

p-BLOCK ELEMENTS

GROUP-IIIA (13) BORON FAMILY

The very word "experiment" refers to a situation where we can tell others what we have done and what we have learned

Niels Bohr

7.1 INTRODUCTION

The elements boron (B), aluminium (Al), gallium (Ga), indium (In) and thallium (Tl) constitute Group III A or 13 of the periodic table. They belong to *p*-block elements. The electronic configuration in their outer most orbit is $ns^2 np^1$. These elements not only show marked similarities among them but also show a very wide variation in properties. Boron is a typical non-metal, aluminium is a metal but shows many chemical similarities to boron and the remaining elements are almost exclusively metallic in character. Although unipositive oxidation state is the characteristic one, for all the members of the group, the unipositive state occurs in compounds of all the elements except boron. In the case of thallium, the unipositive oxidation state is the stable one and in fact it shows similarities to so many elements such as alkali metals, silver and mercury. Hence, it is nick named as **duck-billed platypus** among the elements. A feature of the chemistry of boron is the existence of large number of electron-deficient species which pose formidable problems in valence bond theory. These include not only the hydrides but also organic and metallic derivatives of the hydrides, the metal borides etc.

7.2 ABUNDANCE

Boron and aluminium of this family are considered to be familiar elements whereas gallium, indium and thallium are less familiar elements. Particularly aluminium is abundant in nature. The elements occur in nature in the following proportions in the crust of the Earth.

Boron $3 \times 10^{-4}\%$

Aluminium 8.13%

Gallium $1.5 \times 10^{-3}\%$

Indium $1 \times 10^{-5}\%$

Thallium 10^{-4} to $10^{-5}\%$

Aluminium is of course the most abundant of all the metals and the third most abundant of all the elements. The

comparative scarcity of boron may be partially due to the ease with which the nuclei of its atoms are transmuted by natural bombardment process. Boron is well known, however, because of the existence of concentrated deposits of its compounds particularly in arid regions and because of the desirable properties of many of its compounds which have necessitated large-scale recovery of boron materials. Gallium, indium and thallium never found in concentrated deposits and until recently they were never recovered in sizable quantities.

7.3 OCCURRENCE

Borax $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$ or $\text{Na}_2[\text{B}_4\text{O}_7(\text{OH})_4] \cdot 8\text{H}_2\text{O}$ is the principal source of boron. In India, it occurs in deserts and in the United States, at California. It occurs in hot springs and lakes in volcanic regions. Aluminium occurs mainly as bauxite (a hydrated oxide mineral), cryolite $\text{Na}_3[\text{AlF}_6]$ and also in the aluminosilicate minerals such as mica and feldspar. Gallium, indium and thallium occur in traces in sulphide minerals. Gallium is also found in traces in bauxite.

7.4 ELECTRONIC CONFIGURATION

The electronic configuration of the elements of Group III A are listed in Table 7.1.

From Table 7.1, it follows that all the elements of Group III A have three electrons in their valence shell, two electrons in the *s*-orbital and one electron in the *p*-orbital, i.e., $ns^2 np^1$.

Owing to the similarity in electronic configuration of their outermost shell, they closely resemble in their physical and chemical properties. However, it is to be noted that the penultimate shell of those five elements differ in the configuration; the penultimate shell of boron is having s^2 configuration; the penultimate shell of other three elements aluminium is having $s^2 p^6$ and those of other three elements are having $s^2 p^6 d^{10}$ electrons. Thus, boron is expected to differ from aluminium and further boron and aluminium are

Table 7.1 Electronic configuration of Group III A elements

Element	Atomic number	Electronic configuration	Electronic configuration of valence shell
B	5	$1s^2 2s^2 2p^1$	$2s^2 2p^1$
Al	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	$3s^2 3p^1$
Ga	31	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^1$	$4s^2 4p^1$
In	49	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^1$	$5s^2 5p^1$
Tl	81	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14} 5s^2 5p^6 5d^{10} 6s^2 6p^1$	$6s^2 6p^1$

expected to have some similar properties because penultimate shell of the B and Al has noble gas kernel and are different from other three elements.

