

Subject: Chemistry
Class: XI
Chapter: The s-Block Elements
Top concepts

1. The s-block elements of the periodic table are those in which the last electron enters the outermost s-orbital
2. Elements of group 1 are called alkali metals and elements of group 2 are called alkaline earth metal
3. Group 1 elements: Alkali metals
 - (i) Group 1 elements are called alkali metals because they form hydroxides on reaction with water which are strongly alkaline in nature
 - (ii) The general electronic for alkali metals is [noble gas] ns^1

Electronic configuration of alkali metals:

Element	Symbol	Atomic Number	Electronic configuration
Lithium	Li	3	[He] $2s^1$
Sodium	Na	11	[Ne] $3s^1$
Potassium	K	19	[Ar] $4s^1$
Rubidium	Rb	37	[Kr] $5s^1$
Caesium	Cs	55	[Xe] $6s^1$
Francium	Fr	87	[Rn] $7s^1$

(iii) Trends in properties of alkali metals:

1. Atomic and ionic radii: Alkali metals have the largest atomic and ionic radii in their respective periods of the periodic table. On moving down the group, the atomic and ionic radii increase

Explanation: As we move in a period, the atomic radius and ionic radius tend to decrease due to increase in the effective nuclear charge. Therefore alkali metals have largest atomic and ionic radii in their respective group. On moving down the group there is increase in the number of shells. Thus there is an increase in distance between nucleus and outermost electrons which eventually increases atomic and ionic radii

2. Ionisation enthalpies: Alkali metals have the lowest ionization enthalpy in each period. Within the group, the ionization enthalpies of alkali metals decrease down the group

Explanation: Atoms of alkali metals are largest in their respective periods and therefore, the valence electrons are loosely held by the nucleus. By losing the valence electron, they acquire stable noble gas configuration. This accounts for their ease to lose electrons and hence they, have low ionization enthalpies

3) The second ionization enthalpies of alkali metals are very high.

Explanation: When an electron is removed from the alkali metals, they form monovalent cations which have very stable electronic configurations (same as that of noble gases). Therefore it becomes difficult to remove the second electron from the stable noble gas configurations and hence their second ionization enthalpy values (IE_2) are very high

4. Melting and boiling points: Alkali metals are soft and have low melting and boiling points

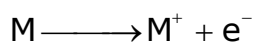
Explanation: Alkali metals have only one valence electron per metal atom and therefore, the energy binding the atoms in the crystal lattice of the metal is low. Thus, the metallic bonds in these metals are not very strong and consequently, their melting and boiling points decrease on moving down from Li to Cs

5. Density: Densities of alkali metals are quite low as compared to other metals. The densities increase on moving down the group. But K is lighter than Na

Explanation: The densities of metallic elements depend upon the type of packing of atoms in metallic state and also on their size. The alkali metals have close packing of metal atoms in their lattice and because of the large size of their atoms, they have low densities. On moving down the group from Li to Cs, there is increase in atomic size as well as atomic mass. But the increase in atomic mass is more and compensates the increase in atomic size. As a result, the densities (mass/volume) of alkali metals gradually increase from Li to Cs. K is lighter than Na due to increase in atomic size of K

6. Electropositive or metallic character: All the alkali metals are strongly electropositive or metallic in character

Explanation: The electropositive character of an element is expressed in terms of the tendency of its atom to release electrons:



As alkali metals have low ionization enthalpies, their atoms readily lose their valence electron. These elements are, therefore, said to have strong electropositive or metallic character. Since, the ionization energies decrease down the group, the electron releasing tendency or electropositive character is expected to increase down the group

7. Oxidation states: All the alkali metals predominantly exhibit an oxidation state of +1 in their compounds

Explanation: Alkali metals have only one electron in their valence shell and therefore can lose the single valence electron readily to acquire the stable configuration of a noble gas. Since the second ionization energies are very high, they cannot form divalent ions. Thus, alkali metals are univalent and form ionic compounds

