NEET Revision Notes
Chemistry
The S-Block Elements

Introduction

The s-Block elements of the modern periodic table are those elements in which the last electron enters the s-orbital. Because s-orbital has an occupancy of only two electrons, there are only two groups: group-1 and group-2 that belong to the s-Block of the modern periodic table.

Group-1 elements consist of six metals which are collectively called alkali metals; as, on reaction with water, they form hydroxide which is strongly alkaline in nature.

Group-1 elements are Lithium, Sodium, Potassium, Rubidium, Caesium, Francium.

Group-2 elements also include six metals and they are commonly called alkaline earth metals (except Beryllium) because their oxides and hydroxides are alkaline in nature and their metal oxides are found in earth crust.

Group-2 elements are Beryllium, Magnesium, Calcium, Strontium, Barium, Radium.

Group-1 elements

Electronic configuration: All the group-1 elements have one valence electron; ns^1. This single electron in the relation to these elements make them highly electropositive metals.

They readily lose the electron to have a monovalent M^+ ion. Therefore, they never exist in the free State.

Electronic configuration of alkali metals is shown below:

$$Li = 1s^2 2s^1$$

$$Na = 1s^2 2s^2 2p^6 3s^1$$

$$K = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$$

$$Rb = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$$

$$Cs = [Xe] 6s^1$$

$$Fr = [Rn] 7s^1$$

Atomic radii:

Out of all the elements, alkali metals have the largest size in their particular period. As we move down the group, the new subshells are added to the atomic structure of the element hence the atomic radius increases.
If we compare the size of the monovalent ion and the parent atom, the monovalent ion is smaller than the parent atom $M^{+} < M$ because after the elimination of the outermost electron the nuclear charge of attraction over the valence electrons increases and it tends to hold the remaining electrons more tightly and as a result the size of ion decreases.

Ionisation enthalpy:

- The ionisation therapy of group-1 elements is relatively low and it decreases down the group from Lithium to Caesium.
- The reason behind this trend is the effect of increasing size as an increase in atomic radii outweighs the increasing nuclear charge, and the valence electron is well screened from the nuclear charge.
- So it becomes easy to eliminate the electron from the outermost shell and less amount of energy is necessary.

Hydration enthalpy:

- The hydration enthalpy of elements depends upon its charge density.
- As we move down the group from Lithium to Caesium, the charge density of elements decreases which results in a decrease in hydration enthalpy.
- Therefore, lithium-ion has maximum degree of hydration and as a result lithium salts are mostly hydrated.

Physical properties:

- All the group-1 elements are soft and light metals.
- Having the large size these elements have low density which increases down the group but there’s an exception that potassium is lighter than sodium.
- Alkali metals have low boiling and melting points because of their weak metal-metal bonding.
- The alkali metals and their salts give characteristic colour to an oxidising flame.

- When the alkali metals are heated, the heat from the flame excites the outermost electron to a higher energy level and when the excited electron comes back to its ground state there is an emission of radiation in the visible region of the spectrum.

Chemical properties:

1. Reactivity towards air:

Alkali metals tarnish in dry air because of the formation of their oxides which in turn react with the moisture of air to form hydroxide. In the presence of oxygen alkali metals burn vigorously and form their respective oxide.

Out of six elements of group-1, Lithium forms monoxide while sodium forms peroxide and other metals form superoxide. Superoxide ions are only stable in the presence of large cations such as potassium rubidium and caesium. The reaction for the formation of oxides of alkali metals are shown below:

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

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

$$M + O_2 \longrightarrow MO_2\,(\text{superoxide})$$

$$M = K, Rb, Cs$$

As the alkali metals are highly reactive towards air and water they are normally kept in kerosene oil.

2. Reactivity towards water:

When the alkali metals react with water, hydroxide is formed along with the evolution of hydrogen gas. The reaction is as follows:

$$2M + 2H_2O \longrightarrow 2M^+ + 2OH^- + H_2$$

$$M = \text{alkali metal}$$

3. Reactivity towards hydrogen:

When alkali metals react with hydrogen metal hydrides are formed which are ionic in nature and have high melting points.

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

4. Reactivity towards halogen:

Alkali metals readily react with halogens to form ionic halides M^+X^-. Exception is that lithium halide has covalent nature as lithium has high polarisation capability and has a tendency to distort electron clouds around the negative halide ion. Therefore, the anion with large size can easily be distorted, lithium halides are covalent in nature.

