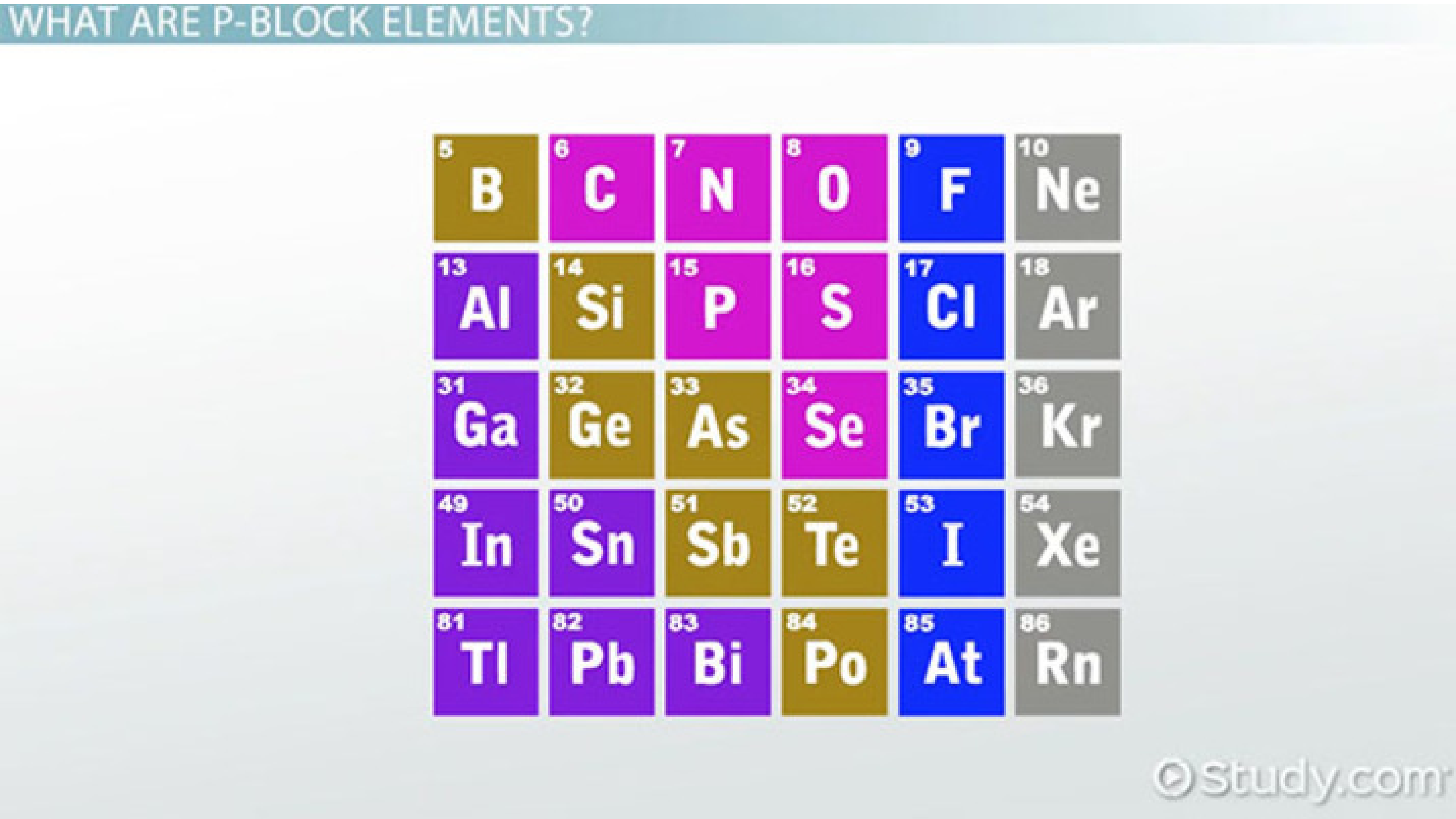
WHAT ARE P-BLOCK ELEMENTS?
5 B
6 C
7 N
8 O
9 F
10 Ne
13 Al
14 Si
15 P
16 S
17 Cl
18 Ar
31 Ga
32 Ge
33 As
34 Se
35 Br
36 Kr
49 In
50 Sn
51 Sb
52 Te
53 I
54 Xe
81 Tl
82 Pb
83 Bi
84 Po
85 At
86 Rn
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Characteristics of p-Block Elements

1. Size of atoms or ions. Size of an atom of *p*-block elements decreases on moving from left to right in a period while the same increases on descending a group (see Table 3.10). As already said, the size of an ion is expressed as ionic radius. Ionic radii, like atomic radii, also decrease on moving from left to right in a period and increase on moving down the group (See Table 3.10).

