1. Details of Module and its structure

Module Detail	
Subject Name	Chemistry
Course Name	Chemistry 02 (Class XI – Semester 2)
Module Name/Title	The p- block elements: The Boron family – Part 1
Module Id	Kech_21101
Pre-requisites	Knowledge of the concept of periodic table, periodic properties of elements in the periodic table and atomic physical and chemical properties of the *s*- block elements.
Objectives	After going through this lesson, the learners will be able to: 1. Understand the atomic, physical and chemical properties of the group 13 elements. 2. Describe the chemistry of some important compounds of boron. 3. Understand the important application and uses of boron, aluminium and their compounds
Keywords	*p*- block elements, Inert pair effect, Anomalous properties of Boron, Borax, Orthoboric acid Diborane

2. Development Team

Role	Name	Affiliation
National MOOC Coordinator (NMC)	Prof. Amarendra P. Behera	CIET, NCERT, New Delhi
Program Coordinator	Dr. Mohd. Mamur Ali	CIET, NCERT, New Delhi
Course Coordinator (CC) / PI	Prof. R. K. Parashar	DESM, NCERT, New Delhi
Course Co-Coordinator / Co-PI	Dr. Deepty Gupta	CIET, NCERT, New Delhi
Subject Matter Expert (SME)	Dr. Amit Kumar	Dyal Singh College, University of Delhi
Review Team	Dr. Sulekha Chandra	Zakir Husain Delhi College, University of Delhi
	Dr Sanjeev Kumar Mishra	Zakir Husain Delhi College, University of Delhi

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1. **Introduction**

In *p*-block elements the last electron enters the outermost *p* orbital. As we know that the number of *p* orbitals is three and, therefore, the maximum number of electrons that can be accommodated in a set of *p* orbitals is six. Consequently there are six groups of *p*–block elements in the periodic table numbering from 13 to 18. Boron, carbon, nitrogen, oxygen, fluorine and helium head the groups. Their valence shell electronic configuration is ns^2np^{1-6} (except for He). The inner core of the electronic configuration may, however, differ. The difference in inner core of elements greatly influences their physical properties (such as atomic and ionic radii, ionisation enthalpy, etc.) as well as chemical properties. Consequently, a lot of variation in properties of elements in a group of *p*-block is observed. The maximum oxidation state shown by a *p*-block element is equal to the total number of valence electrons (i.e., the sum of the *s*- and *p*-electrons). Clearly, the number of possible oxidation states increases towards the right of the periodic table. In addition to this so called group oxidation state, *p*-block elements may show other oxidation states which normally, but not necessarily,

differ from the total number of valence electrons by unit of two. The important oxidation states exhibited by *p*-block elements are shown in Table 1.

Table 1: General Electronic Configuration and Oxidation States of p-Block Elements

Group	13	14	15	16	17	18
General electronic configuration	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6 ($1s^2$ for He)
First member of the group	B	C	N	O	F	He
Group oxidation state	+3	+4	+5	+6	+7	+8
Other oxidation states	+1	+2, − 4	+3, − 3	+4, +2, −2	+5, + 3, +1, −1	+6, +4, +2

In boron, carbon and nitrogen families the group oxidation state is the most stable state for the lighter elements in the group. However, the oxidation state two unit less than the group oxidation state becomes progressively more stable for the heavier elements in each group. The occurrence of oxidation states two unit less than the group oxidation states are sometime attributed to the 'inert pair effect'.

