CH10 – S BLOCK ELEMENTS

Q 10.1 What are the common physical and chemical features of alkali metals?

Ans :

Physical properties:

(1) The alkali metal is soft, so we can cut them easily. We can able to cut the sodium metal even by using a knife.

(2) Generally, the alkali metal is lightly coloured, and mostly, they appear as silvery white.

(3) Its atomic size is larger, so its density is low. The density of the alkali metal increases down in the group from Li to Cs, except for K, which has low density than sodium.

(4) Alkali metal is weak in its metallic bonding, so they have low boiling and melting points.

(5) The salts present in alkali metals exposes colour to flames because the heat of the flame excites an electron, which is located on the outer orbital, to a higher energy level. In this excited state of electrons, getting reversed back to the ground level, the emission of excess energy in the form of radiation falls into the visible region.

(6) Metals like K and Cs lose electrons when they get irradiated with light and also display a photoelectric effect.

Chemical properties:

(1) Alkali metal reacts with water and forms oxides and hydroxides. So, the reaction will be more spontaneous while moving down the group.

(2) Alkali metal reacts with water and forms dihydrogen and hydroxides.

General reaction: $2M + 2H_20$

$2M^+ + 2OH^- + H_2$

(3) Dihydrogen reacts with alkali metals and forms metal hydrides. The hydrides from this have higher melting points, and they are solids which are ionic.

$2M + H_2$

$2M^+ H^-$

(4) Alkali metals directly react with halogens and form ionic halides, except for Li.

$2M + CL_2$

$2MCl$ (M = Li ,K, Rb, Cs)

It has the ability to easily distort the cloud of the electron, which is around the –ve halide ion, because lithium-ion is smaller in size. Hence, Lithium halide is naturally covalent.

(5) Alkali metals are very stronger reducing agents. This increases as we move down the group, except for lithium. Due to its high hydration energy, it results in a strong reducing agent among all alkali metals.

(6) To result in a blue-coloured solution (deep blue), which is naturally conducting, they get dissolved in liquid ammonia.

$M + (x + y) NH_3$

$[M(NH_3)_x]^+ + [e(NH_3)_y]^-$

Q10.2

Discuss the general characteristics and gradation in properties of alkaline earth metals

Ans 10.2

General characteristics:

(i) (Noble gas) ns^2 is the electronic configuration of alkaline earth metal.

(ii) To occupy the nearest inert gas configuration, these metals lose two of their electrons, and so their oxidation state is +2.

(iii) The ionic radii and atomic radii are smaller than alkali metals. When they move down towards the group, there is an increase in ionic radii and atomic radii due to a decrease in effective nuclear charge.

(iv) The ionisation enthalpy is low because the alkaline earth metals are larger in size. The first ionisation enthalpy is higher than the metals of group 1.

(v) They appear in lustrous and silvery white. They are soft as alkali metals.

(vi) Factors that cause alkaline earth metals to contain high boiling and melting points:

(*) The atoms of alkali metals are larger than that of alkaline earth metals.

(*) Strong metallic bonds are formed by two valence electrons.

(vii) Ca- brick red, Sr- crimson red and Ba-apple green result in colours to flames.

Electrons are bounded strongly to get excited in Be and Mg. Therefore, they do not expose any colours to the flame. Alkali metals are more reactive than alkaline earth metals.

Chemical properties:

(i) Reaction with air and water: Due to the formation of an oxide layer on their surface, beryllium and Mg are most inert to water and air.

(a) Beo and Be_3N_2 are formed when powdered Be is burnt in the air.

(b) For the formation of MgO and Mg_3N_2, Mg is burnt in the air with dazzling sparkle since Mg is more electropositive.

(c) The formation of respective nitrides and oxides is by instant reaction of Sr, Ca, and Ba with air.

(d) Ca, Sr, and Ba can able to react vigorously even with water which is cold.

(ii) When they react with halogens, halides are formed at high temperatures.

$M + X_2$

MX_2 (X = F,CL,Br,I)

(iii) Except Be, all the alkaline earth metals react with hydrogen to form hydrides.

(iv) Alkaline earth metals instantly react with acids to form salts with the liberation of hydrogen gas.