Table 3.10. Covalent radii (in Å) of p-block elements. The values shown for the positive or negative oxidation states given in parentheses are ionic radii.

Group→ / Period ↓	III A	IV A	V A	VI A	VII A	Zero
2	B 0.82 0.20 (+3) —	C 0.77 0.15 (+4) 2.60 (–4)	N 0.75 0.11 (+5) 1.71 (–3)	O 0.73 0.09 (+6) 1.40 (–2)	F 0.72 0.07 (+7) 1.36 (–1)	Ne 0.71 — —
3	Al 1.18 0.50 (+3) —	Si 1.11 0.41 (+4) 2.71 (–1)	P 1.06 0.34 (+5) 2.12 (–3)	S 1.02 0.29 (+6) 1.84 (–2)	Cl 0.99 0.26 (+7) 1.81 (–1)	Ar 0.98 — —
4	Ga 1.26 0.62 (+3) 1.13 (+1)	Ge 1.22 0.53 (+4) 0.93 (+2)	As 1.20 0.47 (+5) 2.22 (–3)	Se 1.16 0.42 (+6) 1.98 (–2)	Br 1.14 0.39 (+7) 1.95 (–1)	Kr 1.12 — —
5	In 1.44 0.81 (+3) 1.32 (+1)	Sr 1.41 0.71 (+4) 1.12 (+2)	Sb 1.40 0.62 (+5) 2.45 (–3)	Te 1.36 0.56 (+6) 2.21 (–2)	I 1.33 0.50 (+7) 2.16 (–1)	Xe 1.31 — —
6	Tl 1.48 0.95 (+3) 1.40 (+1)	Pb 1.47 0.84 (+4) 1.20 (+2)	Bi 1.46 0.74 (+5) 1.20 (+3)	Po 1.46 — —	At 1.45 — —	Rn — — —

2. Ionisation energies. First ionisation energy of *p*-block elements are given in Table 3.11 from which it can be seen that these values generally increase on moving from left to right in a period and decrease on descending the group.

The increase in the period is not uniform *i.e.* there are certain elements which show irregular trends. For example, in 2*nd* period Be and N have higher ionisation energy values than their next neighbours namely B and O. Similarly in 3*rd* period Mg and P have higher values than Al and S A detailed explanation

Table 3.11. First ionisation energies (in KJ/mole) of p-block elements.

Group→ / Period ↓	III A	IV A	V A	VI A	VII A	Zero
2	B 800.6	C 1086.4	N 1402.3	O 1314.0	F 1681.0	Ne 2080.7
3	Al 577.6	Si 786.5	P 1011.8	S 999.6	Cl 1251.1	Ar 1520.5
4	Ga 578.8	Ge 762.2	As 944.0	Se 940.9	Br 1139.9	Kr 1350.7
5	In 558.3	Sn 708.6	Sb 831.6	Te 869.3	I 1008.4	Xe 1170.4
6	Tl 589.3	Pb 715.5	Bi 703.3	Po 812.0	At —	Rn 1037.0

for this abnormal behaviour shown by Be, N, Mg and P has been given in the chapter on *Periodic Properties*. (Chapter No 4.)

3. Electropositive (or metallic or basic) character of the elements and the nature of their oxides. Electropositive character of an element depends on the magnitude of its ionisation energy. With the increase of ionisation energy, the metallic character of the elements decreases. Thus :

(*i*) Since the ionisation energy of *p*-block elements increases on moving from left to right in a period, the metallic character of the elements decreases. For example the elements lying at the extreme right of each period (*i.e.* halogens) show minimum electropositive character and hence are non-metals.

