Inorganic Chemistry

Periodic Properties

Prof. K N Upadhya
326 SFS Flats
Ashok Vihar Phase IV
Delhi -110052

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Introduction

The most fascinating aspect of inorganic chemistry as well as its most difficult problem is the diversity in the chemistry of over one hundred elements. The tool that helped the inorganic chemistry systematize the chemistry of elements is the **periodic classification** of elements. The first comprehensive classification of elements was made independently by Dimitri Mendeleev in Russia and Lother Meyer in Germany. While Lother Meyer was primarily interested in the relationships of physical properties against atomic weights, Mendeleev based his system of classification on chemical characterization as a periodic function of atomic weights – the periodic law. After an elucidation of atomic structure and the discovery of isotopes, it was realized that the atomic number of an element is a more fundamental property than its atomic weight. Thus in its revised form the periodic law is stated as:

The properties of elements are periodic functions of their atomic numbers.

In other words, when the elements are listed in the order of increasing atomic numbers, the elements having closely similar properties will fall at definite intervals along the list. This tabular arrangement of elements having similar properties placed in vertical columns is called a *periodic table*. A large number of versions of the periodic table have been proposed but the one that is easiest to use and which is most closely related to the electronic structure of the atoms is the so called *long form*. Let us have a brief look at this form of the periodic table first.

1. Long Form of Periodic Table

This may be regarded as extended form of the Mendeleev's periodic table. Subgroups A and B of Mendeleev's tables are separated in this table. The vertical columns are numbered from 1 to 18; each vertical column represents a chemical family or group. The groups headed by the members of the two 8-members horizontal rows are designated as main group elements. Thus under the present system the main groups are Groups 1, 2,13,14,15,16,17 and 18. The middle section of the periodic table from group 3 to group 12 is occupied by the transition elements. Each horizontal row constitutes a period; the lengths of the period vary. There is a very short period containing only 2 elements, hydrogen and helium. This is followed by two short periods of 8 elements each and two long periods of 18 elements. The next, i.e. sixth period includes 32 elements. The last period is apparently incomplete. With this arrangement, elements in the same vertical column have similar properties.

To keep the periodic table from being excessively large, the two series of inner transition elements (4 *f* and 5 *f*) are placed in separate rows at the bottom of the table. (See Table 1)

Table 1: Long Form of Periodic Table

	S **block**												***P*** **block**					
Group	1	2											13	14	15	16	17	18
1	^{1}H																^{1}H	^{2}He
2	^{3}Li	^{4}Be	d- block										^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
3	^{11}Na	^{12}Mg	3	4	5	6	7	8	9	10	11	12	^{13}Al	^{14}Si	^{15}P	^{16}S	^{17}Cl	^{18}Ar
4	^{19}K	^{20}Ca	^{21}Sc	^{22}Ti	^{23}V	^{24}Cr	^{25}Mn	^{26}Fe	^{27}Co	^{28}Ni	^{29}Cu	^{30}Zn	^{31}Ga	^{32}Ge	^{33}As	^{34}Se	^{35}Br	^{36}Kr
5	^{37}Rb	^{38}Sr	^{39}Y	^{40}Zr	^{41}Nb	^{42}Mo	^{43}Tc	^{44}Ru	^{45}Rh	^{46}Pd	^{47}Ag	^{48}Cd	^{49}In	^{50}Sn	^{51}Sb	^{52}Te	^{53}I	^{54}Xe
6	^{55}Cs	^{56}Ba	^{57}La	^{72}Hf	^{73}Ta	^{74}W	^{75}Re	^{78}Os	^{77}Ir	^{78}Pt	^{779}Au	^{80}Hg	^{81}Tl	^{82}Pb	^{83}Bi	^{84}Po	^{85}At	^{86}Rn
7	^{87}Fr	^{88}Ra	^{89}Ac															

***f* block**

Lanthanoids	^{58}Ce	^{58} Pr	^{60}Nd	^{61}Pm	^{62}Sm	^{63}Eu	^{64}Gd	^{65}Tb	^{66}Dy	^{67}Ho	^{68}Er	^{69}Tm	^{70}Yb	^{71}Lu
Actinoids	^{90}Th	^{91}Pa	^{92}U	^{93}Np	^{94}Pu	^{95}Am	^{96}Cm	^{97}Bk	^{98}Cf	^{99}Es	^{100}Fm	^{101}Md	^{102}No	^{103}Lr

i. Chief Merits of the Long Form of Periodic Table

(i) This periodic table clearly reflects the electronic structure of elements. Elements falling in a particular column have a certain pattern of electronic structure and physical and chemical properties.