$M + 2HCl$

$MCl_2 + H_{2(X)}$

(e) Reducing Nature: Alkaline earth metals are strong reducing agents like alkali metals, but the reducing power is less when compared to alkali metals. In general, the reducing character increases from top to bottom.

(f) Solutions in liquid ammonia: Alkaline earth metals dissolve in liquid ammonia to give deep blue-black solutions like alkali metals.

Q 10.3

Why are alkali metals not found in nature?

Ans 10.3

Sodium, caesium, lithium, francium, potassium and rubidium all together comprise the alkali metals. They consist of only one electron on their valence shell, which is lost easily due to their low ionising energies. So, alkali metals are not found naturally in their elemental state as they are highly reactive.

Q 10.4

Find out the oxidation state of sodium in Na_2O_2.

Ans 10.4

Let the oxidation state of Na be y.

In the case of peroxides, the oxidation state of oxygen is -1.

Therefore,

$2(y) + 2(-1) = 0$

$2y - 2 = 0$

$2y = 2$

$y = +1$

Therefore, the oxidation state of Na is +1.

Q 10.5

Explain why is sodium less reactive than potassium.

Ans 10.5

On moving down the group in the alkali metals, the size of the atom increases and the effect of the nuclear charge gets decreases. Due to these factors, the electron of potassium, which is located outer, gets lost easily as compared to Na. Therefore, potassium is more highly reactive than sodium.

Q 10.6

Compare the alkali metals and alkaline earth metals with respect to (i) ionisation enthalpy (ii) basicity of oxides, and (iii) solubility of hydroxides.

Ans 10.6

Sr. No	Alkaline earth metals	Alkali Metals
a)	Solubility of hydroxide: They are less soluble compared to alkali metals, as it has high lattice energy and higher charge densities that account for higher lattice.	Solubility of hydroxide: They are more soluble compared to alkaline earth metals.
b)	Ionization Enthalpy: They have a smaller atomic size and higher effective nuclear charge compared to alkali metals, which causes their 1st ionisation enthalpy to be higher than that in alkali metals, but the 2nd ionisation enthalpy is less than that of alkali metals.	Ionization Enthalpy: They have large atomic sizes compared to alkaline earth metals, so they have less 1st ionisation enthalpy, so they lose valance electrons very easily.
c)	Basicity of oxides: Their oxides are quite basic but less as compared to those of alkali metals, as they are less electro-positive than alkali metals.	Basicity of oxides: Their oxides are basic in nature, as they are highly electropositive, which makes their oxides highly ionic.

Q 10.7

In what ways lithium shows similarities to magnesium in its chemical behaviour?

Ans 10.7

Similarities between lithium and magnesium:

(i) Lithium and magnesium react slowly with cold water.
(ii) Oxides of lithium and magnesium are less soluble in H_2O. Also, the hydroxides of both decompose at high temperatures.

(iii) Nitrides are formed from both lithium and magnesium when they react with N_2.

(iv) Neither Li nor Mg forms superoxides or peroxides.
(v) Both the carbonates of lithium and magnesium are naturally covalent. They decompose on heating.

(vi) They do not form bicarbonates which are solid.
(vii) Both $MgCl_2$ and LiCl are soluble in ethanol because they are naturally covalent.
(viii) Both $MgCl_2$ and LiCl are naturally deliquescent. They crystallise as hydrates from aqueous solutions.
Eg.,
and
Q 10.8
Explain why can alkali and alkaline earth metals not be obtained by chemical reduction methods.
Ans 10.8
By using a stronger reducing agent, the oxides of metals get reduced by the process called chemical reduction. Alkaline earth metals and alkali metals are strong among the reducing agents. No stronger reducing agent is available than them. Therefore, alkaline earth metals and alkalis cannot be obtained by chemical reduction of their oxides.
Q 10.9
Why are potassium and caesium, rather than lithium, used in photoelectric cells?
Ans 10.9

Lithium, potassium, and caesium are all alkali metals. But still, potassium and caesium are used in photoelectric cells and not Lithium because Li is smaller in size when compared to the other two.
On the other hand, caesium and potassium have low ionisation energy. Therefore, they lose electrons easily. This property is utilised in photoelectric cells.

