s-Block Elements & Compounds

PROPERTIES OF s-BLOCK ELEMENTS Section - 1

Group - I

Introduction :

All the alkali metals have loosely held one s-electron in the outermost shell which they can readily lose to give monovalent (M^+) cation having stable noble gas configuration. Due to their tendency of loosing s-electron easily, they have low ionization energy and high metallic character. The size of atoms and ions of alkali metals increases down the group.

Physical Properties :

(i) Electropositive Character

Alkali metals are highly electropositive in nature and electropositive character increases down the group.

(ii) Ionization Energy

Alkali metals have low ionisation energy and it decreases down the group. In fact K and Cs are used as cathodes in photoelectric cells.

(iii) Density

The density of alkali metals is quite low as compared to other metals. Li, Na and K are even lighter than water. As we go down the group, the mass and volume of alkali metals increases but mass increases by larger factor than the volume and the resultant effect is that the density increases down the group with an exception that potassium being lighter than Sodium. So, the trend is :

$$Li < K < Na < Rb < Cs$$

(iv) Melting Point and Boiling Point :

Melting point and boiling point of alkali metals decreases down the group. The melting points range from lithium 181 °C to caesium 28.5°C. These are extremely low values for metals, and contrast with the melting points of the transition metals, most of which are above 1000 °C.

(v) Flame Test :

Group I elements give a varied range of colours in their flame tests. Li emits crimson light, Na emits yellow, K emits lilac and Rb, Cs emit violet light.

Chemical Properties :

The alkali metals are highly reactive due to low ionization enthalpy and therefore they never occur in free state. The reactivity of these metals increases down the group.

(i) Reactivity towards air

Li reacts with air forming oxide Li_2O (and some Li_2O_2) and nitride Li_3N. Na reacts with air forming oxide (Na_2O) and peroxide (Na_2O_2). Peroxide is formed in large amount. K, Cs and Rb forms oxide (M_2O), peroxide (M_2O_2) and superoxide (MO_2). Superoxide is formed in large amount. (Where M is K, Cs, Rb).

You can note here that except Li all other alkali metals are forming oxides only, whereas Li form nitride also on burning in air. Li_3N is a ruby red salt which gives LiOH and $NH_3\uparrow$ on dissolving in water while Li and nitrogen on simply heating. The increasing stability of peroxide or super-oxide, as size of the metal ion increases, is due to the stabilization of large anions by larger cations through higher lattice energies.

(ii) Reactivity towards water

Group 1 metals all react vigorously with water liberating hydrogen. The reaction becomes increasingly violent on descending the group.

$$M_{(s)} \longrightarrow M^+_{(aq)} + 1e^- \quad ; \qquad H_2O + 1e^- \longrightarrow OH^- + \frac{1}{2}H_2$$

$$M + H_2O \longrightarrow M^+ + OH^- + \frac{1}{2}H_2 \qquad (M = Li, Na, K, Rb, Cs)$$

Reaction of sodium is so violent that it catches fire and is thus kept in kerosene in the laboratory.

Standard Reduction Potential

Standard Reduction potential of alkali metals M is a value that represents the tendency to gain an electron

$$M^+_{(aq)} + 1e^- \longrightarrow M_{(s)}$$

If standard Reduction potential is positive it means element desires to gain electron and if it is negative it means element desires to loose electron.

Further, lower the standard reduction potential, higher will be the tendency to loose electron.

Standard reduction potential (E°) for an alkali metal represents the overall change :

$$M_{(s)} \longrightarrow M_{(g)} \qquad \text{Sublimation enthalpy}$$

$$M_{(g)} \longrightarrow M^+_{(g)} + 1e^- \qquad \text{Ionization enthalpy}$$

$$M^+_{(g)} + H_2O \longrightarrow M^+_{(aq)} \qquad \text{hydration enthalpy}$$

Lithium has the most negative standard reduction potential than any other element in the periodic table. Why ?

