

## Introduction

The elements in the long form of the periodic table have been divided into four blocks — s, p, d and f - block elements. This classification is based on the type of the atomic orbital in which the outermost electron is located. The s orbital can accommodate two electrons, while the three p orbitals can accommodate a maximum of six electrons. Thus there are two groups of s - block elements — Groups 1 and 2 whose electronic configurations are represented as [Noble gas]  $ns^1$  and [Noble gas]  $ns^2$  respectively. There are six groups of p - block elements — Groups 13 to 18. The s and p block elements are collectively called the main - group elements or representative elements. The d and f - block elements are located between the s and p - block elements.

The s - block elements are located on the left of the periodic table and comprise of highly reactive metals. The p -block elements comprise of both metals and non-metals. On moving from left to right in the periodic table the metallic character decreases, while on moving downwards, in a particular group, it increases. Hence it is apparent that we come across both non-metals and metals in p- block elements with metallic character being more apparent in the heavier members of a group. It is more convenient for us to discuss the s and p - block elements separately since the elements in these two blocks differ significantly in their physical and chemical properties. Our main focus will be on general group trends alongwith comparative study of important compounds.

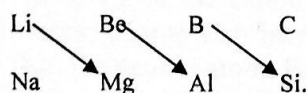
### I. s - block elements

The elements of groups 1 and 2 are called the s- block elements because the outermost electron(s) occupy the s -orbital. They are highly electropositive metals and form ionic compounds. They are referred to as alkali and alkaline- earth metals respectively. Due to their high reactivity the metals do not occur in the free state in nature but are present in form of halides, silicates, nitrates (Group 1) and silicates, carbonates, sulphates and phosphates (Group 2).

The elements show similarities in physical and chemical properties within a group — however the first elements exhibit considerable differences from the rest of the elements of the same group. This anomalous behaviour is mainly due to the following factors — (i) comparatively higher value of the first ionization energy and electronegativity (ii) small size (iii) high polarizing power of the cation, consequently a greater tendency to form covalent compounds (Fajan's rules) and complexes, (iii) inability to display a coordination number exceeding 4 due to non-availability of d - orbitals in the valence shell.

### Diagonal Relationship

A close similarity is observed in certain cases between the first element of a group with the second element of the following group. This is referred to as the diagonal relationship and is observed in the following pairs.



On descending a group, the size increases and the charge on the ion remains the same, thus the polarizing power decreases. On moving across a period, the charge on the ion increases while the size decreases, causing the polarizing power to increase. On moving diagonally the effects of size and charge partly cancel each other, so the polarizing powers are comparable and there is no significant change in properties, particularly for the lighter elements. The

elements along the diagonal also have comparable electronegativities — Li (1.0) & Mg (1.2); Be (1.5) & Al (1.5). However, diagonal similarity is much weaker than the general group similarity.

### Group 1 — the Alkali Metals

The elements in this group are lithium, sodium, potassium, rubidium, caesium and the radioactive, francium. They have one electron in their valence shell ( $ns^1$  configuration). Due to similarity in electronic configuration, they exhibit similarities in chemical properties. They are soft, highly reactive, univalent metals, excellent conductors of heat and electricity and form colourless ionic compounds. Lithium shows considerable differences from the lower members and exhibits diagonal relationship with magnesium.

The physical properties of alkali metals are summarized in Table 1.

**Table 1: Physical Properties of Group 1 Elements**

Property	Li	Na	K	Rb	Cs
Atomic Number	3	11	19	37	55
Electronic Configuration	[He]2s <sup>1</sup>	[Ne]3s <sup>1</sup>	[Ar]4s <sup>1</sup>	[Kr]5s <sup>1</sup>	[Xe]6s <sup>1</sup>
Covalent radius (pm)	123	156	203	216	235
Ionic radius (M <sup>+</sup> ) (pm)	60	95	133	148	169
Ionization Energy (KJmol <sup>-1</sup> )	520	496	419	403	376
Electron Affinity (KJmol <sup>-1</sup> )	60	53	48	47	46
Electronegativity	0.912	0.869	0.734	0.706	0.659
Melting Point (°C)	180.5	97.8	63.2	39.0	28.5
Boiling Point (°C)	1347	881	766	688	705
E <sup>0</sup> (M <sup>+</sup> → M) (V) against S.H.E,	-3.04	-2.71	-2.92	-2.92	-2.92

### General Group Trends

#### Size & Density

The atoms are the largest in the corresponding horizontal periods of the periodic table. On moving down the group the size of the atom increases as additional shells are introduced to accommodate the electrons. This out-weighs the effect of increased nuclear charge. Formation of the cation involves removal of the outermost shell of electrons; the effective nuclear charge also increases and the electrons are held tightly. Thus the cation is smaller than the neutral atom. Even so, the cations are quite large and an increase in size is observed on descending the group. The atoms are large and the atomic mass is low; thus the Group 1 elements have low densities.