Q 10.10

When an alkali metal dissolves in liquid ammonia, the solution can acquire different colours. Explain the reasons for this type of colour change.

Ans 10.10

When the alkali metal is dissolved in liquid ammonia, a deep blue-coloured solution is formed.
The ammoniated electrons absorb energy corresponding to the red region of visible light. Therefore, the transmitted light is deep blue in colour.
Clusters of metal ions are formed at higher concentrations (3M) which causes the solution to attain a copper-bronze colour and a metallic lustre.

Q 10.11

Beryllium and magnesium do not give colour to flame, whereas other alkaline earth metals do so. Why?

Ans 10.11

The valence electrons get excited to a higher energy level when an alkaline earth metal is heated. It radiates energy which belongs to the visible region when this excited electron comes back to its energy level, which is low. The colour is observed here. The electrons are strongly bound in the beryllium and magnesium. The energy required to excite these electrons is very high. When the electron reverts back to its original position, the energy released does not fall in the visible region. Hence, no colour is seen in the flame.

Q 10.12

Discuss the various reactions that occur in the Solvay process.

Ans 10.12

The process of preparing sodium carbonate is called the Solvay process.
Sodium hydrogen carbonate is formed when carbon dioxide gas is bubbled

through a brine solution saturated with ammonia. The obtained sodium hydrogen carbonate is then converted into sodium carbonate.
(i) Brine solution is saturated with ammonia.
This ammoniated brine is filtered for purity removal.
(ii) When carbon dioxide reacts with ammoniated brine, it results in the formation of insoluble sodium hydrogen carbonate.
$(NH_4)_2CO_3 + H_2O + CO_2 \rightarrow 2NH_4HCO_3$
$NH_4HCO_3 + NaCl \rightarrow NaHCO_3\downarrow + NH_4Cl$
(iii) $NaHCO_3$ is obtained by the solution, which contains crystals of $NaHCO_3$ is filtered.
(iv) $NaHCO_3$ is heated strongly to convert it into Na_2CO_3.
$2\ NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$
(v) The carbon dioxide required for the reaction can be obtained by heating limestone. CaO dissolves in water to form calcium hydroxide, which is then transferred to the ammonia recovery tower.
$CaCO_3 \rightarrow CaO + CO_2$
$CaO + H_2O \rightarrow Ca(OH)_2$
(vi) Ammonia is recovered when the filtrate, which is removed after $NaHCO_3$, is mixed with $Ca(OH)_2$ and heated.
The overall reaction taking place in the Solvay process is
Q 10.13
Potassium carbonate cannot be prepared by the Solvay process. Explain why.
Ans 10.13
The Solvay process is not applicable to the preparation of potassium carbonate because potassium carbonate is soluble in water, and it doesn’t precipitate out like sodium bicarbonate.
Q 10.14
Why is Li_2CO_3 decomposed at a lower temperature, whereas Na_2CO_3 at a higher temperature?
Ans 10.14

The electropositive character increases while moving down in the group of alkali metals which results in an increase in the stability of alkali carbonates. Generally, lithium carbonate is not stable when it reacts to heat because lithium carbonate is covalent. Due to the smaller size of lithium-ion, it polarises large carbonate ions, which results in the formation of stable lithium oxide.

This is why sodium carbonate decomposes at high temperatures and lithium carbonate decomposes at low temperatures.

Q 10.15

Compare the solubility and thermal stability of the following compounds of the alkali metals with those of the alkaline earth metals. (a) Nitrates (b) Carbonates (c) Sulphates.

Ans 10.15

(a) Nitrates

Thermal stability

Except for $LiNO_3$, the nitrates of alkali metals get decomposed while heating strongly, which results in the formation of nitrites.

$LiNO_3$, on decomposition, gives oxide.

Like lithium nitrate, alkaline earth metal nitrates also decompose to give oxides.

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

Solubility

Nitrates of both group 1 and group 2 metals are soluble in water.

(b) Carbonates

Thermal stability

The alkali metal carbonates are very stable to heat. But carbonates of lithium decompose and results in the formation of lithium oxide while heating. The carbonates of alkaline earth metals also decompose, which results in the formation of carbon dioxide and oxide while heating.

