S-BLOCK ELEMENT

- S-block elements are those, in which last electron enters the outermost s-orbital.
- As the s-orbital can accommodate only two electrons, this block includes only two groups, group 1/IA and Group 2/IIA.
- General electronic configuration of s-block element is ns^{1-2}, where, n = number of outermost shell.
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Group 1/IA, Alkali metals	Group 2/IIA, Alkaline-earth metals
General configuration — $= ns^1$	General configuration — $= ns^2$
Li = [He] $2s^1$	Be = [He] $2s^2$
Na = [Ne] $3s^1$	Mg = [Ne] $3s^2$
K = [Ar] $4s^1$	Ca = [Ar] $4s^2$
Rb = [Kr] $5s^1$	Sr = [Kr] $5s^2$
Cs = [Xe] $6s^1$	Ba = [Xe] $6s^2$
Fr = [Rn] $7s^1$, (Radioactive element)	Ra = [Rn] $7s^2$, (Radioactive element)

- Group 1 elements are called alkali metals, because they form strongly alkaline oxides and hydroxides.

while dissolved in water

↳ Group 1 elements have loosely held valence electron in their outer shell and form univalent, ionic and colourless compounds.

↳ Group 1 elements are all metals, excellent conductors of electricity and typically soft and highly reactives.

↳ Group 1 elements have single valence electron which is weakly held and is readily removed. In contrast, electrons that are remaining closer to the nucleus, more tightly held and removed only with great difficulties.

↳ Group 2 elements are called alkaline earth metals because their oxides and hydroxides are alkaline in character and these are found on Earth's crust.

↳ Group 2 elements are bivalent and generally form colourless ionic compounds.

↳ Group 2 elements are highly reactive but less reactive than Group 1 elements.

↳ Group 2 elements have two valence electrons

Diagonal relationship:— It has been observered that elements of second period show similarities with the elements of the third period present diagonally to each other, though belonging to different groups. This is called diagonal relationship.

	Group1	Group2	Group13
Second period →	Li	Be	B
Third period →	Na	Mg	Al

Therefore —

Li resembles with Mg and

Be resembles with Al

Properties:—

1) Atomic and Ionic radii —

- Alkali metals have the largest atomic and ionic radii in their respected periods of the periodic table due to low effective nuclear charge.
- On moving down the group, the atomic and ionic radii of s-block elements increase, therefore the first member of each group has the smallest size in its group.

Explanation — On moving down the group, the radii increase due to gradual increase in the number of the shells and the screening effect.

↳ Atomic and ionic radii of the members of the alkaline earth metals are smaller than the corresponding members of the alkali metals.

Explanation — The alkaline earth metals have a higher nuclear charge than alkali metals; therefore the electrons are attracted more towards the nucleus. As a result, their atomic and ionic radii are smaller than those of alkali metals.

2) Ionization Enthalpy — (Energy required to remove an electron from an atom in its gaseous state is known as ionization enthalpy)

↳ Alkali metals have the lowest ionization enthalpy in each period.

Explanation — The atoms of alkali metals are largest in their respective periods and therefore the valence electrons are loosely held by the nucleus. By loosing the valence electron, they acquire the stable noble gas configuration. Therefore, they have low ionization enthalpies.

↳ Within the group, the ionization enthalpies of S-block elements decrease down the group.

Explanation — On moving down the group, the atomic size increases and the magnitude of screening effect (number of inner shells) also increases and consequently, the ionization enthalpy decreases down the group.

↳ ~~The~~ 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 very difficult to remove the second electron from the stable noble gas configurations and hence their second ionization enthalpy values are very high.

$$IE_1 \ll IE_2$$

where, IE_1 = 1st ionization enthalpy of alkali metal.

IE_2 = 2nd " " " " "

For example —

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

$$Na(g) \xrightarrow[(-e^-)]{IE_1} Na^+(g) \xrightarrow[(-e^-)]{IE_2} Na^{2+}(g)$$

Na^+: $(1s^2 2s^2 2p^6)$ = Ne (Noble gas) (Stable configuration)

Na^{2+}: $(1s^2 2s^2 2p^5)$ → unfavorable

Therefore, $IE_1 << IE_2$

↳ The alkaline earth metals have low ionization enthalpies due to fairly large size of the atoms. But a comparison of the ionization enthalpies of group 1 & group 2 show that the element present in the second group have higher values as compared to those of group 1, because they have smaller size than corresponding group 1 elements and electrons are more attracted towards the nucleus of the atoms.