7.5 PHYSICAL PROPERTIES

1. Atomic and ionic radii: The atomic and ionic radii are given in Table 7.2. Atomic and ionic sizes of Group III A elements do not increase regularly. The greater difference in the atomic radius between boron and aluminium is due to the fact that boron has lesser number of electrons (i.e., two electrons) in its inner shell than aluminium (i.e., eight electrons). Thus, in boron, the outer most electrons experience lesser shielding effect and greater nuclear attraction.

Table 7.2 Atomic and ionic radii of Group III A elements

Element	Covalent radius (pm)	Metallic radius (pm)	Ionic radius (pm)	
			M ⁺	M ³⁺
B	80	(88.5)	—	27
Al	125	143	—	53.5
Ga	125	135	120	62
In	150	167	140	80
Tl	155	170	150	88.5

2. Density: As we move down in Group III A, density increases. However, boron and aluminium have comparatively low density. This can be attributed to their lower atomic weights as compared to the remaining elements. Ga is unusual because the liquid expands when it forms solid. Therefore, solid Ga is not denser than liquid Ga.

3. Melting point: The elements of Group III A do not show a regular change in their melting points with increase in atomic number. The melting point of boron is very high because it has the structure of giant covalent polymer in both solid and liquid states. The melting points decrease from B to Ga and then increase. Low melting point of Ga (29.8°) is attributed to the fact that it consists of only Ga₂ molecules; it remains liquid up to 2000°C and hence used in high-temperature thermometry.

4. Boiling point: Boiling points of Group III A elements follow a regular decrease from boron to thallium. This

shows that the strength of bonds holding the atoms in their liquid state decreases from boron to thallium. Note that the boiling point of Ga is in regular order with others whereas the melting point is not. The very low melting point is due to the unusual crystal structure, but the structure no longer exists in the liquid.

5. Heat of sublimation: It decreases regularly on moving down the group.

6. Ionization energies: The ionization energies increase as expected, i.e., first ionization energy < second ionization energy < third ionization energy. The sum of the first three ionization energies for each element is very high.

As the *p*-electron is less tightly held as compared to the *s*-electrons, the first ionization energy has been rather low in each case. The second and third ionization energies have been considerably higher.

Ionization energy decreases from boron to aluminium but does not change appreciably as we move to gallium, indium and thallium. Decrease in ionization energy from boron to aluminium is attributed to the increased size. Further, penultimate shell in both boron and aluminium has inert gas configuration (He configuration in the case of B and Ne configuration in the case of Al) whereas the penultimate shell in all the three remaining elements, i.e., gallium, indium and thallium has 18 electrons [(*n*-1)*s*²*p*⁶*d*¹⁰]. The extra *d*-electrons fail to shield the nuclear charge effectively because shielding by electrons present in various orbitals has been found to be in the order *s* > *p*² > *d* > *f*. Therefore, the outer electrons in the case of gallium, indium and thallium are held more tightly by the nucleus. Consequently, their atoms become smaller and thus their ionization energies become higher than expected. The ionization energy of thallium is further affected because of the poor shielding of 14 *f*-electrons present in the inner penultimate shell, and it is even more than that of Al, Ga and In.

For gallium and indium, the electronic configuration of the species left after the removal of three electrons is [Ar]3*d*¹⁰ and [Kr] 4*d*¹⁰, respectively whereas for thallium the species so formed has the configuration [Xe] 4*f*¹⁴5*d*¹⁰. Thus, the fourth ionization energies of these three elements do not involve the removal of an electron from a noble gas

Table 7.3 Density, melting point, boiling point, electronegativity and heat of sublimation of Group III A elements

Element	Density (g cm ⁻³)	Melting point (K)	Boiling point (K)	Heat of sublimation (kJ mol ⁻¹)	Electro negativity
B	2.35	2453	3923	564	2.0
Al	2.7	933	2740	324	1.5
Ga	5.9	303	2676	273	1.6
In	7.31	430	2353	241	1.7
Tl	11.85	576	1730	179	1.8

Table 7.4 Ionization energies of Group III A elements

	Ionization energies kJ mol ⁻¹			Sum of three
	Ist	IInd	IIIrd	
B	801	2427	3659	6887
Al	577	1816	2744	5137
Ga	579	1979	2962	5520
In	558	1820	2704	5082
Tl	589	1971	2877	5437

configuration and the difference between the fourth and the third ionization energies is not nearly so large as for boron and aluminium.