Solubility

Except for Li_2CO_3, the carbonates of alkali metals are soluble in water and also, while we move down the group, the solubility increases. Carbonates of alkaline earth metals are less soluble in water. The solubility of carbonates in water decreases as the atomic number of the metal ion increases.

(c) Sulphates

Thermal stability

Sulphates of both group 1 and group 2 metals are stable towards heat.

Solubility

Sulphates of alkali metals are soluble in water. But the sulphates of alkaline earth metals show various activities.

$CaSO_4$ Sparingly soluble

$BaSO_4$ Insoluble

$BeSO_4$ Fairly soluble

$SrSO_4$ Insoluble

$MgSO_4$ Soluble

Q 10.16

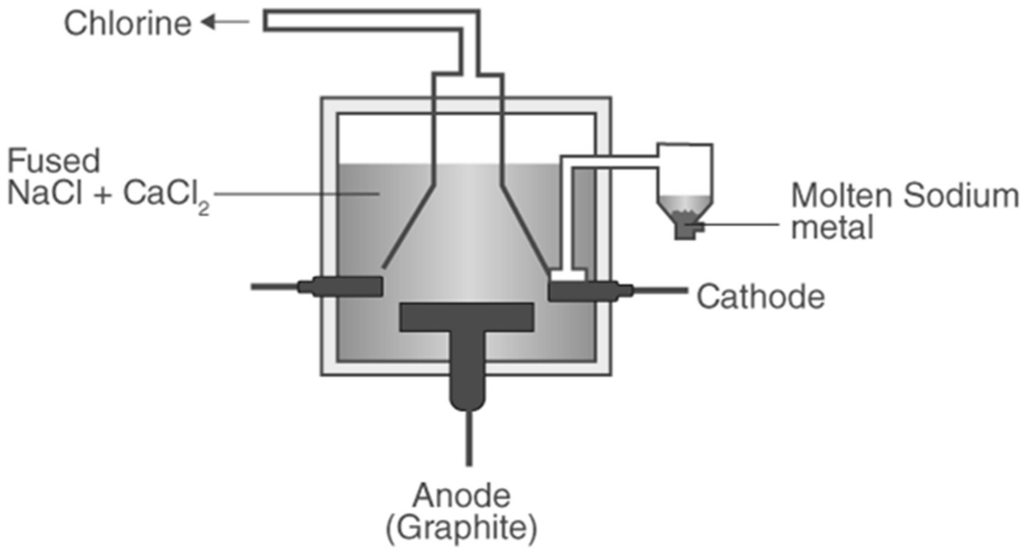
Starting with sodium chloride, how would you proceed to prepare (i) sodium metal (ii) sodium hydroxide (iii) sodium peroxide (iv) sodium carbonate?

Ans 10.16

(i) Sodium metal

Sodium chloride can be converted into sodium by the Downs process.

It can be achieved by electrolysis of fused $CaCl_2$ (60 %) and NaCl (40%) at 1123 K in a special apparatus (Downs cell).



A graphite block is an anode, while steel is made the cathode. Metallic Ca and Na are formed at the cathode. Molten Na is supported by dipping in kerosene.

NaCl

$Na^+ + Cl^-$

(Molten)

At Cathode: $Na^+ + e^-$

Na

At Anode: Cl^-

$Cl + e^-$

Cl + Cl

Cl_2

(ii) Sodium hydroxide

By electrolysing a solution of sodium chloride, we can get Sodium hydroxide. This process is commonly known as the Castner-Kellner process. The process is carried out using a mercury cathode and a carbon anode. Sodium metal, deposited at the cathode, forms an Amalgam by combining with Mercury.

Cathode: $Na^+ + e^-$

Na-Amalgam

$2Na\text{-}Hg + 2H_2O \rightarrow 2NaOH + H_2 + Hg$

Anode: Cl^-

$Cl_2 + e^-$

(iii) Sodium peroxide

After Na metal is obtained from the Downs process, it is heated on Aluminium trays in the presence of air (without CO_2) to form Sodium peroxide.

$2Na + O_{2(air)}$

Na_2O_2

(iv) Sodium carbonate

Sodium hydrogen carbonate is obtained as a precipitate by reacting sodium chloride with ammonium hydrogen carbonate. The resultant crystals can be heated to obtain Sodium Carbonate.