The relative stabilities of these two oxidation state – group oxidation state and two unit less than the group oxidation state – may vary from group to group and will be discussed at appropriate places. It is interesting to note that the non-metals and metalloids exist only in the p-block of the periodic table. The non-metallic character of elements decreases down the group. In fact the heaviest element in each *p*-block group is the most metallic in nature. This change from nonmetallic to metallic character brings diversity in the chemistry of these elements depending on the group to which they belong. In general, non-metals have higher ionisation enthalpies and higher electronegativities than the metals. Hence, in contrast to metals which readily form cations, non-metals readily form anions. The compounds formed by highly reactive non-metals with highly reactive metals are generally ionic because of large differences in their electronegativities. On the other hand, compounds formed between non-metals themselves are largely covalent in character because of small differences in their electronegativities. The change of non-metallic to metallic character can be best illustrated by the nature of oxides they form. The non-metal oxides are acidic or neutral whereas metal oxides are basic in nature. The first member of *p*-block differs from the remaining members of their corresponding group in two major respects. First is the size and all other properties which depend on size. Thus, the lightest *p*-block elements show the same kind of differences

as the lightest *s*-block elements, lithium and beryllium. The second important difference, which applies only to the *p*-block elements, arises from the effect of *d*- orbitals in the valence shell of heavier elements (starting from the third period onwards) and their lack in second period elements. The second period elements of *p*-groups starting from boron are restricted to a maximum covalence of four (using 2*s* and three 2*p* orbitals). In contrast, the third period elements of *p*-groups with the electronic configuration $3s^2 3p^n$ have the vacant 3*d* orbitals lying between the 3*p* and the 4*s* levels of energy. Using these *d*-orbitals the third period elements can expand their covalence above four. For example, while boron forms only $[BF_4]^-$, aluminium gives $[AlF_6]^{3-}$ ion. The presence of these *d*-orbitals influences the chemistry of the heavier elements in a number of other ways. The combined effect of size and availability of *d*- orbitals considerably influences the ability of these elements to form π bonds. The first member of a group differs from the heavier members in its ability to form *p*π–*p*π multiple bonds to itself (e.g., C=C, C≡C, N≡N) and to other second row elements (e.g., C=O, C=N, C≡N, N=O). This type of π- bonding is not particularly strong for the heavier p-block elements. The heavier elements do form π bonds but this involves *d* orbitals (*d*π–*p*π or *d*π–*d*π). As the d orbitals are of higher energy than the *p* orbitals, they contribute less to the overall stability of molecules than does *p*π–*p*π bonding of the second row elements. However, the coordination number in species of heavier elements may be higher than for the first element in the same oxidation state. For example, in +5 oxidation state both N and P form oxoanions : NO_3^- (three-coordination with π– bond involving one nitrogen *p*- orbital) and PO_4^{3-} (four-coordination involving *s*, *p* and *d* orbitals contributing to the π – bond). In this module we will study the chemistry of group 13 elements of the periodic table: The Boron family.

2.   Group 13 elements: The Boron Family

This group elements show a wide variation in properties. Boron is a typical non-metal, aluminium is a metal but shows many chemical similarities to boron, and gallium, indium, thallium and nihonium are almost exclusively metallic in character. Boron is a fairly rare element, mainly occurs as orthoboric acid, (H_3BO_3), borax, $Na_2B_4O_7.10H_2O$, and kernite, $Na_2B_4O_7.4H_2O$. In India borax occurs in Puga Valley (Ladakh) and Sambhar Lake (Rajasthan). The abundance of boron in earth crust is less than 0.0001% by mass. There are two isotopic forms of boron ^{10}B (19%) and ^{11}B (81%). Aluminium is the most abundant metal and the third most abundant element in the earth's crust (8.3% by mass) after oxygen (45.5%) and Si (27.7%). Bauxite, $Al_2O_3.2H_2O$ and cryolite, Na_3AlF_6 are the important minerals of aluminium. In India it is found as mica in Madhya Pradesh, Karnataka, Orissa and Jammu.

Gallium, indium and thallium are less abundant elements in nature. Nihonium has symbol Nh, atomic number 113, atomic mass 286 g mol^{-1} and electronic configuration [Rn] $5f^{14}6d^{10}7s^27p^2$. So far it has been prepared in small amount and half life of its most stable isotope is 20 seconds. Due to these reasons its chemistry has not been established. Nihonium is a synthetically prepared radioactive element. Here atomic, physical and chemical properties of elements of this group leaving nihonium are discussed below.