(*ii*) Since the ionisation energy of *p*-block elements decreases on descending a group, the metallic character of these elements increases in the same direction. The successive increase of metallic character of the elements on descending a group is particularly apparent in groups IV A, V A and VI A which begin with non-metals namely C, N and O and end with metals namely Pb, Bi and Po respectively.

The metallic elements of *p*-block give basic oxides (*e.g.* PbO_2, Bi_2O_3), metalloids give amphoteric oxides (*e.g.* GeO_2, As_2O_3, Sb_2O_3) while non metallic elements give acidic oxides (*e.g.* CO_2, P_4O_{10}, SO_3, Cl_2O_7). The variation of acidic/basic nature of these oxides in the periodic table depends on the metallic (electropositive) character of the element forming the oxide. With the decrease of the metallic character of the elements on moving across a period, the oxides of the elements of the same period become more and more acidic in the same direction. For example the oxides of the elements of *3rd* period become more and more acidic when we proceed from Al_2O_3 to Cl_2O_7.

Group :	III A	IVA	V A	VI A	VII A
Oxides :	Al_2O_3 Amphoteric	SiO_2 Acidic	P_4O_{10} More acidic	SO_3 More acidic	$C1_2O_7$ Most acidic

Asidic nature : —*Acidic nature increasing* →

With the increase of the metallic character of the elements, on descending a group, the oxides of the elements of the same group become more and more basic (*i.e.* less and less acidic) in the same direction. For example the oxides of MO_2 type of the elements of group IV A become more and more basic (*i.e.* less and less acidic) from CO_2 to PbO_2 as shown below :

$$\underbrace{CO_2,\ SiO_2,}_{\text{Acidic}}\quad \underbrace{GeO_2,\ SnO_2,}_{\text{Amphoteric}}\quad \underset{\text{Basic}}{PbO_2}$$

—Basic character increasing →

Similarly the oxides of M_2O_3 type of the elements of group VA become more and more basic (*i.e.* less and less acidic) from N_2O_3 to Bi_2O_3 as shown below :

$$\underbrace{N_2O_3,\ P_2O_3,}_{\text{Acidic}}\quad \underbrace{As_2O_3,\ Sb_2O_3,}_{\text{Amphoteric}}\quad \underset{\text{Basic}}{Bi_2O_3}$$

—Basic character increasing →

4. Electronegativity. Since the electropositive character of *p*-block elements, as we have seen above, decreases on moving in a period and increases down the group, electronegativity which is opposite of electropositive character will show opposite periodic variation, *i.e.* the electronegativity of *p*-block elements decreases down the group and increases in a period as is evident from the values given in Table 3.12.

Table 3.12. Electronegativity values (Pauling's values) of p-block elements.

Group→ / Period↓	III A	IV A	V A	VI A	VII A	Zero
2	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne –
3	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar –
4	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 2.9
5	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
6	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn –

5. Oxidation states and inert electron pair effect. The lighter elements of p-block like those of 2*nd* and 3*rd* period show a variety of oxidation states, both positive and negative, as shown below :

6. Electron affinity. Electron affinity values (EA values) of *p*-block elements are given in Table 3.13. Minus sign indicates the release of energy while plus sign indicates the absorption of energy.

From these values the following points may be noted :

(*i*) On moving from left to right in a period, EA values increase. However, there are certain elements in each period which have abnormal values. For example in 2*nd* period N and Ne have abnormal values. Similarly in 3*rd* period P and Ar have abnormal values. The abnormal values of N (2*nd* period) and P (3*rd* period) have been explained on the basis of the extra stability associated with their half-filled 2*p* (in N) and 3*p* (in P) orbitals. Being extraordinarily stable these orbitals are not able to accept the extra electron. Hence these elements have comparatively smaller EA values. On the basis of extra stability of ns^2p^6 configuration of Ne and Ar, their zero values can be explained.