8. Characteristic flame colouration: All the alkali metals and their salts impart characteristic flame colouration

Explanation: Alkali metals have very low ionization enthalpies. The energy from the flame of Bunsen burner is sufficient to excite the electrons of alkali metals to higher energy levels. Excited state is quite unstable and therefore when these excited electrons come back to their original energy levels, they emit extra energy, which falls in the visible region of the electromagnetic spectrum and thus appear coloured. Characteristic flame colouration by different alkali metals can be explained on the basis of difference in amount of energy absorbed for excitation of the valence electron

9. Photoelectric effect: Phenomenon of ejection of electrons when electromagnetic radiation of suitable frequency strikes metal surface is called photoelectric effect. Alkali metals exhibit photoelectric effect

Explanation: Alkali metals have low ionization enthalpies and therefore, the electrons are easily ejected when exposed to light. Among alkali metals, caesium has lowest ionization enthalpy and hence it shows photoelectric effect to the maximum extent

10. Hydration of alkali metal ions: Alkali metal ions are highly hydrated.

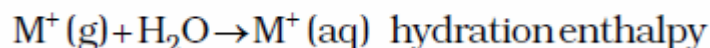
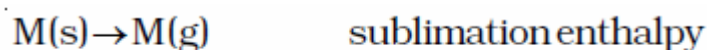
Explanation: Smaller the size of the ion, the greater is the degree of hydration. Thus, Li^+ ion gets much more hydrated than Na^+ ion which is more hydrated than K^+ ion and so on. Therefore, the extent of hydration decreases from Li^+ to Cs^+ . As a result of larger hydration of Li^+ ion than Na^+ ion, the effective size of hydrated Li^+ ion is more than that of hydrated Na^+ ion.

Hydrated ionic radii decrease in the order: $\text{Li}^+ > \text{Na}^+ > \text{K}^+ > \text{Rb}^+ > \text{Cs}^+$

Due to extensive hydration, Li^+ ion has lowest mobility in water.

11. Reducing Nature: Alkali metals are strong reducing agents. This is due to their greater ease to lose electrons

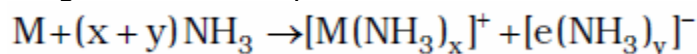
Explanation: Lithium is the strongest reducing agent. Tendency to act as reducing agent depends on energy requirement involved in three processes i.e. sublimation, ionization and hydration.



Lithium being small in size has high ionization enthalpy. On the other hand because of small size it is extensively hydrated and has very high hydration enthalpy. This high hydration enthalpy compensates the high energy needed to remove electron (in second step). Thus Li has greater tendency to lose electrons in solution than other alkali metals. Thus Li is the strongest reducing agent.

12. Solutions in liquid ammonia: Alkali metals dissolve in liquid ammonia giving deep blue solutions which are conducting in nature

Explanation: In solution the alkali metal atom readily loses the valence electron. Both the cation and the electron combine with ammonia to form ammoniated cation and ammoniated electron. 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



13. Basic strength of hydroxides: Hydroxides of alkali metals are strongly basic and basic strength increases down the group

Explanation: The Metal—OH bond in the hydroxides of alkali metals is very weak and it can easily ionize to form M^+ and OH^- ions. This accounts for their basic character. Since the ionization energy decreases down the group, the bond between metal and oxygen becomes weak. Therefore, the basic strength of the hydroxides increases accordingly.

(iv) Important Compounds of alkali metals:

Name of compound	Name of process & Brief about the process	Related chemical equations
Sodium Carbonate (Washing Soda), $Na_2CO_3 \cdot 10H_2O$	Solvay Process: When CO_2 gas is passed through a brine solution saturated with ammonia, sodium bicarbonate is formed. Sodium bicarbonate on heating forms sodium carbonate	$2NH_3 + H_2O + CO_2 \rightarrow (NH_4)_2CO_3$ $(NH_4)_2CO_3 + H_2O + CO_2 \rightarrow 2NH_4HCO_3$ $NH_4HCO_3 + NaCl \rightarrow NH_4Cl + NaHCO_3$ $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$
Sodium Chloride, NaCl	Evaporation of sea water: Evaporation of sea water gives crude salt which contains impurities of $CaSO_4$, Na_2SO_4 etc. To obtain pure sodium chloride,	-