(ii) It separates the most metallic elements from non-metallic elements or weakly metallic elements.

(iii) The positions of the transition elements and relationships among them are quite clear.

(iv) The block-wise division (*s p d f* blocks) of the elements is clearly reflected in this table.

ii. Relationship between the Long Form of Periodic Table and Electronic configurations of Elements

(1) Each period starts with the filling of a new energy level. Two elements hydrogen and helium, for example, constitute the first period with electronic configurations $1s^1$ and $1s^2$ respectively, the only orbital 1*s* (corresponding to n = 1) being full.

(2) Each subsequent period starts with an alkali metal containing one electron in the outermost n s orbital and terminates with a noble gas having 8 electrons (n s^2 n p^6) in the outermost energy level.

(3) Elements with similar electronic configuration in the outermost orbitals are arranged in the same group.

(4) The periodic table with the exception of hydrogen and helium can be divided into four categories depending upon the type of atomic orbitals that are being filled with electrons. These are referred to as *s*-block, *p* – block, *d* – block and *f*-block elements according to whether, the electrons enter s, *p*, *d* or *f* orbitals in atoms of these elements.

This arrangement clearly reflects the periodic nature of electronic configurations of elements, which ultimately gives us an understanding as to why the elements with similar properties fall into such well defined groups. We shall consider the periodicity in electronic configurations first.

iii. Periodicity in Electronic Configurations of Elements

As stated above there is a direct connection between the successive placing of electrons into atomic orbitals and the periodic table. In other words the periodic chemical and physical behaviour of the elements are a natural consequence of the periodic recurrence of the same outermost electronic configuration. This may be clearly understood by examining the electronic configurations of elements through the successive periods of the periodic table.

Electronic configurations of elements in periods

Period 1 The two elements hydrogen ($1s^1$) and helium ($1s^2$) fall in this period. Hydrogen with one electron in the *s* orbital is analogous to the alkali metals and is placed in the Group

1. Helium represents the closure of the period with 1*s* orbital fully occupied and is placed with the noble gases in group 18 of the periodic table.

Period 2 There are 8 elements in this period, all possessing the helium core ($1s^2$). It begins with lithium $2s^1$ and beryllium $2s^2$ after which the 2*p* orbitls start filling from boron $2p^1$ to neon $2p^6$ and the period is complete at neon ($2s^2\ 2p^6$).

Period 3 This period contains the elements Na, Mg, Al, Si, P S, Cl and Ar which have electronic configuration of neon ($1s^2$, $2s^2$, $2p^6$) plus those derived from the regular filling of the 3*s* and 3*p* orbitals as have been described for that of 2*s* and 2*p* orbitals of should be noted that the third quantum level (M) has not received the electrons to its full capacity in this period as the 3*d* orbitals are still unoccupied. The electronic configurations of the first two elements are Na [Ne] 3*s*, Mg [Ne] $3s^2$. There after the 3*p* orbitals start getting electrons and the 3*p* orbitals are full at argon, Ar [Ne] $3s^2\ 3p^6$ which represents the closure of this period .