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

$(NH_4)_2CO_3 + H_2O + CO_2$

$2NH_4HCO_3$

$2NH_4HCO_3 + NaCl$

$NH_4Cl + NaHCO_3$

The resultant crystals can be heated to obtain Sodium Carbonate.

$2NaHCO_3$

$Na_2CO_3 + CO_2 + H_2O$

Q 10.17

What happens when (i) magnesium is burnt in air (ii) quick lime is heated with silica (iii) chlorine reacts with slaked lime (iv) calcium nitrate is heated?

Ans 10.17

(i) When Magnesium is burnt in air, it does so with a dazzling bright light resulting in the formation of Mg_3N_2 and MgO.

$2Mg + O_2$

$2MgO$

$3Mg + N_2$

Mg_3N_2

(ii) Silica (SiO_2) combines with Quick lime (CaO), resulting in the formation of a Slag.

$CaO + SiO_2$

$CaSiO_3$

(iii) Bleaching powder is formed when chlorine is made to react with slaked lime.

$Ca(OH)_2 + 2Cl_2 \rightarrow CaCl_2 + Ca(OCl)_2 + 2H_2O$

(iv) Calcium nitrate, when heated, undergoes decomposition to form calcium oxide.

$2Ca(NO_3)_{2(s)}$

$2CaO_{(s)} + 4NO_{2(g)} + O_{2(g)}$

Q 10.18

Describe two important uses of each of the following : (i) caustic soda (ii) sodium carbonate (iii) quicklime.

Ans 10.18

(i) Caustic soda

(a) Heavily used in soap industries.

(b) Common reagents in laboratories.

(ii) Sodium carbonate

(a) Finds uses in both soap and glass industries.

(b) Also finds use as a water softener.

(iii) Quick lime

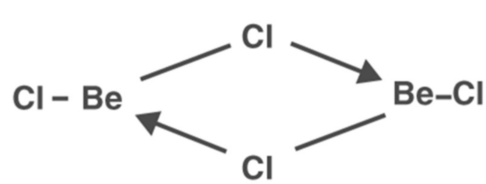
(a) Finds use as a primary material for manufacturing slaked lime.

(b) It helps in the manufacture of cement and glass.

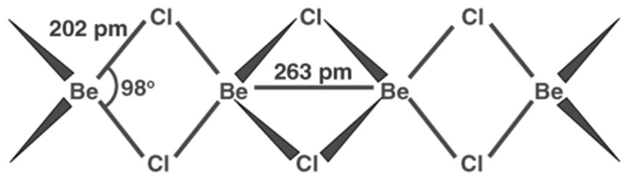
Q 10.19

Draw the structure of (i) $BeCl_2$ (vapour) (ii) $BeCl_2$ (solid).

(i) $BeCl_2$ has a linear structure and exists as a monomer in the vapour state.



(ii) In the solid phase, $BeCl_2$ is a polymer.



Q 10.20

The hydroxides and carbonates of sodium and potassium are easily soluble in water, while the corresponding salts of magnesium and calcium are sparingly soluble in water. Explain.

Ans 10.20

Since the atomic sizes of magnesium and calcium are smaller than that of sodium and potassium, calcium and magnesium form carbonates and hydroxides with higher lattice energies. Thus, they are only sparingly soluble, whereas potassium and sodium are readily soluble due to low lattice energies.

Q 10.21

Describe the importance of the following : (i) limestone (ii) cement (iii) plaster of Paris.

Ans 10.21

Uses of cement:

- Bridge construction
- Plastering
- The most important ingredient in concrete

Uses of Plaster of Paris:

- Used to make casts and moulds
- Used to make surgical bandages

Uses of limestone:

- Preparation of cement and lime
- As a flux in iron ore smelting

Q 10.22

Why are lithium salts commonly hydrated, and those of the other alkali ions are usually anhydrous?

Ans 10.22

Since Lithium has the smallest size among all the alkali metals, it can easily polarise water molecules. Thus, the smaller the size of the ion, the greater its ability to polarise water molecules.

Hence, trihydrated Lithium Chloride and other Lithium salts can be easily polarised. Due to this reason, other alkali metal ions can only form anhydrous salts.