7. Electronegativity: Among the Group III A elements, boron has the maximum electronegativity. It decreases from boron to aluminium as expected. However, from aluminium to thallium, it increases instead of decreasing in minimum to thallium, it increases instead of decreasing contrary to the expectation. This is again attributed to the poor shielding of *d*-electrons in gallium and indium and *d*- and *f*-electrons in thallium (Table 7.3).

8. Oxidation states: The Group III A elements contain three more electrons in their outer most orbit than the stable inert gas (with B and Al) or pseudo inert gas (with Ga, In and Tl) structures. Hence, a uniform +3 oxidation state is expected. This state is characteristic of all the elements. The electronic arrangement in the outermost orbit of these elements $ns^2 np^1$ also suggests +1 oxidation states.

Outer electronic configuration

	ns	np	
Ground state	$\uparrow\downarrow$	\uparrow	+1
Excited state	\uparrow	\uparrow	+3

In thallium, +1 oxidation state is well known. A few unipositive indium compounds mostly halides are known, in the solid state. In boron, there are no evidences for +1 oxidation state. Though aluminium and gallium compounds are prepared at elevated temperatures, they yield most unstable unipositive compounds and the +1 oxidation state of these elements is of least important.

The stability of +1 oxidation state increases more and more when we move down the group from B to Tl. Thus, Tl(I) compounds are more stable than Tl(III) compounds. This is attributed to the inert pair effect. **The two s-electrons in the outer shell tend to remain paired and are not participating in compound formation.** This pair of electrons is called **inert pair** and the effect is called **inert pair effect**. The inert pair effect increases gradually in gallium, indium and thallium compounds. For example, Ga⁺ compounds are unstable, In⁺ compounds are moderately stable, whereas Tl⁺ compounds are most stable. In fact, Tl(I) salts resemble alkali metals because thallium is having large size and low oxidation state. Some points of resemblance are as follows.

- TlOH is soluble in water yielding strong alkaline solution very similar to NaOH.
- Tl(I) cyanide, perchlorate, sulphate, nitrate, phosphate and carbonate are stable and isomorphous with alkali metal salts.
- TlF is having distorted NaCl-type structure whereas other thallos halides crystallize with CsCl structure.
- Like alkali metals, thallium (I) is known to form alums, e.g., Tl₂SO₄ · Al₂(SO₄)₃ · 24H₂O.

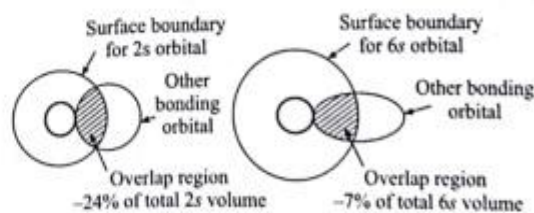


Fig. 7.1 Comparison of overlappings between small (2s) orbital and large (6s) orbital with the orbital of another atom

The stability of unipositive ions in aqueous solutions is consistent only in the case of thallium but for other elements the unipositive oxidation state disproportionates.



Stability of +3 oxidation state decreases regularly as the atomic number increases from boron to thallium.

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The stability of +1 oxidation state of Group III A elements should not be attributed strictly to inert pair effect. The inert pair effect explains what is happening, i.e., two electrons do not participate in bonding. The reason for not participating the two electrons in the outer most *s*-orbital in bonding can be explained basing on energy changes, i.e., excitation energy and the bond strength of the compound formed.

The small sized *2s* orbital of boron overlaps sufficiently with the orbital of another element (X) to yield strong M-X bonds that impart stability to the molecule. However, when we move downwards, the larger orbital (*5s* or *6s*) is involved because of which overlapping is poor giving rise to lower bond energy of the M-X bond.