Table 3.13. Electron affinity values (in KJ/mole) of p-block elements

Group→ / Period ↓	III A	IV A	V A	VI A	VII A	Zero
1						$He \xrightarrow{+e^-} He^-$ ~ 0
2	$B \xrightarrow{+e^-} B^-$ –23	$C \xrightarrow{+e^-} C^-$ –122	$N \xrightarrow{+e^-} N^-$ –20.1	$O \xrightarrow{+e^-} O^-$ –141 $O \xrightarrow{+2e^-} O^{2-}$ +780	$F \xrightarrow{+e^-} F^-$ –322	
3	$Al \xrightarrow{+e^-} Al^-$ –44	$Si \xrightarrow{+e^-} Si^-$ –120	$P \xrightarrow{+e^-} P^-$ –74	$S \xrightarrow{+e^-} S^-$ –200.4 $S \xrightarrow{+2e^-} S^{2-}$ +590	$Cl \xrightarrow{+e^-} Cl^-$ –348.7	$Ar \xrightarrow{+e^-} Ar-$ ~ 0
4	$Ga \xrightarrow{+e^-} Ga^-$ –36	$Ge \xrightarrow{+e^-} Ge^-$ –116	$As \xrightarrow{+e^-} As^-$ –77	$Se \xrightarrow{+e^-} Se^-$ –195 $Se \xrightarrow{+2e^-} Se^{2-}$ +420	$Br \xrightarrow{+e^-} Br^-$ –324.5	$Kr \xrightarrow{+e^-} Kr-$ ~ 0
5	$In \xrightarrow{+e^-} In^-$ –34	$Sn \xrightarrow{+e^-} Sn^-$ –121	$Sb \xrightarrow{+e^-} Sb^-$ –101	$Te \xrightarrow{+e^-} Te^-$ –190.1	$I \xrightarrow{+e^-} I^-$ –295	$Xe \xrightarrow{+e^-} Xe-$ ~ 0

(*ii*) On moving down the group, EA values go on decreasing as is evident from these values for (*a*) C, Si and Ge (Group IV A) (*b*) S, Se and Te (Group VI A) (*c*) O, S and Se for two electrons (Group VI A) and (*d*) Cl, Br and I (Group VII A). However, EA values of some of the elements of 2*nd* period are lower than the EA values of the elements lying just below them in 3*rd* period. For example B < Al, N < P, < O < S and F < Cl. The lower values of the elements of 2*nd* period are explained by saying that, due to the smaller size of the atoms of the elements of 2*nd* period, the addition of an extra electron to these atoms produces high electron density round the resulted anions. This high electron density increases the repulsion between the electrons (called electron—electron repulsion) already present in the relatively compact 2*p* orbital of the 2*nd* shell of these atoms and the electron being added. Due to this electron—elctron repulsion, the atoms of the elements of 2*nd* period show lesser tendency to attract the extra electron from outside and hence lower value of electron affinity for these elements.

7. Oxidising and reducing property. The oxidising property of an element depends on the magnitude of its electron affinity. An element with higher value of electron affinity can easily accept electron(s) and hence can act as a good oxidising agent. Thus an element with higher value of electron affinity will be a stronger oxidising agent than the element with lower value of electron affinity. Now since electron affinity decreases on descending a group and increases along

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Group→ / Period ↓	III A	IV A	V A	VI A	VII A	Zero
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3	$Al \xrightarrow{+e^-} Al^-$ −44	$Si \xrightarrow{+e^-} Si^-$ −120	$P \xrightarrow{+e^-} P^-$ −74	$S \xrightarrow{+e^-} S^-$ −200.4 $S \xrightarrow{+2e^-} S^{2-}$ +590	$Cl \xrightarrow{+e^-} Cl^-$ −348.7	$Ar \xrightarrow{+e^-} Ar^-$ ~ 0
4	$Ga \xrightarrow{+e^-} Ga^-$ −36	$Ge \xrightarrow{+e^-} Ge^-$ −116	$As \xrightarrow{+e^-} As^-$ −77	$Se \xrightarrow{+e^-} Se^-$ −195 $Se \xrightarrow{+2e^-} Se^{2-}$ +420	$Br \xrightarrow{+e^-} Br^-$ −324.5	$Kr \xrightarrow{+e^-} Kr^-$ ~ 0
5	$In \xrightarrow{+e^-} In^-$ −34	$Sn \xrightarrow{+e^-} Sn^-$ −121	$Sb \xrightarrow{+e^-} Sb^-$ −101	$Te \xrightarrow{+e^-} Te^-$ −190.1	$I \xrightarrow{+e^-} I^-$ −295	$Xe \xrightarrow{+e^-} Xe^-$ ~ 0