2.1. *Electronic Configuration*

The outer electronic configuration of these elements is ns^2np^1. A close look at the electronic configuration suggests that while boron and aluminium have noble gas core, gallium and indium have noble gas plus 10 d- electrons, and thallium has noble gas plus 14 f- electrons plus 10 d-electron cores. Thus, the electronic structures of these elements are more complex than for the first two groups of elements discussed in module on s- block elements. This difference in electronic structures affects the other properties and consequently the chemistry of all the elements of this group.

2.2. *Atomic Radii*

On moving down the group, for each successive member one extra shell of electrons is added and, therefore, atomic radius is expected to increase. However, a deviation can be seen. Atomic radius of Ga is less than that of Al. This can be understood from the variation in the inner core of the electronic configuration. The presence of additional 10 d- electrons offer only poor screening effect for the outer electrons from the increased nuclear charge in gallium, consequently, the atomic radius of gallium (135 pm) is less than that of aluminium (143 pm).

2.3. *Ionization Enthalpy*

The ionisation enthalpy values as expected from the general trends do not decrease smoothly down the group. The decrease from B to Al is associated with increase in size. The observed discontinuity in the ionisation enthalpy values between Al and Ga, and between In and Tl are due to inability of d- and f- electrons, which have low screening effect, to compensate the increase in nuclear charge. The order of ionisation enthalpies, as expected, is $\Delta_iH_1 < \Delta_iH_2 < \Delta_iH_3$. The sum of the first three ionisation enthalpies for each of the elements is very high. Effect of this will be apparent when you study their chemical properties.

2.4. *Electronegativity*

Down the group, electronegativity first decreases from B to Al and then increases marginally (Table 2). This is because of the discrepancies in atomic size of the elements.

2.5. *Physical Properties*

Boron is non-metallic in nature. It is extremely hard and black coloured solid. It exists in many allotropic forms. Due to very strong crystalline lattice, boron has unusually high melting point. Rest of the members are soft metals with low melting point and high electrical conductivity. It is worthwhile to note that gallium with unusually low melting point (303K), could exist in liquid state during summer. Its high boiling point (2676 K) makes it a useful material for measuring high temperatures. Density of the elements increases down the group from boron to thallium.

Table 2: Atomic and Physical Properties of Group 13 Elements

| Property		Element				
		Boron B	Aluminium Al	Gallium Ga	Indium In	Thallium Tl
Atomic number		5	13	31	49	81
Atomic mass(g mol^{-1})		10.81	26.98	69.72	114.82	204.38
Electronic Configuration		[He]$2s^22p^1$	[Ne]$3s^23p^1$	[Ar]$3d^{10}4s^24p^1$	[Kr]$4d^{10}5s^25p^1$	[Xe]$4f^{14}5d^{10}6s^26p^1$
Atomic radius/pm^a		(88)	143	135	167	170
Ionic radius M^{3+}/pm^b		(27)	53.5	62.0	80.0	88.5
Ionic radius M$^+$/pm		-	-	120	140	150
Ionization enthalpy (kJ mol^{-1})	Δ_iH_1	801	577	579	558	589
	Δ_iH_2	2427	1816	1979	1820	1971
	Δ_iH_3	3659	2744	2962	2704	2877
Electronegativity^c		2.0	1.5	1.6	1.7	1.8
Density /g cm^{-3} at 298 K		2.35	2.70	5.90	7.31	11.85
Melting point / K		2453	933	303	430	576
Boiling point / K		3923	2740	2676	2353	1730
E$^\ominus$/ V for (M^{3+}/M)		-	−1.66	−0.56	−0.34	+1.26
E$^\ominus$ / V for (M$^+$/M)		-	+0.55	-0.79(acid) −1.39(alkali)	−0.18	−0.34