As the small energy of the M-X bond is not sufficient to compensate for the excitation energy of *s*-electrons, it follows that the larger elements show increasing tendency to form univalent compounds.

Only boron is sufficiently electronegative to show any tendency towards a negative oxidation state. In the borides of the most highly electropositive elements, boron presumably exists in the -3 oxidation state.

9. Nature of bonding: Boron never forms B^{3+} cation because the sum of the three ionization energies is very large. Further because of the very small size of B^{3+} ion (20 pm) in its ionic compounds, the tripositive boron ion will have much polarizing power on the adjacent atoms which results in the covalent character (Fajans rules). Hence, in boron, +3 oxidation state is strictly covalent.

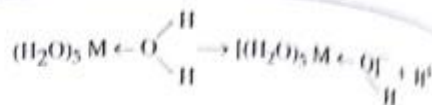
With other elements of Group III A, the +3 oxidation state is largely covalent. Tripositive cations are known in aqueous solutions for all the elements except boron. This is because of the fact that the hydration energies of tripositive cations overcome the ionization energies. Therefore, in aqueous solutions, they exist as hydrated cations and are greatly hydrolysed in solution. For example, in the case of $AlCl_3$, the energy changes are as follows:

$$\begin{aligned} \Delta H_{\text{hydration}} \text{ for } Al^{3+} &= -4665 \text{ kJ mol}^{-1} \\ \Delta H_{\text{hydration}} \text{ for } Cl^- &= -381 \times 3 \text{ kJ mol}^{-1} \\ \text{Total hydration energy} &= -5808 \text{ kJ mol}^{-1} \\ \text{Sum of the first three ionization energies} &= 5137 \text{ kJ mol}^{-1} \text{ for} \end{aligned}$$



Thus, total hydration energy evolved ($-5808 \text{ kJ mol}^{-1}$) more than offsets the ionization energy (5137 kJ mol^{-1}) required to convert Al to Al^{3+} .

The trivalent hydrated metal ions are having six molecules of water which are attached to them strongly giving an octahedral structure. These undergo hydrolysis to form acidic solutions. The strength of the metal oxygen bond in the hydrated ion would be able to weaken the O-H bonds. This causes some hydrolysis and protons get released giving acidic solution.



10. Electropositive character: Boron is a non-metal and its electropositive character is least. Aluminium is metal and is most electropositive. The remaining three elements gallium, indium and thallium are weakly metallic in nature and their electropositive character is less than that of aluminium and decreases from gallium to thallium.

The increase in electropositive character from boron to aluminium is ascribed to the increased size. The *extra* *d* electrons in case of gallium and indium whereas *d* and *f* electrons in the case of thallium do not shield the nuclear charge very effectively and these metals are, therefore, less electropositive. This is illustrated by the increase in ionization energy between aluminium and gallium even though the larger atom would be expected to have a lower value.

11. Electrode potentials: The standard electrode potentials E° for M^{3+}/M are given in Table 7.5.

Table 7.5 Standard electrode potentials E°

	M^{3+}/M (volts)	M^+/M (volts)
B	-0.87*	-
Al	-1.66	+0.55
Ga	-0.56	-0.79**
In	-0.34	-0.18
Tl	+1.26	-0.34

* For $H_2BO_3 + 3H^+ + 3e \longrightarrow B + 3H_2O$

** Value in acidic solution.

The standard electrode potentials E° for M^{3+}/M become less negative from Al to Ga to In. As the free energy change $\Delta G = nFE^\circ$ becomes more positive for the formation of the metal, the reaction $Al^{3+} + 3e \longrightarrow Al$ is not spontaneous. However, the reverse reaction $Al \longrightarrow Al^{3+} + 3e$ occurs spontaneously. As the standard potential becomes less negative down the group, so the reaction $M^{3+} \longrightarrow M$ occurs with less difficulty. Thus, the +3 oxidation state becomes less stable in aqueous solution on moving down the group. In a similar way, the increase in the stability of +1 oxidation state down the group is indicated by E° values for M^+/M . Thus, in thallium, +1 oxidation state is more stable than +3 oxidation state.

