The p-Block Elements

The elements in which the last electron enters in the valence p-sub shell are called the p-block elements. They include elements of the groups 13 to 18. Their general outer electronic configuration is ns^2np^{1-6} (except He which has $1s^2$ configuration).They includes metals, non-metals and metalloids.

Group 15 Elements

Group 15 includes nitrogen (N), phosphorus (P), arsenic (As), antimony (Sb) and bismuth (Bi). As we go down the group, the metallic character increases. Nitrogen and phosphorus are non-metals, arsenic and antimony metalloids and bismuth is a typical metal. The valence shell electronic configuration of these elements is ns^2np^3. The *s* orbital in these elements is completely filled and *p* orbitals are half-filled, making their electronic configuration extra stable.

Covalent and ionic radii increase down the group. There is a considerable increase in covalent radius from N to P. However, from As to Bi only a small increase in covalent radius is observed. This is due to the presence of completely filled *d* or *f* orbitals in heavier members. Ionisation enthalpy decreases down the group due to gradual increase in atomic size. Because of the extra stable half-filled *p* orbitals and smaller size, the ionisation enthalpy of the group 15 elements is much greater than that of group 14 elements.

Oxidation states and trends in chemical reactivity

The common oxidation states of these elements are –3, +3 and +5. The tendencies to exhibit –3 oxidation state decreases down the group due to increase in size and metallic character. The last member of the group, bismuth does not form any compound in –3 oxidation state. The stability of +5 oxidation state decreases and that of +3 state increases (due to inert pair effect) down the group. Nitrogen exhibits + 1, + 2, and + 4 oxidation states also when it reacts with oxygen. Phosphorus also shows +1 and +4 oxidation states in some oxoacids. Nitrogen is restricted to a maximum covalency of 4 since only four orbitals (one *s* and three *p*) are available for bonding.

Oxidation States

Elements:-	**N**	**P**	**As**	**Sb**	**Bi**
Atomic No.:-	**7**	**15**	**33**	**51**	**83**
	-3	**-3**	**-3**	**-3**	**--**
	+5	**+5**	**+5**	**+5**	**+5**
	+3	**+3**	**+3**	**+3**	**+ 3**
	+4	**+4**			
	+1	**+1**			

Notes:-1. Due to non-metallic (N and P) and metalloidic (As and Sb)nature N, P, As, Sb show negative oxidation state (-3)

2. Due to anomalous property "N" shows variables oxidation states.

3. Due to presence of vacant d- sub shell, non-metallic nature and more electronegative value "P" shows variables oxidation states.

4. Due to inert pair effect "As, Sb and Bi" show variables oxidation states.

Anomalous properties of nitrogen

Nitrogen differs from the rest of the members of this group due to its smaller size, high electro negativity, high ionisation enthalpy and non-availability of *d* orbitals. Some of the anomalous properties shown by nitrogen are:

1. Nitrogen has the ability to form ***p*π-*p*π multiple** bonds with itself and with other elements like C and O. Other elements of this group do not form *p*π-*p*π bonds.
2. Nitrogen exists as a diatomic molecule with a triple bond (one *s* and two *p*) between the two atoms. So its bond enthalpy is very high. While other elements of this group are poly atomic with single bonds.
3. The single N–N bond is weak. So the catenation tendency is weaker in nitrogen.
4. Due to the absence of *d* orbitals in its valence shell, the maximum covalency of nitrogen is four
5. N cannot form ***d*π–*p*π bond.** While Phosphorus and arsenic can form ***d*π–*d*π bond** with transition metals and with C and O.

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