↳ Since, the atomic size increases down the group, the ionization enthalpies decrease down the group of alkaline earth metals like alkali metals.

↳ Although, the first ionization enthalpy (IE_1)

values of alkaline earth metals are higher than those of alkali metals, the second ionization enthalpy (IE_2) values of alkaline earth metals are much smaller than those of alkali metals.

Explanation — In case of alkali metals, the second electron is to be removed from a cation which has already acquired a noble gas configuration. On the other hand, in the alkaline earth metals, the second electron is to be removed from a monovalent cation which still has one electron in the outermost shell. Thus the second electron in the alkaline earth metals can be removed easily.

For example —

$$Na(g) \xrightarrow[(-e^-)]{IE_1} Na^+(g) \xrightarrow[(-e^-)]{IE_2} Na^{2+}(g)$$

$Na(g)$: $(1s^2 2s^2 2p^6 3s^1)$

$Na^+(g)$: $(1s^2 2s^2 2p^6)$ ⇓ (stable configuration)

$Na^{2+}(g)$: $(1s^2 2s^2 2p^5)$

$$Mg(g) \xrightarrow[(-e^-)]{IE_1} Mg^+(g) \xrightarrow[(-e^-)]{IE_2} Mg^{2+}(g)$$

$Mg(g)$: $(1s^2 2s^2 2p^6 3s^2)$

$Mg^+(g)$: $(1s^2 2s^2 2p^6 3s^1)$ ⇓ (Not very stable)

$Mg^{2+}(g)$: $(1s^2 2s^2 2p^6)$ ⇓ (Stable configuration)

	IE_1	IE_2
Na ⟶	496 $kJ\,mol^{-1}$	4562 $kJ\,mol^{-1}$
Mg ⟶	737 $kJ\,mol^{-1}$	1450 $kJ\,mol^{-1}$

↳ The third ionization enthalpy (IE_3) values of alkaline earth metals will be very high, as the third electron is to be removed from stable noble gas configuration.

3) Electronegativity — (Tendency to attract electron is known as electronegativity, denoted by 'X')

↳ The electronegativity values decrease down the group for alkali and alkaline earth metals. Hence, the electron releasing tendency or electropositive character increases down the group.

Explanation — As the ionization energies decrease down the group, the electronegativity values decrease down the group and electron releasing tendency or electropositive character increases.

↳ Alkaline earth metals are more electronegative than alkali metals because of comparatively higher ionization energies.

There have been different methods to calculate electronegativity of elements—

a) Pauling's bond energy scale.

b) Mulliken's scale and

c) Alfred - Rochow's scale

a) Pauling's bond energy scale —

Pauling's defined electronegativity as the power of an atom in a molecule to attract electrons to itself. This method makes a use of bond energies; that is the energy required to break a bond to get neutral atoms. He considered the formation of AB molecule by combination of A_2 and B_2 molecules.

$$A_2 + B_2 = 2AB$$

or $$\frac{1}{2} A_2 + \frac{1}{2} B_2 = AB.$$

This reaction may also be written as —

$$\frac{1}{2}(A-A) + \frac{1}{2}(B-B) = A-B$$

If E_{A-A}, E_{B-B} and E_{A-B} are the bond dissociation energies (in kcal) of A-A, B-B and A-B bonds respectively, then —

$$E_{A-B} - \sqrt{(E_{A-A}) \times (E_{B-B})} = \Delta'$$

and $$\boxed{0.182\sqrt{\Delta'} = \chi_B - \chi_A}$$

where, χ_A = electronegativity of atom A

and χ_B = " " " B

Q. Calculate the electronegativity of fluorine from the following data —

$E_{H-H} = 104.2$ k Cal mol^{-1}, $E_{F-F} = 36.6$ k Cal mol^{-1}

$E_{H-F} = 134.6$ k Cal mol^{-1} and $\chi_H = 2.1$

Soln — From Pauling's eqn —

$$\Delta' = E_{H-F} - \sqrt{(E_{H-H}) \times (E_{F-F})}$$

$$= 134.6 - \sqrt{104.2 \times 36.6}$$

$$= 72.85 \text{ k cal mol}^{-1}$$

Therefore, $\chi_F - \chi_H = 0.182 \times \sqrt{\Delta'}$

$$= 0.182 \times \sqrt{72.85}$$

$$= 1.55$$

$$\therefore \chi_F = 1.55 + \chi_H$$

$$= 1.55 + 2.1$$

$$= \underline{3.65}$$

Note :- Also, $\boxed{\Delta' = 30\,(\chi_B - \chi_A)}$ [For molecule AB]

b) Mullikan's scale —

↳ Mullikan suggested that, the electronegativity of an atom is the average of the ionization enthalpy (IE) and electron affinity (EA) of the atom.