Q 10.23

Why is LiF almost insoluble in water, whereas LiCl is soluble not only in water but also in acetone?

Ans 10.23

The lattice energy of LiF is very high compared to the hydration energy of LiF due to the small size of Li^+ ions and F^- ions. So LiF is insoluble in water. For a substance to dissolve in water, its hydration energy must be greater than its lattice energy. In the case of LiCl, the hydration energy is higher than the lattice energy. Hence, LiCl is water soluble. Due to higher polarisation, LiCl has some covalent character. Hence, it is soluble in non-polar solvents such as acetone.

Q 10.24

Explain the significance of sodium, potassium, magnesium and calcium in biological fluids.

Ans 10.24

Sodium (Na):

They are found in our blood plasma and the interstitial fluids around the cells. They help in

(a) Transmission of nerve signals.

(b) They regulate the flow of water across the membranes of the neighbouring cells.

(c) Transport sugars and amino acids from and to cells.

Potassium (K):

They are found mostly in cell fluids in greater quantities.

They help in

(a) Activating enzymes.
(b) Oxidising glucose to form ATP.
(c) Transmitting nerve signals.

Magnesium (Mg) and calcium (Ca):

They are also called macro-minerals, named so because of their abundance in our bodies. Mg helps in
(a) Relaxing nerves and muscles.
(b) Building and strengthening bones.
(c) Maintaining blood circulation in our body.

Ca helps in
(a) Coagulation of blood.
(b) Maintaining homeostasis.

Q 10.25

What happens when:
(i) Sodium metal is immersed in water?
(ii) Sodium metal is heated in a free supply of air?
(iii) Sodium peroxide gets dissolved in water?

Ans 10.25

(i) Sodium reacts to form NaOH and H_2 gas when it is dropped in water. The reaction occurs as shown below:

$2Na_{(s)} + 2H_2O_{(l)}$

$2NaOH_{(aq)} + H_{2(g)}$

(ii) Sodium peroxide is formed when sodium reacts with oxygen while heating it in the presence of air. The reaction proceeds as shown below:

$2Na_{(s)} + O_{2(s)}$

$Na_2O_{2(s)}$

(iii) NaOH and water are formed as a result of the hydrolysis of Sodium peroxide when it is dissolved in water.

$Na_2O_{2(s)} + 2H_2O_{(l)}$

$2NaOH_{(aq)} + H_2O_{2(aq)}$

Q 10.26

Comment on each of the following observations:

(a) The mobilities of the alkali metal ions in an aqueous solution are
$Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$
(b) Lithium is the only metal to form a nitride directly.
(c) E^0 for $M^{2+}_{(aq)} + 2e^-$
M(s) (where M = Ca, Sr or Ba) is nearly constant.

Ans 10.26

(a) The ionic and atomic sizes of the metals tend to increase while going down the alkali group.

The increasing order of the ionic sizes of the alkali metal ions is as shown below:

$Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$

The smaller the size of an ion, the greater its ability to get hydrated. Li^+ ion gets heavily hydrated since it is the smallest in size, whereas Cs^+ has the largest size and is the least hydrated ion. The alkali metal ions, when arranged in the decreasing order of their hydrations, are as shown below:

$Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

The higher the mass of a hydrated ion, the lesser its ionic mobility. Thus, hydrated Li^+ is the least mobile ion, whereas hydrated Cs^+ is the most mobile ion.

The ionic mobility of the alkali metal ions is in the following order:

$Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$

(b) The only metal that can form a nitride directly is Lithium because Li^+ has a smaller size and is easily compatible with the N^{3-} ion. Thus, the lattice energy released is very high, which is enough to overcome the amount of energy needed to form N^{3-} ion.

(c) Electrode potential (E°) of any M^{2+}/M electrode is decided by three factors:

(i) Enthalpy of hydration
(ii) Enthalpy of vaporisation
(iii) Ionisation enthalpy

The cumulative effect of these factors on Ba, Sr, and Ca is almost the same. As a result, their electrode potentials are also the same.

Q 10.27

State why

(a) a solution of Na_2CO_3 is alkaline in nature.

(b) alkali metals are prepared by electrolysis of their fused chlorides.

(c) Sodium is found to be more useful than potassium.