Period 4 The first two elements are potassium and calcium in which the outer electrons occupy the next lowest energy orbitals which is the 4s since it has a lower energy than the 3*d* orbitals. In the next ten elements (scandium to zinc) the five 3*d* orbitals are progressively filled. Thus the outer electronic configurations of these ten elements are

Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
$3d^14s^2$	$3d^24s^2$	$3d^34s^2$	$3d^54s^1$	$3d^54s^2$	$3d^64s^2$	$3d^74s^2$	$3d^84s^2$	$3d^{10}4s^1$	$3d^{10}4s^2$

The next orbitals to be used in building up the elements are the 4p set which are filled in a regular fashion in the elements gallium to argon.; Thus completing the fourth period or the first long period of the periodic table. The elements from scandium to zinc corresponding to the filling of 3*d* orbitals are said to constitute the first series of transition elements. During the filling of 3*d* orbitals one should take note of the uneven build up of these orbitals which occurs at chromium and copper due to extra stability of half-filled (d^5) and filled (d^{10}) *d* orbitls respectively.

Period 5 Broadly speaking this period follows closely the sequence in period 4 except that in the transition series the adoption of an outermost electron by the 4*d* orbitals occurs more often. The first two elements of the period are rubidium and strontium having the electronic configurations [Ar] $4s^2\ 4p^6\ 5s^1$ and [Ar] $4s^2\ 4p^6\ 5s^2$ respectively. The second series of transition elements starts with yttrium (39) and formally ends with cadmium (48). The filling up of 5*p* orbitals starts with indium and is completed at xenon. The period ends with xenon. The successive entry of electrons into 5*s*, 4*d* and 5*p* sets of orbitals is analogous to that for period 4 in general

Period 6 This period begins with caesium and barium corresponding to electrons entering the 6*s* orbitals. Although an electron enters the 5*d* orbitals at lanthanum, the lanthanoid series of elements (cerium to lutetium) corresponds to the filling up of the 4*f* orbitals these *f*-block elements are characterized by a marked similarity in their chemical properties. The third row of *d*-block elements then continues from hafnium to mercury as the 5*d* orbitals are filled with electrons. Finally the 6*p* orbitals are successively occupied from the element thallium to the noble gas radon. The lanthanoids also known as inner transition elements are shown below the framework of the periodic table.

Period 7 This period begins with francium and radium; the next elements are actinium and thorium. The build up of 5*f* then proceeds. The protactinium (atomic number 91) has the electronic configuration *[*Rn] $5f^2$ $6d^1$ $7s^2$ and nobelium (atomic number 102) [Rn]$5f^{14}$ $7s^2$.

The elements corresponding to the filling of electrons into 5*f* orbitals are regarded to constitute the other series of inner transition elements. These are called *actinoids* in many ways analogous to the lanthanoids. Elements beyond 103 continue to receive electrons in 6*d* level; thus another series of transition elements begins which is yet to be completed.

Electronic configurations of elements in groups
Element in the same vertical group have similar electronic configurations, i.e. they have the same number of electrons in the outer orbitals. Electronic configurations of the elements of Group1 are, for example, given below:

3	Lithium	[He] $2s^1$
11	Sodium	[Ne] $3s^1$
19	Potassium	[Ar] $4s^1$
37	Rubidium	[Kr] $5s^1$
55	Caesium	[Xe] $6s^1$

The outer configuration (s^1) is characteristic of the alkali metals, ls^1 (for hydrogen) being an exception.

Similarly it may be seen that each group has a characteristic outer electronic configuration. The alkaline earth metals (Group-2) have s^2 with the exception of $1s^2$ (for helium). The successive groups of non-transition elements have the outer configuration.

Group	13	14	15	16	17	18
Outer configuration	s^2p^1	s^2p^2	s^2p^3	s^2p^4	s^2p^5	s^2p^6

The group-wise configurations for the transition elements also exhibit a fair degree of uniformity. This is not important, however. A correlation between the electronic configurations of the periodic table for the transition elements is better sought by making a reference to their horizontal rows in the periodic table.

The well-known similarities among elements leading to grouping together of such elements as the alkali metals and the halogens clearly correspond to similarities in electronic configuration. Thus Periodic Law although originally entirely empirical is now recognized to be very solidly founded on atomic structures of the elements. It might indeed be restated: *The physical and chemical properties of elements are a function of the electronic configurations of their atoms, which vary in a periodic manner with increasing atomic number.*

2. Periodicities of Some Atomic Properties

Most of the physical and chemical properties of elements are dependent on the electronic structures of their atoms. Since there exists periodicities of electronic structures of atoms, most of the atomic properties of elements are also found to be periodic in nature. Thus the properties such as atomic size, ionization enthalpy, electron gain enthalpy and

electronegativity exhibit a fair degree of periodicity. Now we turn to a consideration of these properties and their relation to positions of elements in the periodic table.

i. Atomic Size: From results of X-rays diffraction studies in solids and from electron diffraction or spectroscopic examination of gaseous molecules, a large amount of information is now available on interatomic distances in the elements and their compounds.