↳ Electron affinity is the amount of energy released when an electron is added to a gaseous isolated atom or ion.

↳ Therefore,

$$\chi_M = \frac{IE + EA}{2}$$

When, values are in eV —, when values are in kJ/mol —

$$\chi_M = \frac{IE + EA}{5.6}, \quad \chi_M = \frac{IE + EA}{540}$$

Q. Calculate the electronegativity of Chlorine atom on Mullikan's scale [$IE_{Cl} = 13.0\ eV$, $EA_{Cl} = 4.0\ eV$]

Soln. —

$$(\chi_M)_{Cl} = \frac{(IE)_{Cl} + (EA)_{Cl}}{5.6}$$

$$= \frac{13 + 4}{5.6}$$

$$= 3.03$$

c) Alfred - Rochow's scale —

↳ Alfred - Rochow's defined electronegativity as the electrostatic force of attraction exerted by the nucleus of an atom on the valence electrons.

$$\chi = 0.359 \times \frac{Z_{eff}}{r^2} + 0.744$$

where,
r = Covalent radius of the atom in anstrom unit

Z_{eff} = Effective nuclear charge.

Q. Calculate the electronegativity of C-atom following Alfred-Rochow's approach. The covalent radius of carbon atom is 0.77 Å. [Z_{eff} = 2.90]

Soln:—

$$\chi = 0.359 \times \frac{Z_{eff}}{r^2} + 0.744$$

$$= \frac{0.359 \times 2.90}{(0.77)^2} + 0.744$$

$$= \underline{\underline{2.50}}$$

Group 1 (Alkali metals)—

	Li	Na	K	Rb	Cs
Atomi radii (Å) →	1.52	1.86	2.27	2.48	2.65
Ionic radii (Å) →	0.76	1.02	1.38	1.52	1.67
First IE ($kJmol^{-1}$) →	520.1	495.1	418.6	402.9	375.6
Second IE ($kJmol^{-1}$) →	7296	4563	3069	2650	2420
Pauling's electronegativity →	1.0	0.9	0.8	0.8	0.7

Group 2 (Alkaline-earth metals) —

	Be	Mg	Ca	Sr	Ba	Ra
Atomi radii (Å) →	1.12	1.60	1.97	2.15	2.22	—
Ionic radii (Å) →	0.31	0.72	1.00	1.18	1.35	1.98
First IE ($kJ mol^{-1}$) →	899	737	590	549	503	509
Second IE ($kJ mol^{-1}$) →	1757	1450	1145	1064	965	979
Pauling's electronegativity →	1.5	1.2	1.0	1.0	0.90	—

↳ Summery —

Across (→):

i) Effective nuclear charge increases

ii) Ionization enthalpy increases

iii) Electronegativity increases

Li	Be
Na	Mg
K	Ca
Rb	Sr
Cs	Ba
Fr	Ra

Down the group (↓):

i) Atomi radii increases

ii) Ionic radii increases

iii) Number of shells increases

iv) Screening effect increases

Up the group (↑):

i) Ionization enthalpy increases

ii) Electronegativity increases.

Right to left (←):

i) Atomi radii increases

ii) Ionic radii increases

iii) Screening effect increases

↳ The melting point (m.p.) and boiling point (b.p.) of alkali metals are very ~~high~~ low due to the presence of interatomic bonds. Weak bonds are due to large atomic size and single valence electron

↳ In the group, on moving from top to bottom, m.p. and b.p. of alkali metals decrease because with increasing size of atoms, the repulsion between non-bonding electrons increase.