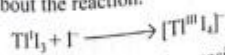
12. Complex formation: On account of their smaller size and greater charge, these elements have a much greater tendency to form complexes than the *s*-block elements. For example, the molecular trihalides or similar species of the elements of Group III A are still capable of accepting a pair of electrons and very large number of complexes such as $[BF_4]^-$, $[AlCl_4]^-$, $[GaCl_4]^-$ and $[InCl_4]^-$ are known. In boron, the second orbit is the outer most orbit and does not contain *d*-subshell but aluminium and other III A elements

will contain *d*-subshells in their outer most orbits. Hence, aluminium and the heavier elements are not restricted to an octet of electrons in their valence shells. Hence, for these elements, coordination numbers higher than four may be formed, e.g., $[\text{AlF}_6]^{3-}$ and $[\text{TlF}_6]^{3-}$. Thus, the covalence of boron is restricted to 4 only while the other elements can exhibit a covalence up to 6. Owing to this reason, boron cannot form complexes such as $[\text{BF}_6]^{3-}$.

B. Aqueous solution chemistry: All the tripositive aquo ions of these elements are acidic, that of aluminium the least and that of thallium the most so. Thus, aqueous solutions of their salts are appreciably hydrolysed, and salts of weak acids (e.g., carbonates and cyanides) cannot exist in contact with water. In acidic solution, aluminium is present as the $[\text{Al}(\text{H}_2\text{O})_6]^{3+}$ ion, as the acidity is decreased, polymeric hydrolysed species such as hydrated, $[\text{Al}_2(\text{OH})_2]^{4+}$ and $[\text{Al}_3(\text{OH})_4]^{5+}$ appear, then $\text{Al}(\text{OH})_3$ is precipitated, and finally, in alkaline solutions, aluminate anions such as $[\text{Al}(\text{OH})_4]^{-}$ and $[\text{Al}(\text{OH})_6]^{3-}$ and polymeric species such as $[\text{Al}(\text{OH})_2]_n$, $\text{AlOAl}(\text{OH})_2$ are formed. The chemistry of gallium is broadly similar to that of aluminium in this respect. Indium and thallium(III) hydroxides, however, are not amphoteric.

Redox potential data show that $\text{Al}^{3+}(\text{aq})$ is much less readily reduced than the other tripositive cations in aqueous solution. This, doubtlessly, arises partly from a more negative hydration free energy of the smaller Al^{3+} ion, but another important contributory factor is the increase in ionization energies between aluminium and gallium and between indium and thallium; there is relatively little variation in atomization enthalpies, and the overall variation in E° is, therefore, quite different from that in two preceding groups.

The E° value of $\text{Tl}^{3+}/\text{Tl}^+$, +1.26 V indicates that it is a powerful oxidant. The value of E° is, however, very dependent upon the anion present, because Tl(I) resembles an alkali metal ion by forming a few stable complexes in aqueous solution (e.g., TlCl , unlike AgCl is not soluble in aqueous ammonia or potassium cyanide), whereas Tl(III) is very strongly complexed by a variety of anions. Thus, TlCl is fairly soluble, E° for the system $[\text{TlCl}_2]^-/\text{TlCl}$ is only +0.9 V. Iodide forms a more stable complex than chloride (soft acid-soft base relation) and at high iodide concentrations $[\text{TlI}_2]^-$ is a stable species even though $E^\circ_{\text{Tl}^{3+}/\text{Tl}^+}$ is much higher than $E^\circ_{\text{I}_2/\text{I}^-}$ (+0.54V) and TlI is sparingly soluble. Thus, $\text{I}_2(\text{I}^- + \text{I}_2)$ in solid TlI, can under these conditions oxidize Tl(I) and bring about the reaction.



In alkaline media, Tl(I) is also easily oxidized, as TlOH is soluble in water and hydrated $\text{Tl}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$ or $2\text{Tl}(\text{OH})_3$ is very sparingly soluble, with K_{sp} about 10^{-45} .

7.8 ANOMALOUS BEHAVIOUR OF BORON

7.8.1 How Boron Differs from other Elements

As usual, the first element boron of the Group III A shows different properties from the rest of the elements because of the following.