Ans 10.27

(a) Sodium bicarbonate and sodium hydroxide are the end products when Na_2CO_3 is hydrolysed. Since the product is alkaline in nature, a solution of Na_2CO_3 is considered to be alkaline in nature.

(b) Chemical reduction cannot be used to prepare alkali metals since they themselves are reducing in nature. Alkali metals are highly electropositive and thus cannot be prepared by displacement reactions. Since they also react with water, these alkali metals cannot be prepared by electrolysis of their aqueous solutions. Thus, alkali metals are mostly prepared by electrolysis of their fused chlorides.

(c) Sodium ions are primarily found in the Blood plasma and the interstitial fluids around the cells, whereas Potassium ions are found within the cell fluids. Sodium ions help in the transmission of nerve signals and also regulate the flow of water and transport sugars and amino acids into the cells.

Thus, Sodium is more important for our survival than potassium.

Q 10.28

Write balanced equations for reactions between:

(a) Na_2O_2 and water

(b) KO_2 and Water

(c) Na_2O and CO_2

Ans 10.28

(a) $Na_2O_2 + 2H_2O \rightarrow 2NaOH + H_2O_2$

(b) $2KO_{2(s)} + 2H_2O_{(l)}$

$2KOH_{(aq)} + H_2O_{2(aq)} + O_{2(aq)}$

(c) $Na_2O_{(s)} + CO_{2(g)}$

Na_2CO_3

Q 10.29

How would you explain the following observations?

(i) BeO is almost insoluble, but $BeSO_4$ is soluble in water.

(ii) BaO is soluble, but $BaSO_4$ is insoluble in water.

(iii) LiI is more soluble than KI in ethanol.

Ans 10.29

(i) The sizes of Be^{2+} and O^{2-} are small and highly compatible with each other. Due to this, a high amount of lattice energy is released during its formation. The hydration energy, when it is made to dissolve in water, is not enough to overcome the lattice energy. Thus, BeO is almost insoluble in water.

Whereas the size of an SO_4^{2-} is large compared to Be^{2+} and there is lesser compatibility and lattice energy which can be easily overcome by the hydration energy. Thus, $BeSO_4$ is easily soluble in water.

(ii) The sizes of Ba^{2+} and SO_4^{2-} are large and highly compatible with each other. Due to this, a high amount of lattice energy is released during its formation. The hydration energy, when it is made to dissolve in water, is not enough to overcome the lattice energy. Thus, $BaSO_4$ is insoluble in water.

Whereas the size of an O^{2-} is small compared to Ba^{2+} and there is lesser compatibility and lattice energy which can be easily overcome by the hydration energy. Thus, BaO is easily soluble in water.

(iii) The lithium-ion has a smaller size, and as a result of that, it has a higher polarising capability. This enables it to polarize the electron cloud around an iodide ion, thus resulting in a greater covalent character in LiI than KI. Thus, LiI is easily soluble in ethanol.

Q 10.30

Which of the following alkali metals has the least melting point?

(a) Na (b) K (c) Rb (d) Cs

Ans 10.30

(d) Cs

Cs has the least melting point of the given alkali metals since it has the largest size. Due to its larger size, the binding capability of Cs is limited, and the lattice energy released during the formation of its compounds is less and can be easily broken.

Q 10.31

Which one of the following alkali metals gives hydrated salts?

(a) Li (b) Na (c) K (d) Cs

Ans 10.31

(a) Li

Li is capable of forming hydrated salts because of its size. Since it is smaller in size, it has a higher charge density and can easily attract water molecules around it and for hydrated salts like $LiCl.2H_2O$. The other alkali metals have a bigger size and lesser charge density and, thus, aren't capable of forming hydrated salts.

Q 10.32

Which one of the alkaline earth metal carbonates is thermally the most stable?

(a) $MgCO_3$ (b) $CaCO_3$ (c) $SrCO_3$ (d) $BaCO_3$

Ans 10.32

(d) $BaCO_3$

Thermal stability is directly proportional to the size of the cation, i.e., the larger the size of the atom, the greater its thermal stability. The biggest cation among the given compounds is Ba.

Thus, $BaCO_3$ will be the most thermal carbonate among the given compounds, followed by $SrCO_3$, $CaCO_3$ and $MgCO_3$.

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