	Li	Na	K	Rb	Cs
m.p. (°C) =	181	98	63	39	28.5
b.p. (°C) =	1347	881	766	688	705

↳ The m.p. and b.p. of alkaline earth metals donot show any trend. Howevere, m.p. and b.p. of these elements are highere than those of alkali metals. This is because, the atoms of alkaline earth metals have smallere size as compared to alkali metals and these elements form strong bond due to two valence electrons.

	Be	Mg	Ca	Sr	Ba	Ra
m.p. (°C) =	1287	649	839	768	727	700
b.p. (°C) =	2500	1105	1494	1381	1850	1700

Anomalous behaviour of Lithium (Li) — Group 1

↳ Lithium shows different properties from other alkali metals due to following reasons—

i) The size of lithium and its ion (Li^+) is the smallest of all the alkali metals and ions.

ii) It has highest ionization enthalpy and least electropositive character.

iii) The polarising power of Li^+ ion is the greatest of all alkali metal ions due to its small size which results in the covalent character of its compounds.

↳ Lithium shows following different properties from other alkali metals —

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

ii) The m.p. and b.p. of Lithium ~~at~~ is much higher than those of other alkali elements.

iii) Lithium is least reactive metal of all alkali metals.

iv) Only Lithium combines with carbon and silicon forming the carbide (Li_2C_2) and the silicide (Li_6Si_2) among all the alkali metals.

v) Lithium forms only monoxide, Li_2O with oxygen, while other alkali metals form superoxide (MO_2) and peroxide (M_2O_2).

vi) Lithium readily reacts with nitrogen forming nitride while other alkali metals do not react –

$$6Li + N_2 \longrightarrow 2Li_3N^-$$

vii) The hydrides of Lithium, (LiH) is more stable as compared to the hydrides of other members of the family.

viii) The salts of Lithium have lower ionic character than salts of other alkali metals; This is because of high polarising power of Li^+ ion.

ix) Li^+ ion has polarizing power and hence forms covalent bond. Thus Lithium salts are soluble in non-polar solvents like alcohol. But the salts of other alkali metals are ionic, therefore insoluble in non-polar solvents.

x) The Lithium ion (Li^+) and its compounds are more heavily hydrated than those of other alkali metal ions and their compounds due to the smallest size of Li^+ ion.

xi) Due to low electropositive character of

Lithium, its various salts are less stable and therefore decompose to give oxides —

$$2LiOH \longrightarrow Li_2O + H_2O$$

$$4LiNO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2$$

$$Li_2CO_3 \longrightarrow Li_2O + CO$$

xii) Lithium when heated with NH_3, it forms imide (Li_2NH), while other alkali metals form amides (MNH_2).

Diagonal relationship of Lithium (Li) and Magnesium (Mg) —

↳ Lithium resembles to Magnesium due to diagonal relationship.

↳ Some common characteristics of Lithium and Magnesium are —

i) The atomic radius of Lithium is closed to that of Magnesium —

Radius of Lithium = 152 pm

" " Magnesium = 160 pm

The ionic radii are also similar —

Li^+ = 76 pm and Mg^{2+} = 72 pm

ii) Both Lithium and Magnesium have similar electronegativities —

$Li = 1.0$, $Mg = 1.2$

iii) Both $LiOH$ and $Mg(OH)_2$ are weak bases.

iv) Unlike the other members of the group, Lithium reacts with N_2 to form Lithium nitride. Magnesium also reacts in a similar way —

$$6Li + N_2 \xrightarrow{heat} 2Li_3N$$
(Lithium nitride)

$$3Mg + N_2 \xrightarrow{heat} Mg_3N_2$$
(Magnesium nitride)

v) Both Lithium and Magnesium harder and lighter elements than other elements in the respective groups.

vi) Both Lithium and Magnesium combine with Oxygen to form monoxides. Other group members form peroxides and superoxides.

$$4Li + O_2 \xrightarrow{heat} 2Li_2O$$

$$2Mg + O_2 \xrightarrow{heat} 2MgO$$

vii) The carbonates of both Lithium and Magnesium metals decompose on heating to form oxide and evolve CO_2.

viii) Both LiCl and $MgCl_2$ are soluble in alcohol.

ix) The nitrates of both Lithium and Magnesium evolve nitrogen dioxide and oxygen on heating –

$$4LiNO_3 \longrightarrow 2Li_2O + 4NO_2 + O_2$$

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

x) The hydroxides of both Lithium (Li) and Magnesium (Mg) metals decompose on strong heating to form respective oxides –

$$2LiOH \xrightarrow{heat} Li_2O + H_2O$$

$$Mg(OH)_2 \xrightarrow{heat} MgO + H_2O$$

Anomalous behaviour of Beryllium (Be) — (Group 2)

↳ Beryllium, the first member of the alkaline earth metals family show anomalous behaviour as compared to magnesium and the rest of the members. It is mainly because of the following reasons ——

i) Small size of the Beryllium atom and Be^{2+} ion.

ii) High ionization enthalpy and electronegativity of Beryllium.