(*ii*) On moving down the group, EA values go on decreasing as is evident from these values for (*a*) C, Si and Ge (Group IV A) (*b*) S, Se and Te (Group VI A) (*c*) O, S and Se for two electrons (Group VI A) and (*d*) Cl, Br and I (Group VII A). However, EA values of some of the elements of *2nd* period are lower than the EA values of the elements lying just below them in *3rd* period. For example B < Al, N < P, < O < S and F < Cl. The lower values of the elements of *2nd* period are explained by saying that, due to the smaller size of the atoms of the elements of *2nd* period, the addition of an extra electron to these atoms produces high electron density round the resulted anions. This high electron density increases the repulsion between the electrons (called electron—electron repulsion) already present in the relatively compact 2*p* orbital of the *2nd* shell of these atoms and the electron being added. Due to this electron—elctron repulsion, the atoms of the elements of *2nd* period show lesser tendency to attract the extra electron from outside and hence lower value of electron affinity for these elements.

7. Oxidising and reducing property. The oxidising property of an element depends on the magnitude of its electron affinity. An element with higher value of electron affinity can easily accept electron(s) and hence can act as a good oxidising agent. Thus an element with higher value of electron affinity will be a stronger oxidising agent than the element with lower value of electron affinity. Now since electron affinity decreases on descending a group and increases along as compared to Sn.

8. Diogonal relationship. Boron which is a *p*-block element and is present in the 2*nd* period of the periodic table shows similarities in properties with silicon which is situated diagonally opposite to it in 3rd period. This type of across-the-street similarity in properties existing between B and Si is called diagonal relationship. This relationship can be explained on the basis of the concept of electronegativity and polarisation of ions.

9. Flame colouration. None of the *p*-block elements or its salt gives characteristic colour in the bunsen flame, when it is heated in it. The reason is that excitation energy given out by the electrons, when they come back to the original (ground) energy level, does not appear in the vissible region of the spectrum.

10. Formation of hydrides. Many of the elements of *p*-block form hydrides. The stability of the hydrides of the elements of a given group decreases as we descend the group. For example the stability of the hydrides of group VA decreases from NH_3 to BiH_3.

$$NH_3 > PH_3 > AsH_3 > SbH_3 > BiH_3$$

—Stability decreases →

Again we know that, since the groups IV A, V A and VI A begin with non-metals (C, N and O) and end with metals (Pb, Bi and Po), the hydrides of non-metals are more stable than those of metals.

The aqueous solution of these hydrides show acidic character which increases from top to bottom in a group. For example the acidic nature of the hydrides *viz*., HF, HCl, HBr and HI in aqueous solution increases from HF to HI.

$$HF < HCl < HBr < HI$$

—Acidic character increases →

11. Formation of halides. *p*-block elements also form halides. The covalent character of the halides containing different cations belonging to the *same period* of the periodic table but same halide ion increases from left to right in the period. For example in 2*nd* period BCl_3 is partially ionic while CCl_4 is covalent. Similarly in 5*th* period $SbCl_3$ is partly ionic whereas $TeCl_4$ is essentially covalent. The increase in covalent character from left to right in a period is due to the increase of the polarising power of the cation to polarise the anion due to the decrease of the size of the cations from left to right in a period (*One of the Fajans's rules*).

The covalent character of the halides containing the same cation but different anions belonging to the same group increases from top to bottom in a group. For example AlF_3 has essentially ionic character. $AlCl_3$ has intermediate character while $AlBr_3$ and AlI_3 have essentially covalent character.

AlF_3	$AlCl_3$	$AlBr_3$, AlI_3
Essentially ionic character	Intermediate character	Essentially covalent character

—Covalent character increasing →

The increase in covalent character from top to bottom is due to the increase of the polarisability of the anions or due to the increase of the size of the anions from top to bottom in a group (*One of the Fajans's rules*).