Anomalous properties of lithium:

- The first element of group-1; lithium shows anomalous behaviour in certain properties because of its exceptionally small size of both parent atom and monovalent ion and its high polarising power i.e. high charge by radius ratio.
- As a result, lithium compounds have increased covalent character due to which its solubility in organic solvents is also high.
- Lithium has maximum hydration enthalpy and its salts are often found in hydrated form.
- Among all the group-1 elements lithium is least reactive but the strongest oxidising agent.

Diagonal relationship of lithium to magnesium:

- Lithium and magnesium have similar properties because of their similar sizes

  $Li = 152\,pm$ $\quad\quad Mg = 160\,pm$

  $Li^{+} = 76\,pm$ $\quad\quad Mg^{+2} = 72\,pm$
- Both the elements react slowly with water and their oxides and hydroxide are much soluble than the oxides and hydroxides of other elements of their particular group.
- Both lithium and magnesium nitride compounds by direct combination with nitrogen. Li_3N and Mg_3N_2.
- The oxides of lithium and magnesium Li_2O and MgO do not combine with excess oxygen to give any superoxide. The halide pound of lithium and magnesium, LiCl and $MgCl_2$ are soluble in ethanol.

Some important compounds of sodium:
Sodium carbonate:

- It is also called as washing soda; $Na_2CO_3.10H_2O$
- The process by which sodium carbonate is prepared is called Solvay process where concentrated solution of NaCl is saturated with ammonia.

- Firstly, ammonia reacts with water and carbon dioxide gas to form ammonium carbonate followed by formation of ammonium hydrogen carbonate.
- Ammonium hydrogen carbonate is subjected to react with NaCl which results in the formation of sodium hydrogen carbonate along with ammonium chloride as by-product.
- On heating the crystals of Sodium hydrogen carbonate, sodium carbonate is formed.

The reaction for the process can be written as follows:

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

$$(NH_4)_2CO_3 + H_2O + CO_2 \longrightarrow 2NH_4HCO_3$$

$$NH_4HCO_3 + NaCl \longrightarrow NH_4Cl + NaHCO_3$$

$$2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$$

Properties:

- Sodium carbonate exists in a white crystalline solid form and is a decahydrate. $Na_2CO_3.10H_2O$
- Sodium carbonate is also called washing soda and has high water solubility. On heating sodium carbonate, it loses its water of crystallisation and forms monohydrate from decahydrate.

$$Na_2CO_3.10H_2O \xrightarrow{375K} Na_2CO_3.H_2O + 9H_2O$$

$$Na_2CO_3.H_2O \xrightarrow{>373K} Na_2CO_3 + H_2O$$

- Heating at higher temperature, above 373K, monohydrate changes completely to anhydrous condition and a white powder-like compound is formed called soda ash.

Uses:

- Washing soda is used in water softening, laundering and cleaning purposes.
- It is widely used in the manufacture of glass, borax, soap and caustic soda.
- Washing soda is also used in paper, textile and paint industries.
- Along with all this sodium carbonate is an important Lab reagent for both qualitative and quantitative analysis of elements.

Sodium chloride:

- Sodium chloride, commonly known as common salt, is obtained by the evaporation of seawater.

- The raw NaCl is prepared by crystallisation of brine solution, consisting of impurities such as sodium sulphate, calcium sulphate, calcium chloride and magnesium chloride.
- To obtain pure sodium chloride, the crude salt is dissolved in a minimum amount of water and filtered several times to remove insoluble impurities.
- Then the solution is saturated with hydro chloride gas and as a result crystals of pure sodium chloride or obtained while impurities of calcium and magnesium chloride are soluble in water so they remain in solution.

Properties:

Sodium chloride is a water soluble compound and its solubility does not increase with increase in temperature. NaCl has a melting point of 1081K.

Uses:

NaCl is widely used for domestic purpose while it is commercially used for the preparation of sodium carbonate, sodium hydroxide and Na_2O_2.

Sodium hydroxide:

- Sodium hydroxide is also called as caustic soda. NaOH is commercially prepared by the process of electrolysis of NaCl in Castner-Kellner cell.
- In the process, the brine solution is electrolysed using mercury as cathode and carbon as anode.
- Sodium metal is discharged at the cathode combines with mercury to form sodium amalgam while chlorine gas is evolved at a note.