↳ Some of the important differences between Beryllium and other members of the family are —

i) Beryllium is harder than other alkaline earth metals.

ii) Beryllium possesses higher b.p. and m.p. as compared to other alkaline earth metals.

iii) Beryllium doesnot react with water even at elevated temperature. On other hand, other alkaline earth metals react with boiling water.

iv) Beryllium doesnot combine directly to form hydride ~~wea~~ whereas other alkaline ~~each~~ earth metals do.

v) The oxide and hydroxide of Beryllium are amphoteric while those of other elements in the group are basic.

vi) Beryllium doesnot exhibit co-ordination number more than four because its valence shell has only four orbitals. The other members of the group can have co-ordination number of six by making use of d-orbitals present in them.

vii) Beryllium carbide with water gives methane while Magnesium carbide gives acetylene —

$$Be_2C + 2H_2O \longrightarrow 2BeO + \underset{\text{(methane)}}{CH_4}$$

$$MgC_2 + 2H_2O \longrightarrow Mg(OH)_2 + \underset{\text{(acetylene)}}{C_2H_2}$$

viii) Because of high ionization enthalpy and small size, Beryllium forms covalent compounds while other members of the group form ionic compounds. Because of co-valent character, salts of Beryllium are easily hydrolysed —

eg:- $BeCO_3 + 4H_2O \longrightarrow [Be(H_2O)]^{2+} + CO_3^{2-}$

Diagonal similarities of Beryllium (Be) and Aluminium (Al) —

↳ Beryllium shows some similarities with Aluminium due to diagonal relationship. Some common examples are —

i) Both Beryllium and Aluminium have same electronegativity values (Be = 1.5 and Al = 1.5) and their charge/radius ratios (Be = 0.064 and Al = 0.060) are similar indicating similar field strengths.

ii) Both Beryllium and Aluminium form covalent compounds.

iii) Both Beryllium and Aluminium form halides that have similar solubilities.

iv) BeO and Al_2O_3 are amphoteric, i.e., they dissolve in acids as well as in bases —

$$BeO + 2HCl \longrightarrow BeCl_2 + H_2O$$

$$BeO + 2NaOH \longrightarrow Na_2BeO_2 + H_2O$$

(Sodium berellate)

$$Al_2O_3 + 6HCl \longrightarrow 2AlCl_3 + 3H_2O$$

$$Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$$

(Sodium meta-aluminate)

v) Both Beryllium and Aluminium metals are weakly electropositive in nature.

vi) The carbides of Beryllium and Aluminium liberates methane with water —

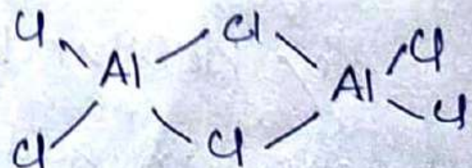
$$Be_2C + 2H_2O \longrightarrow 2BeO + CH_4$$

$$Al_4C_3 + 6H_2O \longrightarrow 2Al_2O_3 + 3CH_4$$

vii) Salts of Beryllium and Aluminium metals form hydrated ions, for example — $[Be(H_2O)_4]^{2+}$ and $[Al(H_2O)_6]^{3+}$ in aqueous solutions.

viii) $BeCl_2$ like Al_2Cl_6 has a bridged polymeric structure in vapour phase as shown below —

And



Both these chlorides are soluble in organic solvents and are strong Lewis acids. They are used as Friedel Craft catalysts.

x) Because of similarity in charge/size ratios both Beryllium and Aluminium have strong tendency to form complexes. For example —

$[Be(C_2O_4)_2]^{2-}$ and BeF_4^{2-} are tetrahedral complexes formed by Beryllium.

And Aluminium forms octahedral complexes like AlF_6^{3-} and $[Al(C_2O_4)_3]^{3-}$