^{a}*Metallic radius,* ^{b} *6-coordination,* ^{c} *Pauling scale,*

2.6. *Chemical Properties*

Oxidation state and trends in chemical reactivity

Due to small size of boron, the sum of its first three ionization enthalpies is very high. This prevents it to form +3 ions and forces it to form only covalent compounds. But as we move from B to Al, the sum of the first three ionisation enthalpies of Al considerably decreases, and is therefore able to form Al^{3+} ions. In fact, aluminium is a highly electropositive metal.

However, down the group, due to poor shielding effect of intervening d and f orbitals, the increased effective nuclear charge holds ns electrons tightly (responsible for inert pair effect) and thereby, restricting their participation in bonding. As a result of this, only p- orbital electron may be involved in bonding. In fact in Ga, In and Tl, both +1 and +3 oxidation states are observed. The relative stability of +1 oxidation state progressively increases for heavier elements: Al<Ga<In whereas the +3 oxidation state is highly oxidising in character. The compounds in +1 oxidation state, as expected from energy considerations, are more ionic than those in +3 oxidation state. In trivalent state, the number of electrons around the central atom in a molecule of the compounds of these elements (e.g., boron in BF_3) will be only six. Such electron deficient molecules have tendency to accept a pair of electrons to achieve stable electronic configuration and thus, behave as Lewis acids. The tendency to behave as Lewis acid decreases with the increase in the size down the group. BCl_3 easily accepts a lone pair of electrons from ammonia to form $Bcl_3 \leftarrow NH_3$.

Planar   Tetrahedral
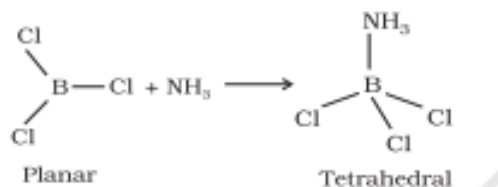
AlCl₃ achieves stability by forming a dimer
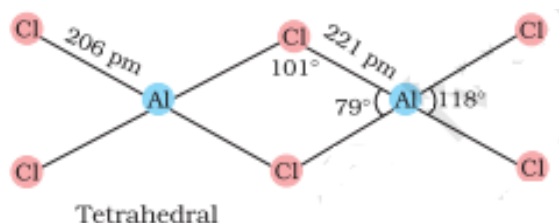
Tetrahedral

In trivalent state most of the compounds being covalent are hydrolysed in water. For example, the trichlorides on hyrolysis in water form tetrahedral $[M(OH)_4]^-$ species; the hybridisation state of element M is sp^3. Aluminium chloride in acidified aqueous solution forms octahedral $[Al(H_2O)_6]^{3+}$ ion. In this complex ion, the $3d$ orbitals of Al are involved and the hybridisation state of Al is sp^3d^2.

Problem 1: Standard electrode potential values, $E^{\ominus}$ for Al^{3+}/Al is –1.66 V and that of Tl^{3+}/Tl is +1.26 V. Predict about the formation of M^{3+} ion in solution and compare the electropositive character of the two metals.

Solution: Standard electrode potential values for two half cell reactions suggest that aluminium has high tendency to make Al^{3+}(aq) ions, whereas Tl^{3+} is not only unstable in solution but is a powerful oxidising agent also. Thus Tl^+ is more stable in solution than Tl^{3+}. Aluminium being able to form +3 ions easily, is more electropositive than thallium.

