$$

$$2Na_2O_2\ (s) + 2H_2O\ (l) \longrightarrow 4NaOH\ (aq) + O_2\ (g)$$

$$O^{2-}\ (aq) + H_2O\ (l) \longrightarrow 2OH^-\ (aq)$$

30. Justify that BeO is amphoteric in nature.

Ans: BeO is amphoteric in nature since it reacts with both acids and bases.

$$BeO\ (s) + H_2SO_4\ (aq) \longrightarrow BeSO_4\ (s) + H_2O\ (l)$$

$$BeO\ (s) + 2\ NaOH\ (aq) \longrightarrow Na_2BeO_2\ (aq) + H_2O\ (l)$$

Sodium beryllate

31. What is the effect of heat on alkali and alkaline earth metal hydroxides?

Ans: 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.

$$2LiOH\ (s) \longrightarrow Li_2O\ (s) + H_2O\ (l)$$

$$Mg\ (OH)_2\ (s) \longrightarrow MgO\ (s) + H_2O\ (l)$$

32. What is lime water and milk of magnesia?

Ans: A saturated solution of $Ca(OH)_2$ in water is called lime water and is used as a test for CO_2. A suspension of $Mg(OH)_2$ in water is called milk of magnesia and it is used for the treatment of acidity in stomach.

33. Why lime water turns milky by passing CO_2 gas but becomes clear with excess of CO_2?

Ans: Lime water turns milky by passing CO_2 gas due to formation of insoluble $CaCO_3$

$$Ca(OH)_2\ (aq) + CO_2\ (g) \longrightarrow CaCO_3\ (s) + H_2O\ (l)$$

On further addition of CO_2 insoluble $CaCO_3$ changes into soluble $Ca(HCO_3)_2$

$$CaCO_3\ (s) + H_2O\ (l) + CO_2\ (g) \longrightarrow Ca(HCO_3)_2(aq)$$

34. What is the trend of solubility of alkaline earth metal hydroxides?

Ans: The solubility of alkaline earth metal hydroxides in water increases down the group. $Be(OH)_2$ is quite insoluble. $Mg(OH)_2$ is sparingly soluble while $Ba(OH)_2$ is more soluble. This increase in solubility is due to low lattice energy of hydroxides which is, in turn, due to higher ionic size.

35. Tell about stability of alkali metal carbonates and their solubility in water.

Ans: 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. The decomposition is made easy because the electrostatic attraction in converting from carbonate to oxide is considerable.

36. Why is it difficult to decompose potassium carbonate?

Ans: In case of large cation like K^+ in K_2CO_3, the gain in electrostatic attraction is relatively much less and the decomposition is difficult.

37. What is washing soda?

Ans: At temperature below 35.2°C, Na_2CO_3 crystallizes out from water as $Na_2CO_3.10H_2O$, which iscalled washing soda.

38. Why the solution of sodium carbonate in water is basic in nature?

Ans: The solution of Na_2CO_3 in water is basic due to hydrolysis of carbonate ion.

$$Na_2CO_3\ (s) + 2H_2O \longrightarrow 2NaOH\ (aq) + H_2CO_3\ (aq)$$

NaOH is a strong base and H_2CO_3 is a weak acid. So the solution is alkaline in nature.

39. Solution of Na_2O in water is alkaline. Justify the statement.

Ans: When Na_2O is dissolved in water, it produces NaOH which is a strong alkali. Hence, solution of Na_2O in water is alkaline.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

40. What is the action of litmus with aqueous solution of Na_2CO_3?

Ans: The aqueous solution of $Na_2CO_3$is alkaline due to formation of strong alkali NaOH and it turns litmus solution red.

$$Na_2CO_3\ (s) + 2H_2O \longrightarrow 2NaOH\ (aq) + H_2CO_3\ (aq)$$

41. Mention the trend of solubility of alkaline earth metal carbonates.

Ans: Unlike the alkali metal carbonates, the alkaline earth metal carbonates are only very 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.

$$CaCO3\ (s) \longrightarrow CaO\ (s) + CO_2\ (g)$$

42. What is the relation between ease of decomposition and size of ion?

Ans: 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.