The reaction for the above process can be written as follows:

$$Cathode: Na^+ + e^- \xrightarrow{Hg} Na - \text{amalgam}$$

$$Anode: Cl^- \longrightarrow \frac{1}{2}Cl_2 + e^-$$

The sodium amalgam formed is then treated with water to form sodium hydroxide and hydrogen gas is evolved in the reaction.

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

Properties:

- Sodium hydroxide is a water soluble compound that gives a strong alkaline solution.
- It is hygroscopic in nature and reacts with CO_2 in the atmosphere to form sodium carbonate.
- Sodium hydroxide appears as a white translucent solid which melts at 591K.

Uses:

- Sodium hydroxide is widely used in the manufacture of soap, paper, artificial silk and a number of other chemicals.
- It is useful in petroleum refining and purification of bauxite ore.
- Sodium hydroxide is also utilised in textile industries and for the preparation of pure fats and oils in the food industry.
- It is an important laboratory reagent.

Sodium hydrogen carbonate; baking soda

- Baking soda is synthesised from sodium carbonate solution saturated with carbon dioxide gas.
- This reaction results in the formation of white crystalline powder which is less soluble and easily separates out.

The reaction for the above can be written as follows:

$$Na_2CO_3 + H_2O + CO_2 \longrightarrow 2NaHCO_3$$

Properties and uses:

- Sodium hydrogen carbonate is a mild antiseptic for treating skin infections and it is also used in fire extinguishers.
- Sodium hydrogen carbonate decomposes on heating and generates bubbles of carbon dioxide.
- To this property it is widely used in making cakes and pastries.

Biological importance of sodium and potassium:

Sodium: sodium ions are a major component of our living system, they are found mainly on the outside of cells in the interstitial fluid.

- These irons play a crucial role in the transmission of nerve signals, regulating the flow of water across cell membranes and transport of compounds like sugar and amino acids into the cell.
- Potassium ions are the abundant cations present in cell fluids. T
- They act as an activator for many enzymes and participate in the process of ATP production.
- Sodium and potassium both the ions are the key elements of the sodium-potassium pump that operates across the cell membrane.

Group-2 elements
Electronic configuration:

Alkaline earth metals have two electrons in their s-orbital valence shell; they represent the general electronic configuration of valence shells as ns^2.
Electronic configuration of group two elements are as follows:

$Be = 1s^2 2s^2$

$Mg = 1s^2 2s^2 2p^6 3s^2$

$Ca = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

$Sr = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2$

$Ba = [Xe] 6s^2$

$Ra = [Rn] 7s^2$

Atomic radii and ionic radii:

As compared to group-1 elements the elements of group-2 are smaller in size. This is because of the increased nuclear charge in these elements. On moving down the group the size of elements increases with the increase in atomic number.

Ionisation enthalpy:

- The alkaline earth metals have low ionisation enthalpy due to their large size of atoms. As the atomic size increases down the group, ionisation enthalpy decreases.
- If we compare the first ionisation enthalpies of group-1 and group-2 elements of corresponding period, the alkaline earth metal will have higher ionisation enthalpy than the metal of corresponding group-1.
- This is because alkaline earth metal has small size than alkali metal and two electrons or completely filled in s-orbital of valence shell so it is comparatively difficult to eliminate electrons from the fully filled s-orbital than the single electron from the s-orbital of alkali metal.
- The second ionisation enthalpies of group-2 metals are smaller than that of corresponding group one metal.

Hydration enthalpy:

- Like group-1 elements, the hydration enthalpy of alkaline earth metal ions also decreases with the increase in ionic size.
- Therefore, beryllium ion will have maximum hydration enthalpy.

- In comparison to group-1 elements, the hydration enthalpy of alkaline earth metal ions is larger than those of alkali metal ions.
- As a result, the halides of alkaline earth metals, magnesium chloride and calcium chloride are more extensively hydrated than the compounds of alkali metals. Exception: NaCl and KCl do not form any such hydrates.

Physical properties:

- Group-2 elements are silvery, lustrous, relatively soft but harder in nature than group-1 elements.
- The melting point and boiling point of these metals are relatively higher than group-1 elements due to their smaller size but this trend is not systematic.
- As they have low ionisation enthalpy, they are electropositive metals and their electropositive character increases down the group from Be to Ba.
- Group-2 elements also impart characteristic colour respective to the flame as the electrons are excited to higher energy level and when they draw back to the ground state energy is emitted in the visible region of spectrum.
- Out of all elements of group-2, beryllium and magnesium do not impart any colour to flame because electrons are so tightly bound that they do not get excited and no characteristic flame is observed.
- These elements have higher electrical and thermal conductivity.