(i) Reactivity towards Air

Boron is unreactive in crystalline form. Aluminium forms a very thin oxide layer on the surface which protects the metal from further attack. Amorphous boron and aluminium metal on heating in air form B_2O_3 and Al_2O_3 respectively. With dinitrogen at high temperature they form nitrides.

$$2E(s) + 3O_2(g) \xrightarrow{\Delta} 2E_2O_3(s)$$

$$2E(s) + N_2(g) \xrightarrow{\Delta} 2EN(s)$$
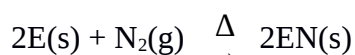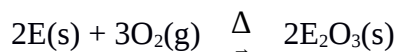
(E = element)

The nature of these oxides varies down the group. Boron trioxide is acidic and reacts with basic (metallic) oxides forming metal borates. Aluminium and gallium oxides are amphoteric and those of indium and thallium are basic in their properties.

(ii) *Reactivity towards Acids and Alkalies*

Boron does not react with acids and alkalies even at moderate temperature; but aluminium dissolves in mineral acids and aqueous alkalies and thus shows amphoteric character. Aluminium dissolves in dilute HCl and liberates dihydrogen.

$$2Al(s) + 6HCl\ (aq) \rightarrow 2Al^{3+}\ (aq) + 6Cl^-\ (aq) + 3H_2\ (g)$$
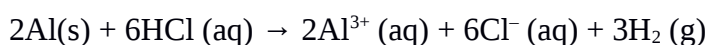
However, concentrated nitric acid renders aluminium passive by forming a protective oxide layer on the surface.

Aluminium also reacts with aqueous alkali and liberates dihydrogen.

$$2Al(s) + 2NaOH(aq) + 6H_2O(l) \rightarrow 2Na^+[AlH_4]^-(aq) + 3H_2(g)$$
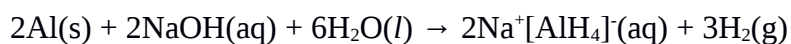
Sodium tetrahydroxoaluminate(III)

(iii) *Reactivity towards Halogens*

These elements react with halogens to form trihalides (except TlI_3).

$$2E(s) + 3\ X_2\ (g) \rightarrow 2EX_3\ (s) \qquad (X = F, Cl, Br, I)$$
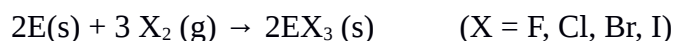
Problem 2: White fumes appear around the bottle of anhydrous aluminium chloride. Give reason.

Solution: Anhydrous aluminium chloride is partially hydrolysed with atmospheric moisture to liberate HCl gas. Moist HCl appears white in colour.

3. **Important Trends and Anomalous Properties of Boron**

Certain important trends can be observed in the chemical behaviour of group 13 elements. The tri-chlorides, bromides and iodides of all these elements being covalent in nature are

hydrolysed in water. Species like tetrahedral $[M(OH)_4]^-$ and octahedral $[M(H_2O)_6]^{3+}$, except in boron, exist in aqueous medium. The monomeric trihalides, being electron deficient, are strong Lewis acids. Boron trifluoride easily reacts with Lewis bases such as NH_3 to complete octet around boron.

$F_3B + :NH_3 \rightarrow F_3B \leftarrow NH_3$

It is due to the absence of d orbitals that the maximum covalence of B is 4. Since the d orbitals are available with Al and other elements, the maximum covalence can be expected beyond 4. Most of the other metal halides (e.g., $AlCl_3$) are dimerised through halogen bridging (e.g., Al_2Cl_6). The metal species completes its octet by accepting electrons from halogen in these halogen bridged molecules.

Problem 3: Boron is unable to form BF_6^{3-} ion. Explain.

Solution Due to non-availability of d orbitals, boron is unable to expand its octet. Therefore, the maximum covalence of boron cannot exceed 4.

4. *Some Important Compounds of Boron*

Some useful compounds of boron are borax, orthoboric acid and diborane. We will briefly study their chemistry.

4.1. *Borax*

It is the most important compound of boron. It is a white crystalline solid of formula $Na_2B_4O_7 \cdot 10H_2O$. In fact it contains the tetranuclear units $[B_4O_5(OH)_4]^{2-}$ and correct formula; therefore, is $Na_2[B_4O_5(OH)_4].8H_2O$. Borax dissolves in water to give an alkaline solution.