43. Tell about the solubility of nitrates of both alkali and alkaline earth metals.

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

44. What is the effect of heat on nitrates of alkali and alkaline earth metals?

Ans: Nitrates of Li, Mg, Ca and Ba decompose on heating to give O_2, NO_2andthe metallic oxide, whereas, nitrates of Na and K decompose to give different products.

$$4LiNO_3\ (s) \longrightarrow 2Li_2O\ (s) + 4NO_2\ (s) + O_2\ (g)$$

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

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

$$2NaNO_3\ (s) \longrightarrow 2NaNO_2\ (s) + O_2\ (g)$$

45. Prove that decomposition of lithium nitrate gives different products than nitrates of other alkali metals?

Ans: Decomposition of lithium nitrate gives different products than nitrates of other alkali metals as shown below:

$$4LiNO_3\ (s) \longrightarrow 2Li_2O\ (s) + 4NO_2\ (s) + O_2\ (g)$$

$$2NaNO_3\ (s) \longrightarrow 2NaNO_2\ (s) + O_2\ (g)$$

46. What happens when:

a. Li_2CO_3 is heated
b. Na_2CO_3 is heated

Ans: a. $Li_2CO_3 \longrightarrow 2Li_2O + CO_2$
b. $2Na_2CO_3 \longrightarrow$ No decomposition on heating

47. How Plaster of Paris is prepared? OR How gypsum is converted into Plaster of Paris?

Ans: 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.

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

48. Tell about solubility of alkali and alkaline earth metal sulphates.

Ans: All the alkali metals give sulphates and they are all soluble in water. The solubilities of sulphates of alkaline earth metals gradually decrease down the group. $BeSO_4$ and $MgSO_4$ are fairly soluble in water. $CaSO_4$ is slightly soluble, while $SrSO_4$ and $BaSO_4$ are almost insoluble.

49. What are the advantages of Down's cell?

Ans: Following are the advantages of Down's cell:

(a) The metallic fog is not produced.

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

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

50. What are the problems faced and their solutions during the working of Nelson's

cell/diaphragm cell?

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

1. Chlorine produced can react with hydroxide ions in cold giving hypochlorite 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.

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 keeps the current following through the external current.

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 prevents the possibility of hydroxides ions to reach the anode.

51. What is the role of gypsum in plant growth?

Ans: 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.

52. What is the role of sulphur in plant growth?

Ans: 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 not a constituent of chlorophyll, plants deficient in sulphur exhibit a pale green colour. The root systems of several plants have 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.

53. What is dead burnt gypsum?

Ans: Gypsum must not be heated too strongly as the anhydrous salt is then formed which absorbs water slowly. Such plaster is called 'Dead burnt'.

54. What are the uses of Plaster of Paris?

Ans: Following are the uses of Plaster of Paris:

1. Plaster of Paris is used for making plaster walls, casts of statuary, coins, etc.

2. It is used in surgery, Plaster of Paris bandages are used for holding in place fractured bones after they have been set.

55. Mention the varieties of Plaster of Paris.

Ans: Following are the varieties of Plaster of Paris:

(1) Cement Plaster

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

(2) Hard Finish Plasters

These are made by the calcination 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. Gypsum is also used as filler in paper industries.

56. Why 2% gypsum is added in cement?

Ans: About 2% of gypsum is added during grinding of cement clinkers which prevents the cement from hardening too rapidly. It increases the setting time of cement.

57. How acidic soil is treated?

Ans: Large quantities of calcium oxide are used in agriculture for neutralizing acidic soils. It has been found that application of lime (CaO) to acidic soils increases the amount of readily soluble

phosphorus.

58. What is the role of calcium in plant growth?

Ans: 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 microorganisms that produce nitrates. The effect of calcium on the supply of available phosphorus in the soil is of 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.

59. What is lime mortar?

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

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

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

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

60. What is white-wash?

Ans: A suspension of the calcium hydroxide is called milk of lime and is used as a white-wash.

61. How lime and sand are used to make glass?

Ans: The ability of lime to react with sand at high temperature forming calcium silicate ($CaSiO_3$) serves as an important basis for glass manufacture. Lime is added in sand and on heating gets converted to calcium silicate. Addition of lime in glass increases the hardness and chemical durability of glass.