	the crude salt is dissolved in minimum amount of water and filtered to remove insoluble impurities. The solution is then saturated with HCl gas. Crystals of pure sodium chloride separate out.	
Sodium Hydroxide (Caustic Soda), NaOH	Electrolysis of NaCl in Castner- Kellner cell: A brine solution is electrolysed using a mercury cathode and a carbon anode. Na metal discharged at the cathode combines with mercury to form sodium amalgam. Chlorine gas is evolved at the anode. Sodium amalgam on treatment with water forms sodium hydroxide	$\text{Cathode : Na}^+ + \text{e}^- \xrightarrow{\text{Hg}} \text{Na - amalgam}$ $\text{Anode : Cl}^- \rightarrow \frac{1}{2}\text{Cl}_2 + \text{e}^-$ $2\text{Na-amalgam} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + 2\text{Hg} + \text{H}_2$
Sodium Hydrogencarbonate (Baking Soda), NaHCO ₃	NaHCO ₃ is made by saturating a solution of sodium carbonate with carbon dioxide	$\text{Na}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2 \rightarrow 2\text{NaHCO}_3$

4. Group 2 elements: Alkaline earth metals

(i) Group 2 elements are called alkaline earth metals because their oxides and hydroxides are alkaline in nature and these metal oxides are found in the earth's crust

(ii) The general electronic configuration for alkaline earth metals is [noble gas] ns²

Electronic configuration of alkaline earth metals:

Element	Symbol	Atomic Number	Electronic configuration
Beryllium	Be	4	[He]2s ²
Magnesium	Mg	12	[Ne]3s ²
Calcium	Ca	20	[Ar]4s ²
Strontium	Sr	38	[Kr]5s ²
Barium	Ba	56	[Xe]6s ²

Radium	Ra	88	$[\text{Rn}]7s^2$
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(iii) Trends in properties of alkaline metals and their comparison with alkali metals:

1. Atomic and ionic radii: Atomic and ionic radii of alkaline earth metals increases down the group and are smaller than the corresponding members of the alkali metals

Explanation: Alkaline earth metals have a higher nuclear charge and therefore, the electrons are attracted more towards the nucleus. As a result, their atomic and ionic radii are smaller than those of alkali metals. Atomic and ionic radii increase on moving down the group, due to screening effect and increases in the number of the shells.

2. Ionisation enthalpies: Alkaline earth metals have low ionization enthalpies due to their large size

Explanation: Although IE_1 values of alkaline earth metals are higher than those of alkali metals, the IE_2 values of alkaline earth metals are much smaller than those of alkali metals

In case of alkali metals (for e.g. Na) the second electron is to be removed from a cation which has already acquired a noble gas configuration. However in the alkaline earth metals (for e.g. Mg), the second electron is to be removed from a monovalent cation, i.e. Mg^+ : $(1s^2 2s^2 p^6 3s^1)$ which still has one electron in the outermost shell. Thus, the second electron in Mg can be removed easily.

3. Melting and boiling points: Alkaline earth metals have higher melting and boiling points than the corresponding alkali metals

Explanation: Atoms of alkaline earth metals have smaller size as compared to alkali metals. Due to two valence electrons in alkaline earth metals, metallic bond is stronger than alkali metals. Hence alkaline earth metals have higher melting and boiling points than corresponding alkali metals

4. Electropositive or metallic character: The electropositive character increases down the group i.e., from Be to Ba but alkaline earth metals are not as strongly electropositive as the alkali metals

Explanation: Because of the low ionization enthalpies of alkaline earth metals, they are strongly electropositive in nature. However, these are not as strongly electropositive as the alkali metals of group 1 because of comparatively higher ionization energies

5. Oxidation states: All the alkaline earth metal exhibits an oxidation state of +2 in their compounds

Explanation: Alkaline earth metals exhibit oxidation state of +2 as they can lose two electrons and form bivalent ions. Following are the reason for the easy removal of two electrons