$Na_2B_4O_7 + 7H_2O \rightarrow 2NaOH + 4H_3BO_3$
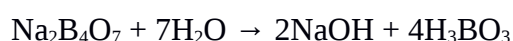
Orthoboric acid

On heating, borax first loses water molecules and swells up. On further heating it turns into a transparent liquid, which solidifies into glass like material known as borax bead.

$Na_2B_4O_7.10H_2O \xrightarrow{\Delta} Na_2B_4O_7 \xrightarrow{\Delta} 2NaBO_2 + B_2O_3$
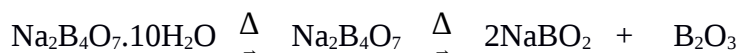
Sodium metaborate   Boric anhydride

The metaborates of many transition metals have characteristic colours and, therefore, borax bead test can be used to identify them in the laboratory. For example, when borax is heated in a Bunsen burner flame with CoO on a loop of platinum wire, a blue coloured $Co(BO_2)_2$ bead is formed.

4.2. *Orthoboric acid*

Orthoboric acid, H_3BO_3 is a white crystalline solid, with soapy touch. It is sparingly soluble in water but highly soluble in hot water. It can be prepared by acidifying an aqueous solution of borax.

$$Na_2B_4O_7 + 2HCl + 5H_2O \rightarrow 2NaCl + 4B(OH)_3$$
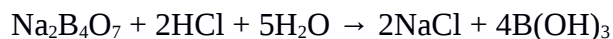
It is also formed by the hydrolysis (reaction with water or dilute acid) of most boron compounds (halides, hydrides, etc.). It has a layer structure in which planar BO_3 units are joined by hydrogen bonds as shown in Fig. 1. Boric acid is a weak monobasic acid. It is not a protonic acid but acts as a Lewis acid by accepting electrons from a hydroxyl ion:

$$B(OH)_3 + 2HOH \rightarrow [B(OH)_4]^- + H_3O^+$$

On heating, orthoboric acid above 370K forms metaboric acid, HBO_2 which on further heating yields boric oxide, B_2O_3.

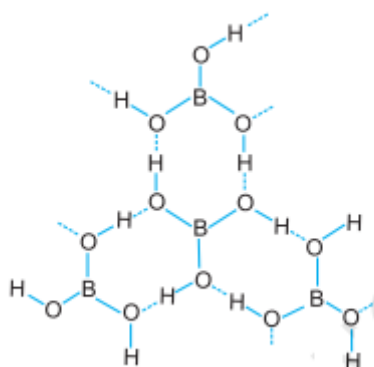$$H_3BO_3 \xrightarrow{\Delta} HBO_2 \xrightarrow{\Delta} B_2O_3$$

Fig. 1: Structure of boric acid; the dotted lines represent hydrogen bonds

Problem 4: Why is boric acid considered as a weak acid?

Solution: Because it is not able to release H^+ ions on its own. It receives OH^- ions from water molecule to complete its octet and in turn releases H^+ ions.

4.3. *Diborane, B_2H_6*

The simplest boron hydride known is diborane. It is prepared by treating boron trifluoride with $LiAlH_4$ in diethyl ether.

$$4BF_3 + 3\ LiAlH_4 \rightarrow 2B_2H_6 + 3LiF + 3AlF_3$$
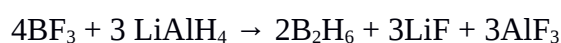
A convenient laboratory method for the preparation of diborane involves the oxidation of sodium borohydride with iodine.

$$2NaBH_4 + I_2 \rightarrow B_2H_6 + 2NaI + H_2$$

Diborane is produced on an industrial scale by the reaction of BF_3 with sodium hydride.