Many elements including the noble gases and most metals are monoatomic in the elemental state, discrete atoms, not chemically bound to one another are present in the solid state. If we consider the simple picture of atoms as spheres of a definite radius which pack together so that adjacent atoms touch each other, the measured atomic distance would correspond to twice the atomic radius. This concept of a fixed size for an atom conflicts with the wave-mechanical viewpoint which regards the electronic density in atoms as extending indefinitely from the nucleus approaching zero approximately. Strictly speaking, it is impossible to define the size of the atoms. Nevertheless, it is helpful in discussing the structure of the elements (and their covalent compounds) to assign a consistent set of radii which represent the relative sizes of the atoms. Atomic radii were assigned by Pauling from the inter-atomic distances observed in a number of elements. For a few non-metals such as nitrogen and oxygen, the atomic radius is calculated from interatomic distances in compounds containing a single covalent bond between non-metalic atoms. For example, the radius of nitrogen atom is derived from the N-N instance in hydrazine $H_2N\text{-}NH_2$ and that of oxygen from the O-O distance in hydrogen peroxide, HO – OH. Values for atomic radii of metal atoms are half the interatomic distance in the solid state. For example, the distance between two adjacent copper atoms in solid is 256 pm; hence the metallic radius of copper is assigned a value of 128 pm. For simplicity, the term atomic radius is referred to both covalent and metallic radius depending on whether the element is a non-metal or a metal.

The atomic radii of a few elements are listed in Table. 2a. two trends are obvious. The atomic size decreases across a period as illustrated for the elements of the second period because within the period the outer electrons are in the same valence shell and the effective nuclear charge increases as the atomic number increases resulting in the increased attraction of electrons to the nucleus.

Within a family or vertical column of the periodic table, the atomic radius increases regularly with atomic number as shown in Table 2b for alkali metals and halogens. As we descend a group, the principal quantum number (n) increases and the valence electrons are farther from the nucleus. Consequently the atomic size increases as reflected in the atomic radii. There is, however an exception to the vertical trend which occurs with the elements just following the lanthanoids. The fact that these elements do not have the expected increase in their atomic sizes as their congeners in the proceeding periods is the effect of lanthanoid contraction. This phenomenon is associated with the filling of 4*f* orbitals before 5*d* which results in a regular decrease in atomic radii. The net result of the lanthanaoid contraction is that the 4*f* and 5*d* series of elements have almost similar atomic radii.

Table 2a: Atomic Radii / pm across the period

Atom	Atomic radius	Atom	Atom radius
Li	152	Na	186
Be	111	Mg	160
B	88	Al	143
C	77	Si	117
N	70	P	110
O	74	S	104
F	72	Cl	99

Table 2b: Atomic Radii / pm down the Group across a Family

Atom	Atomic radius	Atom	Atom radius
Li	152	F	72
Na	186	Cl	99
K	231	Br	114
Rb	244	I	133
Cs	262	At	140

ii. **Ionic and Crystal Radii**: In an ionic crystal, the ions may be regarded as though they are in contact with one another and hence the measured inter-atomic distance corresponds to the sum of the radii of the cation and anion. As in the case of atomic radii, it is advantageous to have some idea of relative size of ions and several attempts have been made to arrive at a suitable set of ionic radii. The set most widely used is due to Pauling. He took the experimental values of interionic distances in a number of crystals namely NaF, KCI, RbBr and CsI and deduced from them a set of ionic radii which closely reproduce the observed interionic distances in many other compounds.