(i) Divalent cations have stable noble gas configuration

(ii) In solution, the +2 ions of alkaline earth metals are extensively hydrated and the high hydration energies of E^{2+} ions make them more stable than E^+ ions. It is observed that the amount of energy released when E^{2+} ion is dissolved in water is much more than that for E^+ ions. This large amount of extra energy released in the hydration of +2 ions is more and compensates the second ionization enthalpy required for the formation of such E^{2+} ions

(iii) In the solid state, the divalent cations form stronger lattices than monovalent cations and therefore, a lot of energy called lattice enthalpy is released. It is the greater lattice enthalpy of E^{2+} ion which compensates for the high second ionization enthalpy and is responsible for its greater stability as compared to E^+ ion

6. Hydration enthalpy: Hydration enthalpies of alkaline earth metal ions decrease down the group and are larger than those of the corresponding alkali metal ions

Explanation: This is due to smaller size of alkaline earth metal ions as compared to corresponding alkali metal ions of the same period. Due to smaller size, the alkaline earth metal ions are strongly attracted by polar water molecules and hence have larger hydration enthalpies. The hydration enthalpies of alkaline earth metal ions decrease with increase in ionic size down the group.



7. Basic strength of hydroxides: Hydroxides of alkaline earth metals are less basic than alkali metals of the corresponding period.

Explanation: Lesser basic strength of hydroxides of alkaline earth metal hydroxides is due to their high ionization enthalpies, small ionic size and dipositive charge on the ions

As a result, the Metal—O bond in these hydroxides is relatively stronger than that of corresponding alkali metals and therefore, does not break easily. Therefore, they are less basic than corresponding alkali metals.

(iii) Important Compounds of alkaline earth metals:

Name of compound	Name of process & Brief about the process	Related chemical equations
Calcium Oxide or Quick Lime, CaO	It is prepared by heating limestone (CaCO_3) in a rotary kiln at 1070-1270 K. The carbon dioxide is removed as soon as it is produced to enable the reaction to proceed to completion	$\text{CaCO}_3 \xrightleftharpoons{\text{heat}} \text{CaO} + \text{CO}_2$

Calcium Hydroxide (Slaked lime), Ca(OH)_2	It is prepared by adding water to quicklime, CaO	$\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$
Calcium Carbonate, CaCO_3	It is prepared by passing carbon dioxide through slaked lime. Excess of carbon dioxide should be avoided since this leads to the formation of water soluble calcium hydrogencarbonate	$\text{Ca(OH)}_2 + \text{CO}_2 \rightarrow \text{CaCO}_3 + \text{H}_2\text{O}$
	It is prepared by the addition of sodium carbonate to calcium chloride.	$\text{CaCl}_2 + \text{Na}_2\text{CO}_3 \rightarrow \text{CaCO}_3 + 2\text{NaCl}$
Calcium Sulphate (Plaster of Paris), $\text{CaSO}_4 \cdot \frac{1}{2} \text{H}_2\text{O}$	It is obtained when gypsum, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$, is heated to 393 K	$2(\text{CaSO}_4 \cdot 2\text{H}_2\text{O}) \rightarrow 2(\text{CaSO}_4) \cdot \text{H}_2\text{O} + 3\text{H}_2\text{O}$

5. Diagonal relationship: It is observed that some elements of second period show similarities with the elements of third period present diagonally to each other though belonging to different groups. This is called diagonal relationship

Cause of diagonal relationship: Reason is the similarity in properties such as electronegativity, ionization enthalpy size, charge/radius ratio, etc. between the diagonal elements. On moving from left to right across a period, for example, the electronegativity increases and while moving down group electronegativity decreases. Therefore, on moving diagonally, the two opposing tendencies almost cancel out and the electronegativity values remain almost same as we move diagonally. Thus, the diagonal pairs have many similar properties.