$$Li_{(s)} \xrightarrow{\Delta_a H_{Li} = 159\,KJ/mol} Li_{(g)} \xrightarrow{\Delta_{ion} H_{Li} = 520.2\,kJ/mol} Li^+_{(g)} + 1e^- \xrightarrow{\Delta_{hyd} H_{Li} = -520} Li^+_{(aq)} + 1e^-$$

Total energy absorbed $= \Delta a H_{Li} + \Delta_{ion} H_{Li} + \Delta_{hyd} H_{Li} = 159 + 520.2 - 520 = 159.2$ kJ/mol.

$$Na_{(s)} \xrightarrow{\Delta_a H_{Na} = 107\,KJ/mol} Na_{(g)} \xrightarrow{\Delta_{ion} H_{Na} = 495.8\,KJ/mol} Na^+_{(g)} + 1e^- \xrightarrow{\Delta_{hyd} H_{Na^+} = -406\,K/mol} Na^+_{(aq)} + 1e^-$$

Total energy absorbed $= \Delta_a H_{Na} + \Delta_{ion} H_{Na} + \Delta_{hyd} H_{Na^+} = 107 + 495.8 - 406 = 196.8$ kJ/mol

Hence total energy absorbed in $Li_{(s)} \longrightarrow Li^+_{(aq)} + 1e^-$ is less and also least as compared to any other element. Therefore it has most negative standard reduction potential.

Li has most negative standard reduction potential (or highest tendency for $Li_{(s)} \longrightarrow Li^+_{(aq)} + 1e^-$). It seems that reaction of Li with water should be most vigorous. But it is surprising that Li reacts less vigorously with water than other alkali metals. The explanation lies in the kinetics (that is the rate at which the reaction proceeds) rather than in the thermodynamics (that is total amount of energy absorbed). You will study more about kinetics in upcoming modules.

(iii) Reducing Character

Among alkali metals, Lithium has strongest reducing character and sodium has least reducing character and rest are almost the same.

[Lower the standard reduction potential (considering the sign also) higher is the reducing character]

(iv) Solution in liquid Ammonia

Liquid Ammonia is also a good polar solvent next to water. Ammonia gas (b. p. = – 33°C) is condensed to give liquid ammonia. Both water and ammonia undergo self - ionization :

$$2H_2O \rightleftharpoons H_3O^+ + OH^- \quad ; \quad 2NH_3 \rightleftharpoons NH_4^+ + NH_2^-$$

When small amount of sodium is added in liquid ammonia (i.e. : Dilute Solutions of Na in liquid Ammonia), sodium metal looses an electron to give sodium cation, both of these ions are solvated by liquid ammonia to give a blue coloured solution which is highly conducting.

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

The blue colour of the solution is due to the 'ammoniated electron' which absorbs energy in the visible region of light and thus imparts blue colour to the solution. The conducting nature is also mainly due to solvated electron.

The solution is paramagnetic in nature and on standing slowly liberates hydrogen resulting in the formation of amide.

$$e^- + NH_{3(\ell)} \longrightarrow NH^-_{2(ammoniated)} + \frac{1}{2}H_{2(g)}$$

Concentrated solutions of Na in liquid ammonia is metallic bronze in colour and diamagnetic in nature due to formation of metal ion clusters.

Group - II

Introduction :

All the alkaline earth metals have two s-electron in the outermost shell which if they lose, they will give divalent (M^{2+}) cation having stable noble gas configuration. The size of atoms and ions of alkaline earth metals increases down the group.

Physical Properties :

(i) Electropositive character

Alkaline earth metals are highly electropositive in nature and Electropositive character increase down the group.

(ii) Ionization Enthalpy

Second ionization enthalpy is very high than first ionization enthalpy. It is due to the fact that extracting an electron from a positive ion becomes difficult. Both first and second ionization energy decreases down the group.

(iii) Density

The density of alkaline earth metals is high as compared to alkali metals. And as we go down the group both the mass and volume of alkali metals increases but the resultant effect that is the density do not show a regular change. As we move down the group, it first decreases (upto Ca) and then increases. So, the trend is : $Ba > Sr > Be > Mg > Ca$

(iv) Melting Point and Boiling Point

Melting point and boiling point of alkaline earth metals is higher than the alkali metals. As we move down the group, Melting point and Boiling point don't show a regular change.