$$2BF_3 + 6\ NaH \xrightarrow{450\ K} B_2H_6 +\ 6\ NaF$$
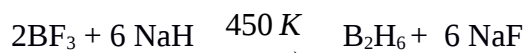
Diborane is a colourless, highly toxic gas with a b.p. of 180 K. Diborane catches fire spontaneously upon exposure to air. It burns in oxygen releasing an enormous amount of energy.

$$B_2H_6 + 3O_2 \rightarrow B_2O_3 + 3\ H_2O \qquad \Delta_c H^\ominus = -\ 1976\ kJ\ mol^{-1}$$
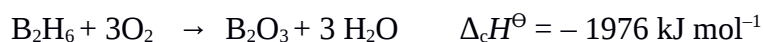
Most of the higher boranes are also spontaneously flammable in air. Boranes are readily hydrolysed by water to give boric acid.

$$B_2H_6\ (g) + 6H_2O(l) \rightarrow 2B(OH)_3\ (aq) + 6H_2\ (g)$$
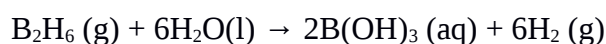
Diborane undergoes cleavage reactions with Lewis bases(L) to give borane adducts, $BH3 \cdot L$

$$B_2H_6 + 2\ NMe_3 \rightarrow 2BH_3{:}NMe_3$$

$$B_2H_6 + 2\ CO \rightarrow 2BH_3{:}CO$$
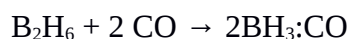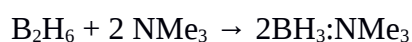
Reaction of ammonia with diborane gives initially $B_2H_6.2NH_3$ which is formulated as $[BH_2(NH_3)_2]^+[BH_4]^-$; further heating gives borazine, $B_3N_3H_6$ known as "inorganic benzene" in view of its ring structure with alternate BH and NH groups

$$3B_2H_6 + 6NH_3 \rightarrow 3\ [BH_2(NH_3)_2]^+[BH_4]^- \xrightarrow{Heat} 2B_3N_3H_6 + 12\ H_2$$
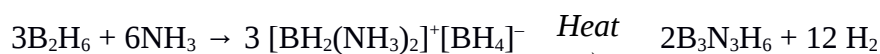
The structure of diborane is shown in Fig. 2(a). The four terminal hydrogen atoms and the two boron atoms lie in one plane. Above and below this plane, there are two bridging hydrogen atoms. The four terminal B-H bonds are regular two centre-two electron bonds while the two bridge (B-H-B) bonds are different and can be described in terms of three centre–two electron bonds shown in Fig. 2 (b).
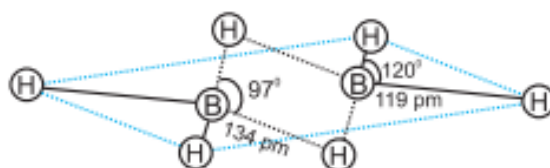

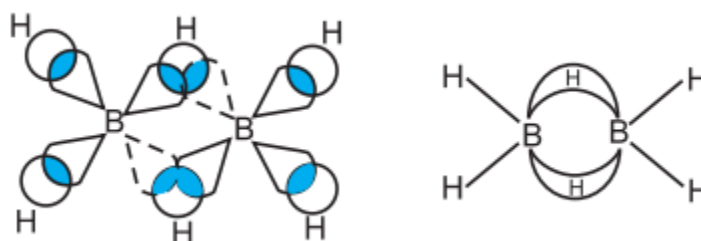Fig. 2(a) The structure of diborane, B_2H_6

Fig. 2(b:) Bonding in diborane. Each B atom uses sp^3 hybrids for bonding. Out of the four sp^3 hybrids on each B atom, one is without an electron shown in broken lines. The terminal B-H bonds are normal 2-centre-2- electron bonds but the two bridge bonds are 3-centre-2-electron bonds. The 3-centre-2-electron bridge bonds are also referred to as banana bonds.