An additional factor is that ionic radii vary with the coordination number of the ions. Usually the radius increases with increase in coordination number of the ions. Hence, when comparing ionic radii, we should compare like with like and use value for a single coordination number (typically six). Very extensive list of consistent values on thousands of compounds particularly oxides and fluorides are now available and some are given in Table3.

Table 3: Ionic radii (pm)*

Li^+	Be^{2+}	B^{3+}		O^{2-}	F^-
59 (4)	27 (4)	12 (4)		135 (2)	128 (2)
76 (6)				138(4)	131(4)
Na^+	Mg^{2+}	Al^{3+}		140 (6)	133 (6)
99(4)	49(4)	39(4)			
102 (6)	72(6)	53(6)			
K^+	Ca^{2+}	Ga^{3+}			
138(6)	100(6)	62(6)			
Rb^+	Sr^{2+}	In^{3+}	Sn^{2+}	Sn^{4+}	
149(6)	116(6)	79(6)	122(8)	69(6)	
160(8)	125(8)	92(8)			
Cs^+					
167(6)					
174(8)					

* Numbers in parentheses are coordination numbers of ions.

The general trends for ionic radii are the same as for metallic radii. Thus ionic radii increase in going down a group;

$Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$

but with the lanthanoid contraction restricting the increase among the heaviest ions, the radii of the ions of the same charge decrease across a period.

$Ca^{2+} > Mn^{2+} > Zn^{2+}$

When an ion occurs in different environments with different coordination numbers, its radius increases as the coordination number increases:

$2 < 4 < 6 < 8 < 10 < 12$

If an element can exist with different oxidation numbers, for a given coordination number its ionic radius deceases with increasing positive charge;

$Fe^{2+} > Fe^{3+}$

Since a positive charge indicates a reduced number of electrons, and hence more dominant nuclear attraction, cations are usually smaller than anions.

When we find some atoms and ions which contain the same number of electrons, we call them isoelectronic . For example O^{2-}, F^-, Na^+ and Mg^{2+} have the same number of electrons (10). Their radii would be different because of their different nuclear charges. The cation with greater number of charge will have smaller radius because of the greater attraction of the electrons to the nucleus. Anions with the greater negative charge will have larger radius. In this case the greater repulsion of the electrons will outweigh the nuclear charge and the ion will expand in size. Thus the increasing order of the sizes of these ions is

$Mg^{2+} < Na^+ < F^- < O^{2-}$

iii Ionization Enthalpy: A quantitative measure of the tendency of an element to lose electron is given by its ionization enthalpy. It represents the energy required to completely remove the most loosely bound electron from an isolated gaseous atom in its ground state. In other words the first ionization enthalpy for an element X is the enthalpy change (Δ H)for the process

$$X_{(g)} \longrightarrow X_{(g)}^+ + e^-_{(g)}$$

It is possible to remove the second, third, fourth, etc electrons and the succeeding enthalpies are the second, third, fourth, etc. ionization enthalpies.

The ionization enthalpy is one of the few fundamental properties of an atom which we are able to measure directly. The magnitude of ionization enthalpy gives a quantitative measure of the stability of the electronic structure of the isolated atom. The important factors which govern ionization enthalpy are:

(i) the magnitude of nuclear charge,
(ii) the distance of the outermost electron from the nucleus, that is, the atomic radius,
(iii) the shielding effect of the underlying shells of electrons, and
(iv) the extent to which the outer electron penetrates the charge cloud set up by the lower-lying electrons.

With regard to the last effect it is found that the degree of penetration of electrons in a given principal quantum shell decreases in the order $s > p > d > f$. This corresponds to the extent of binding of the various electrons. An n *s* electron is more tightly bound than an n *p* electron, which, in turn is more tightly bound than an n *d* electron, etc.

Shielding effect (constant):
Due to its greater extent of penetration, an *s* electron more effectively shields the nucleus than a *p* electron and a *p* electron more effectively shields the nucleus than a *d* electron and so on. Utilizing this basic idea it has been possible to evaluate a shielding constant(s) for a particular electron. This constant is a measure of the extent to which the outer electrons are able to shield the nucleus from the chosen electron. To determine the shielding constants write down the configuration of electrons in the following order and groupings.