Chemical properties :

(i) Reactivity towards air

Alkaline earth metals react with air forming an oxide and a nitride.

$$M_{(s)} \xrightarrow{\text{air}} MO + M_3N_2 \quad (\text{where } M = Be, Mg, Ca, Sr, Ba)$$

The case of above reaction depends on electropositive character and hence increases down the group. In the case of Mg, it burns with a dazzling white light which is used to provide light in flash photography using light bulks.

(ii) Reactivity towards water

$$\text{OHC}: M_{(s)} \longrightarrow M^{2+}_{(aq)} + 2e^- \qquad \text{RHC}: 2H_2O_{(\ell)} + 2e^- \longrightarrow H_2 + 2OH^-$$

Now, to check the ease of reaction for different metals, we have to check the Oxidation Half Cell reaction and this reaction refers to the reducing ability (Standard Reduction Potential) of the metal.

And, Since for alkaline earth metals reducing character increases down the group, the ease of reaction of alkaline earth metals with water increases down the group.

Experimentally, it is seen that Be reacts only with steam and Magnesium can react with both hot water and steam and Ca, Sr, Ba react even with cold water rapidly.

(iii) Solutions in liquid ammonia

In liquid ammonia, group II metals form bright blue dilute solutions containing solvated electrons and metal hexa ammoniates. The metal hexammoniates form ammides on heating which further form nitrides and NH_3 concentrated solutions are bronze coloured.

(iv) Reducing character

Less is the standard reduction potential more will be reducing character and since the standard reduction potential decreases down the group, the reducing character of alkaline earth metals increases down the group.

IN-CHAPTER EXERCISE - A

Choose the correct alternative. Only One choice is correct. However, question marked with '*' may have more than one correct option.

***1.** *Highly pure dilute solution of sodium in liquid ammonia.*

(A) *Shows blue colour* (B) *Exhibits electrical conductivity*

(C) *Produces sodium amide* (D) *Produces H_2 gas.*

2. *Sodium emits bright D-lines in _________ region of spectrum* :

(A) *Red* (B) *Indigo* (C) *yellow* (D) *Green*

***3.** *The compound (s) formed upon combustion of Na in excess air is (are)* :

(A) Na_2O_2 (B) Na_2O (C) NaO_2 (D) $NaOH$

4. *Which of the following properties of Group I elements decreases on going down the group.*

(A) *Cohesive energy* (B) *Reactivity with water*

(C) *Standard reduction potential* (D) *Reducing character*

(E) *Density* (F) *Melting and boiling points.*

5. *Among the given statements the incorrect one is* :

(A) *Be differs much more from other alkali earth metals than Li does from others alkali metals*

(B) *Be forms a very strong complex* $[Be(H_2O)_4]^{2+}$

(C) *Be generally forms covalent compounds*

(D) *Be usually has more than four water of crystallisation associated with it*

IMPORTANT CONCEPTS Section - 2

In this section we will discuss some important concepts related to s-block elements.

Lattice enthalpy

It is the energy required to separate one mole of a solid ionic compound completely in gaseous ions.

Lattice Enthalpy of breaking solid NaCl in gaseous Na^+ and Cl^- is +788 KJ/mol.

In a system of +ve and –ve ion the energy required to separate the two ions is inversely proportion to the distance between the centres of the two ions.

$$\Rightarrow \quad \text{LE} \propto \frac{1}{r^+ + r^-}$$

(r^+ is the radius of +ve ion and r^- is the radius of –ve ion.)

For a given anion and different cations as we move down the group the size of cation increase and consequently the interionic distance between cation and anion increases. Hence, the lattice enthalpy would decrease down the group.

Also you may note that as the charge of ion increases the Lattice Enthalpy increases. Therefore, Lattice Enthalpy of group II metal salts is more as compared to corresponding group I metal salts.