- (i) Small size and high charge make the ion B^{3+} highly polarizing power. Thus, it does not exist. Almost all boron compounds are covalent.
- (ii) Boron does not have *d*-orbitals. Thus, its coordination number is limited to four, whereas the other elements can have a coordination number of six.
- (iii) Boron does not exhibit the "inert pair" effect.
- (iv) Boron combines with metals forming borides whereas other elements do not combine. They form alloys with other metals.
- (v) Boron cannot be attacked by non-oxidizing acids such as HCl whereas others are attacked.
- (vi) Boron does not decompose water or steam whereas other elements of Group III A decompose hot water or steam.
- (vii) Boron is non-metal and bad conductor of electricity but other elements are metals and good conductors of electricity.
- (viii) Boron exhibits allotropy but others elements do not exhibit allotropy.
- (ix) Boron never forms B^{3+} ion, but other elements can form M^{3+} ions.
- (x) Boron forms a large number of volatile hydrides which are electron-deficient compounds whereas other elements form only one polymeric hydride. Thallium does not form hydride.
- (xi) Boron halides are monomeric whereas the halides of the other elements are dimeric.

7.8.2 Similarities between Boron and Aluminium

- (i) **Electronic configuration:** Both boron and aluminium have the same outer electronic configuration $ns^2 np^1$.

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- (ii) **Oxidation state:** Both these elements exhibit +3 oxidation state. However, boron exhibits -3 oxidation state in metal borides.
- (iii) **Covalency:** Both these elements form covalent compounds. However, aluminium may form electrovalent compounds with strong electron accepting groups or atoms.
- (iv) **Formation of oxides:** Both these elements form similar sesquioxide of the type M_2O_3 .
- (v) **Formation of chlorides:** Both these elements form chlorides of the type MCl_3 , when heated in a current of chlorine or by passing chlorine over the heated mixture of their oxides and charcoal.



These chlorides are covalent and readily hydrolysed in water.



- (vi) **Formation of nitrides:** Both the elements when heated with nitrogen or ammonia form nitrides.



These nitrides undergo decomposition when heated with steam or sodium hydroxide liberating ammonia gas.

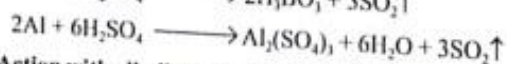


- (vii) **Formation of sulphides:** Both these elements react with sulphur at high temperature to form sulphides which undergo hydrolysis by water.

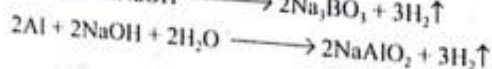


- (viii) **Formation of alkyl compounds:** Both these elements form similar organic compounds with alkyl radicals.

- (ix) **Action with conc. H_2SO_4 :** Both these elements react with concentrated sulphuric acid to form sulphur dioxide



- (x) **Action with alkalis:** They react with alkalis to evolve H_2 .



7.8.3 Dissimilarities of Boron and Aluminium

Boron and aluminium are more different than similar in their properties. This is due to the presence of two electrons in the penultimate shell of boron whereas there are eight electrons in the penultimate shell of aluminium. Boron differs from aluminium in the following respects.

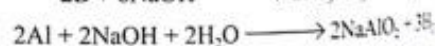
- (i) Boron is a typical non-metal whereas aluminium is a metal.
- (ii) Boron exhibits allotropy whereas aluminium does not exhibit allotropy.
- (iii) Crystalline boron is very hard whereas aluminium is sufficiently soft.
- (iv) Aluminium is very good conductor of heat and electricity whereas boron is a bad conductor.
- (v) Boron has very high melting point as compared with aluminium.
- (vi) The maximum covalence shown by boron is 4 whereas aluminium shows a maximum covalence of 6.
- (vii) Boron is not attacked even by steam whereas aluminium decomposes steam liberating hydrogen.



- (viii) Dilute acids have no action on boron but concentrated HNO_3 oxidizes it to boric acid. Aluminium evolves H_2 gas from dil. HCl and H_2SO_4 but concentrated HNO_3 renders aluminium passive.