Chemical properties:

The reactivity of group-2 metals increases down the group but they are less reactive than the corresponding group-1 metals.

Reactivity towards air:

- Beryllium and magnesium are inert towards oxygen and water because they form an oxide film on their surface.
- The powdered beryllium burns vigorously on ignition in air to form beryllium oxide and beryllium nitride.
- Similarly, Magnesium is highly electropositive metal and it burns with dazzling flame in air to form magnesium oxide and magnesium nitride.
- However, calcium, strontium and barium readily reacts with air to form their corresponding oxides and nitrides.

Reactivity towards halogens:

The alkaline earth metals combine with halogen only at high-temperature to form their metal halide compounds.
The reaction for formation of metal Halide is as follows:

$$M + X_2 \longrightarrow MX_2$$
$$(M = f, Cl, Br, I)$$

Reactivity towards dihydrogen:

- Out of all the group-2 elements, beryllium does not combine with hydrogen while others react with hydrogen to form their corresponding hydrides MH_2.
- Beryllium hydride can be prepared by the reduction process of $BeCl_2$ with $LiAlH_4$. The reaction is as follows:

  $$2BeCl_2 + LiAlH_4 \longrightarrow 2BeH_2 + LiCl + AlCl_3$$

Reactivity towards acids:
Group-2 metals readily react with acid and liberate hydrogen gas. Reaction is as follows:

$$M + 2HCl \longrightarrow MCl_2 + H_2$$

Anomalous behaviour of beryllium:

- Beryllium shows anomalous behaviour as it has exceptionally small size as compared to other members of the group due to which it has high and hydration enthalpy and forms largely covalent compounds which can easily be hydrolysed.
- Beryllium is unable to exhibit coordination number more than four as it lacks d-orbital while the other group elements can exceed their coordination number to six by making use of their d-orbitals.
- The oxides and hydroxides of beryllium are amphoteric in nature while the oxides and hydroxides of other elements are not.

Diagonal relationship between beryllium and aluminium:

- Both beryllium ion and aluminium ion have similar charge by radius ratio hence they both resemble in certain ways.

- Both of them were not readily react with acid because of the oxide film present on their surface.
- Both the elements have strong tendency to form complexes with fluorine.
- The chloride compound of beryllium and aluminium have structure in such a manner that chlorine is bridged in between in the vapour phase.

Some important compounds of calcium:
Calcium oxide, lime

Calcium oxide is also called quicklime and is commercially prepared by heating calcium carbonate in a Rotary kiln at a temperature of $1070-1270K$.
Reaction for the same is shown as below:

$$CaCO_3 \overset{heat}{\leftrightarrows} CaO + CO_2$$

Quicklime is a white amorphous solid having melting point of 2870K. On exposure to atmosphere it readily absorbs moisture and CO_2.

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

$$CaO + CO_2 \longrightarrow CaCO_3$$

Properties and uses:

- Calcium oxide is an important material in the manufacturing of cement and is the cheapest form of alkali.
- It is also utilised in manufacturing of sodium carbonate from caustic soda.
- Commercially it is utilised in the purification process of sugar and in manufacturing of dye and other synthetic chemicals.

Calcium carbonate, lime stone

Limestone can be prepared by addition of sodium carbonate to calcium chloride or by passing carbon dioxide gas through slaked lime. The reaction for both the processes is shown below:

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

$$CaCl_2 + Na_2CO_3 \longrightarrow CaCO_3 + 2NaCl$$

Properties:
Calcium carbonate is a fluffy powder and is insoluble in water. On heating above 1200K it gets decomposed into calcium oxide and carbon dioxide is evolved in the reaction.

$$CaCO_3 \xrightarrow{1200K} CaO + CO_2 .$$

Uses:

- Calcium carbonate is used as a building material in the form of marble.
- Calcium carbonate along with magnesium carbonate can be utilised as a flux in the extraction process of iron metal.
- It can also be used as an antacid, filler chemical in cosmetics and is a constituent of chewing gum.

Biological importance of magnesium and calcium:

Magnesium is required as the core factor by enzymes that utilise ATP in phosphate transfer. Magnesium is also constituted in the chlorophyll pigment of plants.

Calcium is the mean constituent of bones and teeth present in human body. It plays a crucial role in neuro muscular function, inter-neuronal transmission, blood coagulation and cell membrane integrity.