Boron also forms a series of hydridoborates; the most important one is the tetrahedral $[BH_4]^-$ ion. Tetrahydridoborates of several metals are known. Lithium and sodium tetrahydridoborates, also known as borohydrides, are prepared by the reaction of metal hydrides with B_2H_6 in diethyl ether.

$2MH + B_2H_6 \rightarrow 2\,M^+\,[BH_4]^-$ (M = Li or Na)

Both $LiBH_4$ and $NaBH_4$ are used as reducing agents in organic synthesis. They are useful starting materials for preparing other metal borohydrides.

5.   Uses of Boron and Aluminium and their Compounds

Boron being extremely hard refractory solid of high melting point, low density and very low electrical conductivity, finds many applications. Boron fibres are used in making bullet-proof vest and light composite material for aircraft. The boron-10 (^{10}B) isotope has high ability to absorb neutrons and, therefore, metal borides are used in nuclear industry as protective shields and control rods. The main industrial application of borax and boric acid is in the manufacture of heat resistant glasses (e.g., Pyrex), glass-wool and fibreglass. Borax is also used as a flux for soldering metals, for heat, scratch and stain resistant glazed coating to earthenwares and as constituent of medicinal soaps. An aqueous solution of orthoboric acid is generally used as a mild antiseptic. Aluminium is a bright silvery-white metal, with high tensile strength. It has a high electrical and thermal conductivity. On a weight-to-weight basis, the electrical conductivity of aluminium is twice that of copper. Aluminium is used extensively in industry and everyday life. It forms alloys with Cu, Mn, Mg, Si and Zn. Aluminium and its alloys can be given shapes of pipe, tubes, rods, wires, plates or foils and, therefore, find uses in packing, utensil making, construction, aeroplane and transportation industry. The use of aluminium and its compounds for domestic purposes is now reduced considerably because of their toxic nature.

Summary

p-Block of the periodic table is unique in terms of having all types of elements – metals, non-metals and metalloids. There are six groups of p-block elements in the periodic table numbering from 13 to 18. Their valence shell electronic configuration is $ns^2\,np^{1-6}$ (except for

He). Differences in the inner core of their electronic configuration greatly influence their physical and chemical properties. As a consequence of this, a lot of variation in properties among these elements is observed. In addition to the group oxidation state, these elements show other oxidation states differing from the total number of valence electrons by unit of two. While the group oxidation state is the most stable for the lighter elements of the group, lower oxidation states become progressively more stable for the heavier elements. The combined effect of size and availability of d orbitals considerably influences the ability of these elements to form π-bonds. While the lighter elements form pπ–pπ bonds, the heavier ones form dπ–pπ or dπ–dπ bonds. Absence of d orbital in second period elements limits their maximum covalence to 4 while heavier ones can exceed this limit.

Boron is a typical non-metal and the other members are metals. The availability of 3 valence electrons ($2s^2\ 2p^1$) for covalent bond formation using four orbitals (2s, 2px, 2py and 2pz) leads to the so called electron deficiency in boron compounds. This deficiency makes them good electron acceptor and thus boron compounds behave as Lewis acids. Boron forms covalent molecular compounds with dihydrogen as boranes, the simplest of which is diborane, B_2H_6 . Diborane contains two bridging hydrogen atoms between two boron atoms; these bridge bonds are considered to be three-centre two-electron bonds. The important compounds of boron with dioxygen are boric acid and borax. Boric acid, $B(OH)_3$ is a weak monobasic acid; it acts as a Lewis acid by accepting electrons from hydroxyl ion. Borax is a white crystalline solid of formula $Na_2\ [B_4O_5\ (OH)_4\ ]{\cdot}8H_2O$. The borax bead test gives characteristic colours of transition metals.

Aluminium exhibits +3 oxidation state. With heavier elements +1 oxidation state gets progressively stabilised on going down the group. This is a consequence of the so called inert pair effect.