- (ix) Boron dissolves in fused alkalis evolving hydrogen whereas aluminium reacts with hot alkali solutions



Borates are very stable as compared to the aluminates

- (x) Boron forms two types (B_nH_{n+4} and B_nH_{n+6}) hydrides whereas aluminium does not form such hydrides.

- (xi) The halides of boron are covalent in nature and are hydrolysed by water giving boric acid. Aluminium chloride in solution gives Al^{3+} ions.

- (xii) Oxide and hydroxide of boron are acidic whereas aluminium are amphoteric.

- (xiii) Boron combines with metals to form borides of Mg_3B_2 but aluminium forms alloys only.

- (xiv) Boron forms many covalent compounds as compared with aluminium.

7.8.4 Resemblance between Boron and Silicon: Diagonal Relationship

Boron shows more resemblance to silicon because of diagonal relationship as shown below.

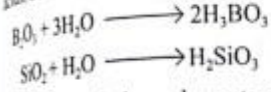
Non-metallic character: Boron and silicon are typical non-metals. They both have high melting points and are bad conductors of electricity.

Allotropy: Both these elements exhibit allotropy (amorphous and crystalline) crystalline forms are

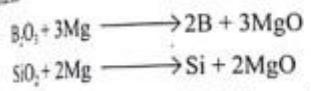
Density, atomic volume and electronegativity: Densities and the atomic volumes of both the elements are low. Their electronegativities are almost similar (B = 2.0; Si = 1.8)

Their ionization energies (kJ mol⁻¹) are also almost similar (B = 801; Si = 786).

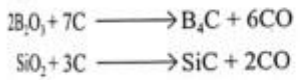
Oxides: Boron and silicon burn in air or oxygen to form stable and acidic oxides B₂O₃, SiO₂. These oxides in association with water yield corresponding acids, boric acid and silicic acid. Both are weak acids.



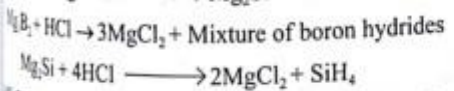
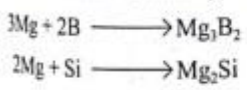
Extraction: Both these elements can be obtained by the reduction of their respective oxides with magnesium.



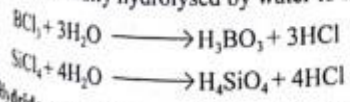
Carbides: The oxides of both these elements when fused with carbon form carbides; B₄C and SiC. These are very hard substances and are used as abrasives.



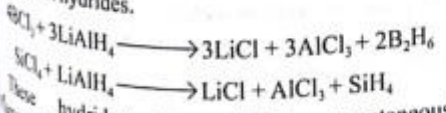
Reaction with metals: Both combine with metals to form borides and silicides, which are decomposed by dilute acids to form volatile hydrides.



Reaction with halogens: Both boron and silicon form halides with halogens. Fluorides of both are colourless fuming gases. Chlorides BCl₃ and SiCl₄ are liquids, which are readily hydrolysed by water to acids.

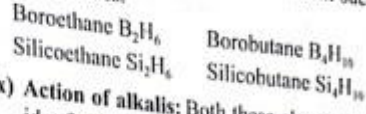


Hydrides: Both boron and silicon form a number of covalent hydrides.

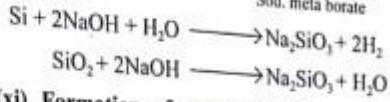
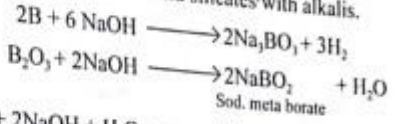


These hydrides are volatile, spontaneously inflammable and readily hydrolysed. These hydrides are known as boranes and silicoalkanes or silanes. Boron

appears to act as a tetra-covalent element such as silicon in these hydrides.



(x) **Action of alkalis:** Both these elements and their oxides form borates and silicates with alkalis.



(xi) **Formation of esters:** Both these elements form volatile esters of the type B(OR)₃ and Si(OR)₄ with alcohols.