Following pairs exhibit diagonal similarity:

- (i) Lithium – Magnesium
- (ii) Beryllium – Aluminium
- (iii) Boron – Silicon

6. Some important reasons:

(1) Lithium carbonate decomposes at a lower temperature whereas sodium carbonate decomposes at a higher temperature

Explanation: Lithium carbonate decomposes at a lower temperature whereas sodium carbonate decomposes at a higher temperature because lithium carbonate is unstable to heat. Lithium ion is smaller in size than sodium ion and thus polarizes the carbonate ion to a greater extent than sodium ion leading to the formation of more stable lithium oxide and carbon dioxide.

(2) Potassium carbonate cannot be prepared by Solvay process

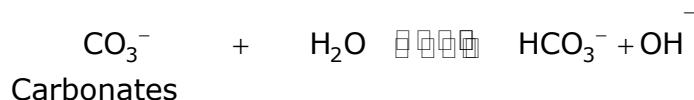
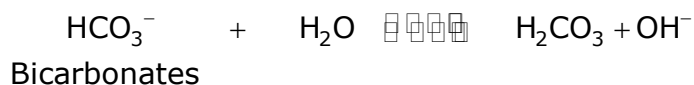
Explanation: Unlike NaHCO_3 , the intermediate KHCO_3 formed during reaction in Solvay process, is highly soluble in water and thus cannot be taken out from solution to obtain K_2CO_3 . Hence, K_2CO_3 cannot be prepared by Solvay process.

(3) Alkali and alkaline earth metals cannot be obtained by chemical reduction methods

Explanation: Alkali and alkaline earth metals are themselves very strong reducing agents and therefore cannot be reduced by chemical reduction methods

(5) An aqueous solution of carbonates and bicarbonates of alkali metals is alkaline

Explanation: The aqueous solution of carbonates and bicarbonates of alkali metals is alkaline due to the hydrolysis of these salts which produces hydroxide ions.



(5) LiCl is more covalent than KCl

Explanation: According to Fajan's rule smaller the size of cation and larger the size of anion greater is the covalent character of ionic bond. Li is small in size than K, thus Li^+ has a high charge density. Thus polarizing power of Li^+ is higher than K^+ , hence LiCl is more covalent than KCl.

(6) BaO is soluble but BaSO_4 is insoluble in water

Explanation: Size of O^{2-} ion is smaller than SO_4^{2-} . Since a bigger anions stabilizes bigger cation more than a smaller cation stabilizes a bigger anion,

lattice enthalpy of BaO is smaller than BaSO₄. BaO is soluble as hydration energy is more than lattice energy but BaSO₄ (as hydration energy is less than lattice energy) is insoluble in water.

(7) Solubility of alkaline earth metal carbonates and sulphates in water decrease down the group

Explanation: The size of anions being much larger compared to cations, the lattice enthalpy will remain almost constant within a particular group. Since the hydration enthalpies decrease down the group, solubility will decrease as found for alkaline earth metal carbonates and sulphates.

(8) Alkali metals are not found in nature

Explanation: Alkali metals are highly reactive because of low ionization enthalpy value and therefore are not found in nature. They are present in combined state only in form of halides, oxides etc.

(9) Sodium is less reactive than potassium

Explanation: Ionization Energy of potassium is less than sodium because of large size or less effective nuclear charge. Thus, potassium is more reactive than sodium

(10) NaOH is a stronger base than LiOH

Explanation: The M-OH bond in hydroxides of alkali metal is very weak and can easily ionize to form M⁺ ions and OH⁻ ions. This accounts for the basic character. Since ionization energy decreases down the group bond between metal and oxygen becomes weak. Therefore basic strength of hydroxides increases accordingly. Thus NaOH is a stronger base than LiOH

(11) Alkali metals are kept in paraffin or kerosene

Explanation: Alkali metals react explosively with water forming metal hydroxides along with hydrogen. Hydrogen gas released immediately catches fire. Thus alkali metals are highly sensitive towards air and water and hence are kept therefore in kerosene or paraffin oil

(12) Except for Be and Mg, the alkaline earth metal salts impart characteristic colours to the flame

Explanation: Beryllium and magnesium atoms are comparatively smaller and their ionization energies are very high. Hence, the energy of the flame is not sufficient to excite their electrons to higher energy levels. These elements, therefore, do not give any colour in Bunsen flame