(1*s*), (2*s*, 2*p*) (3*s*, 3*p*) (3*d*) (4*s*, 4*p*) (4*f*) (5*s*, 5*p*) etc.

Each grouping is assigned a contribution per electron. The sum of all such contributions that affect a given electron is the shielding constant(*s*). The assigned contributions are outlined below:

(1) Electrons in any groups to the right of the group of chosen electron contribute nothing to the *S*.
(2) All other electrons in the group considered contribute to an extent of 0.35 each (except in the 1*s* group, where 0.30 is used).
(3) All the electrons in the (n-1) contribute $S = 0.85$ each.
(4) All the electrons in (n -2) or lower shells $S = 1$ each.
(5) All the electron in groups lying lower in sequence than the (n *d*) or (n*f*) group contribute $S = 1$ each.

The above set of empirical rules called Slater's rules are useful guide for estimating the effective nuclear charges experienced by electrons in different orbitals. Lets us consider the application of these rules in a few cases. Take, for example potassium, the electronic configuration of which is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$. The effective nuclear charge (Z^*) experienced by 4*s* electron of potassium is

$$Z^* = Z - S = 19 - [(0.85 \times 8) + (1.00 \times 10)] = 2.20$$

Take an other example (4s) in zinc, the electronic configuration of which is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$

$$Z^* = Z - S = 3\text{- - } [(o.35 \times 1) + (0.85 \times 18) + (1 \times 10)] = 4.35$$

The ionization enthalpies of the gaseous atoms are plotted as a function of atomic number in Fig. 1. It is clear that there is a periodicity in the values of the ionization enthalpies that parallel the chemical properties of the elements. For the non – transition elements, there are fairly simple trends with respect to ionization enthalpy and position in the periodic table within a given family, increasing n trends to cause reduced ionization enthalpy because of the combined effects of size and shielding. The transition and post- transition elements show some anomalies in this regard and these will be discussed later.

Within a given period, there is a general tendency for the ionization enthalpy to increase with increase in atomic number. This is a result of the tendency for effective unclear charge to increase progressing from left to right. There are two other factors which prevent this from being monotonic. One is the change in type of orbital which occurs as one goes from Group 2 (*s* – orbital) to Group 13 (*p*- orbital). The second is the exchange energy between electrons of like spins. This stabilizes a system of parallel electron spins because electrons having the same spin tend to avoid each other as a result of Pauli's exclusion principle. The electrostatic repulsions between electrons are thus reduced.