#### Ionization Energy

The first ionization energies of the atoms are the lowest in the corresponding period. The atoms are large and the outermost electron is loosely held. Moreover; loss of this electron gives a stable noble gas configuration. On moving down the group, the ionization energy

decreases as the size increases. The second ionization energy is extremely high as it is difficult to remove an electron from a cation having noble gas configuration. The elements form unipositive cations, they are highly electropositive and the metallic character increases down the group.

### Melting and Boiling Points

These metals are soft and can be cut with a knife. At room temperature they adopt a body-centred cubic lattice. They have one electron that can participate in metallic bond formation and thus they have low cohesive energy, which is reflected in low melting and boiling points. On moving down the group, the metallic bonding becomes weaker and the melting and boiling points decrease.

### Electronegativity and Type of Bond

These elements have low electronegativity; consequently when they react with other elements a large electronegativity difference is possible and ionic bonds are formed.

### Flame Colours

The alkali metal atoms have low ionization energy and may emit electrons when irradiated with light, hence they find use in photocells. Electrons may also be excited to higher energy levels, for example in the flame test. When an alkali metal salt is introduced to a Bunsen burner flame, a characteristic colour is observed. (Li – red, Na – golden-yellow, K – lilac). The colour arises from electronic transitions, which occur due to excitation of electrons from valence shell to higher excited state energy levels. When the excited electrons return to the original level, the extra energy is emitted and lies in the visible range of the electromagnetic spectrum.

### Chemical Properties

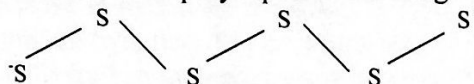
The alkali metals are very similar in their chemical properties, which are governed by the ease with which they can lose their valence electron. They are excellent reducing agents and react with water forming hydrogen.



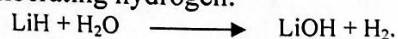
The reaction is highly exothermic.

They react with oxygen to form oxides, peroxides or superoxides depending on the metal. Lithium forms the oxide ( $\text{Li}_2\text{O}$ ), sodium, the peroxide ( $\text{Na}_2\text{O}_2$ ) while others form the superoxide ( $\text{MO}_2$ ,  $\text{M} = \text{K, Rb, Cs}$ ). Lithium is the only element that forms a nitride by direct combination with nitrogen.

Alkali metals react with sulphur forming sulphides e.g.  $\text{Na}_2\text{S}$  and polysulphides  $\text{Na}_2\text{S}_n$  where  $n = 2$  to  $6$ . The polysulphides have zig-zag chain of sulphur atoms.

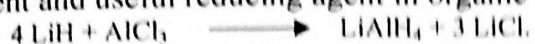


The metals react with hydrogen forming ionic hydrides,  $\text{M}^+\text{H}^-$ . The stability of the hydrides decreases down the group. The hydrides are ionic solids and are the only example of species containing  $\text{H}^-$  ion. During electrolysis hydrogen is liberated at the anode. They react with water liberating hydrogen.





Lithium hydride reacts with aluminum chloride to form lithium aluminium hydride, an excellent and useful reducing agent in organic chemistry



Sodium hydride reacts with trimethyl borate to form another useful compound, sodium borohydride



The alkali metals react violently with halogens to form halides. The reactivity of an alkali metal towards a particular halogen increases down the group. The reactivity of halogen towards an alkali metal follows the order:  $\text{F}_2 > \text{Cl}_2 > \text{Br}_2 > \text{I}_2$ . The general reactions of alkali metals are summarized in Table 2.

**Table 2: Reaction of Group 1 Metals**

Reagent	Reaction	Remarks
$\text{O}_2$	$\text{M} + \text{O}_2 \longrightarrow \text{M}_x\text{O}_y$	$x = 2, y = 1$ for Li (monoxide); $x = 2, y = 2$ for Na (peroxide) $x = 1, y = 2$ for others (superoxide)
S	$2\text{M} + \text{S} \longrightarrow \text{M}_2\text{S}$	Polysulphides also formed
$\text{H}_2\text{O}$	$\text{M} + 2\text{H}_2\text{O} \longrightarrow \text{MOH} + \text{H}_2$	Slow for Li, explosive for K
$\text{H}_2$	$2\text{M} + \text{H}_2 \longrightarrow 2\text{MH}$	Ionic hydrides formed
$\text{X}_2$ (halogen)	$2\text{M} + \text{X}_2 \longrightarrow 2\text{MX}$	All halogens react
$\text{N}_2$	$6\text{Li} + \text{N}_2 \longrightarrow 2\text{Li}_3\text{N}$	Only Li reacts
ROH	$2\text{M} + 2\text{ROH} \longrightarrow 2\text{MOR} + \text{H}_2$	Slow for Li, Vigorous for others
$\text{NH}_3$	$\text{M} + \text{NH}_3(\text{e}) \longrightarrow [\text{M}(\text{NH}_3)_x]^+ + \text{e}^- (\text{NH}_3)$	Excellent reducing agents