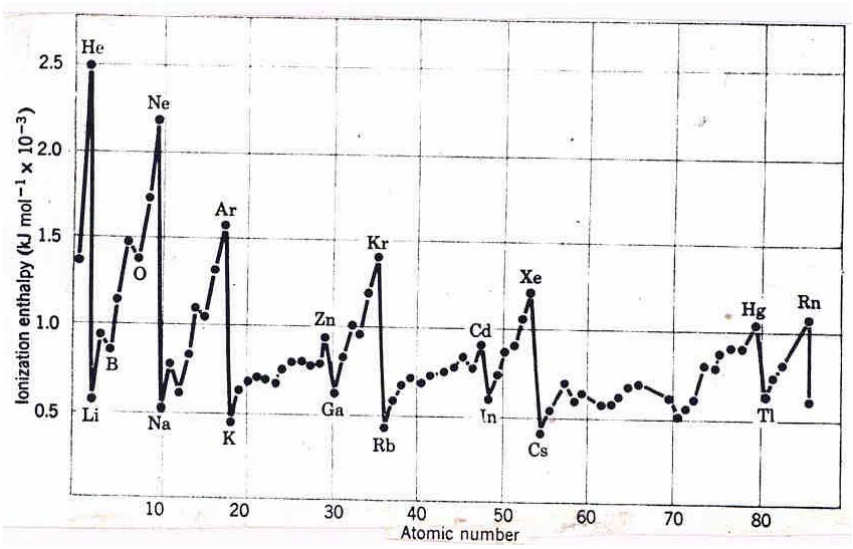

Fig. 1: Ionization enthalpy of elements as function of atomic number

In the light of above let us analyze trends in ionization enthalpy for some elements (Fig.1) The ionization enthalpy of helium is greater than that of hydrogen and this is due to the increase in nuclear charge from 1 to 2. But the increase is not as great as might be expected. Changing nuclear charge from 1 to 2 should increase the ionization enthalpy from 1310 kJ mol^{-1} (for hydrogen) to 5240 kJ mol^{-1}. The fact that the observed ionization enthalpy for helium is only 2369 is a result of the repulsion of the two electrons, which makes the helium atom less stable than might be expected. Other notable exceptions are met for p^1 and p^4 configuration where ionization enthalpy is smaller than expected because of the greater ease of removal of an electron to give a stable configuration involving empty or half –filled orbitals. The ionization enthalpy for boron is less than that for beryllium. This is an indication that *p*-electrons tend to be slightly higher in enthalpy than *s*-electrons of the same principal quantum number and thus require less enthalpy for their removal. The addition of second and third 2 *p*-electrons in carbon and nitrogen is accompanied by the increase in ionization enthalpy which may again be attributed to the increasing nuclear charge. To understand the slight drop at oxygen, the filling of *p* –orbitals is to be understood. The outer electronic configuration of nitrogen is $2s^2\ 2p_x^1\ 2p_y^1\ 2p^1{}_z$. The fourth electron in oxygen must be placed in a *p*-orbital which has already an electron in it. Apparently the extra repulsion which results from two electrons occupying the same orbital offsets the increased nuclear charge, and the ionization enthalpy of oxygen is slightly less than that of nitrogen. As the fifth and sixth *p*-electrons are added, the effect of increasing nuclear charge overcomes electron repulsion and the ionization enthalpy rises to a maximum at neon.

Fig.1 shows that the trend set in the second period is repeated in the next period. The ionization enthalpies of aluminum and sulphur are found to be exceptionally low.

Very little change in the ionization enthalpy occurs across a given transition series. For that matter very little change occurs among all the transition elements. This would appear to be the result of the interplay of the several factors. While the size is remaining relatively constant, the increased nuclear charge is offset by increased shielding by the additional electrons entering lower lying *d* orbitals.

As would be expected, the increase in atomic number in a given family produces a decrease in ionization enthalpy. This is associated with an increase in size, while the same type of outer configuration is maintained. This indicates that the effect of increased nuclear charge is

more than counterbalanced by the size increase and the presence of more shielding electrons. There is, however, an exception to the vertical trend, and it occurs with the elements just following the lanthanoids.

The fact that these elements have higher ionization enthalpies than their congeners in the preceding period is another effect of the lanthanoid contraction, which results from the relatively large increase in nuclear charge without expansion to higher shells.

Considering, the post-transition elements we find that ionization enthalpies for zinc, cadmium and mercury are highest in the respective series and then suddenly a drop occurs at gallium, indium and thallium. The peaks at zinc, cadmium and mercury and the minima at gallium, indium and thallium are most likely the result of filled *s* orbitals stability and the very effective screening of the *s* pair.

Summarizing the trends in the ionization enthalpies of elements in the periodic table the following generalizations may be noted:

1. As electrons are added to the orbitals of the same principal quantum number in successive elements, the ionization enthalpy increases due to the increase in nuclear charge.
2. Electrons of the highest principal quantum number are shielded from the nucleus by the inner or core electrons.
3. When several *p* or *d*-orbitals are available one electron enters each orbital until all are half-filled (Hund's rule). The half-filled set of orbitals with all spins the same seems to be particularly stable; addition of another electron often results in the decrease of ionization enthalpy.
4. For elements in the same periodic group with the exception of the transition elements, there is a tendency for ionization enthalpy to decrease with the increasing atomic number.

Table 4 gives the ionization enthalpies of some of the selected elements. (The ionization enthalpies of most of the elements are given in Appendix). The increased enthalpy required to remove a second electron, the second ionization enthalpy, or a third electron, a third ionization enthalpy can easily be understood in terms of the greater amount of enthalpy required to remove an electron from (a + n) species than from [(a+ (n –1)] species.

Table 4: Ionization Enthalpies of gaseous atoms of the first 30 elements (kJ mol^{-1})*

Atomic Number	Element	I	II	III	IV
1.	H	1310			
2.	He	2369	5242		
3.	Li	519	7290	11800	
4.	Be	898	1755	14830	24996
5.	B	799	2424	3659.9	2098
6.	C	1085	3250	4615	6316
7.	N	1400	2854	4573	7465
8.	O	1312	3388	5296	7461
9.	F	1679	3372	6040	8410
10.	Ne	2078	3959	6270	9367
11.	Na	495	4560	6905	9530
12.	Mg	737	1449	7725	10538
13.	Al	577	1814	2742	11566
14.	Si	785	1575	3226	4351
15.	P	1061	1894	2742	11566

16.	S	999	2257	3373	4566
17.	Cl	1254	2294	3846	5141
18.	Ar	1519	2663	3943	5764
19.	K	418	3066	4598	5872
20.	Ca	589	1144	4937	6479
21.	Sc	632	1243	2386	7106
22.	Ti	660	1314	2713	4168
23.	V	648	1371	2863	4598
24.	Cr	652	1602	2984	4765
25.	Mn	716	1508	3248	
26.	Fe	761	1560	2954	
27	Co	757	1643	3229	
28.	Ni	735	1750	3389	
29	Cu	744	1956	3550	
30.	Zn	905	1731	3287	

* To a good approximation, the values at 298 K may be obtained by adding 6.21 kJ per electron removed to the listed values.

iv. Electron Gain Enthalpy: It is found that some gaseous negative ions are energetically stable, though extremely reactive. For example, the energy required to remove an electron from a fluoride ion, F^- is 332.3 kJ mol^{-1}.

$$F^-(g) \rightarrow F(g) + e^- \quad \Delta H = + 332.3 \text{ kJ mol}^{-1}.$$

Plainly this energy is the ionization enthalpy of the fluoride ion. However, such quantities are written in terms of the reverse process, the capture of an electron by a neutral atom. The energy released is known as the electron gain enthalpy (Δeg H). Electron gain enthalpy provides a measure of the case with which an atom adds an electron to form anion as represented by the following equation:

$$X(g) + e^- \rightarrow X^-(g) : \Delta H = \Delta \text{eg } H$$

Depending upon the nature of element the process of adding an electron to the atom can be either exothermic or endothermic. For many elements, the energy is released when an electron is added to the atom, the electron gain enthalpy is negative. For example, the halogens (Group 17) have very high negative electron gain enthalpies as they can attain stable noble gas configuration by picking up an electron. On the other hand the noble gases with completely filled outer *s* and *p* orbitals have large positive electron gain enthalpies because the electron has to enter the next higher principal quantum level leading to an unstable electronic configuration. No element has a negative second electron gain enthalpy. For example the formation of dinegative ion like O^{2-} from O^- is an endothermic process, i.e. the electron gain enthalpy of O^- is not negative.

The magnitude of electron gain enthalpy is dependent upon various factors. In particular, the atomic size and effective nuclear charge are important. In general, the negative electron gain enthalpy decreases with increasing atomic radius and increases with decreased screening by inner *d* electrons. Also the value will depend upon the type of orbital that the added electron enters; other things remaining constant, the negative electron gain enthalpy is greatest for an electron entering an *s* orbital and decreases for *p*, *d* and *f* orbital. In fact in the common monoatomic anions the electrons enter a *p* orbital (except for H, where the 1*s* orbital receives electron).

Like ionization enthalpy, electron gain enthalpy is also a periodic property. However, the periodicity in electron gain enthalpy is much less pronounced than in ionization enthalpy. Note that both the negative electron gain enthalpy and the ionization enthalpy increase

toward the upper – right hand corner of the periodic table because one is defined in terms of energy released and the other in terms of needed. Table 5 shows that negative electron gain enthalpies of the halogen atoms are greater than those of the other elements. The decrease of negative electron gain enthalpy from chlorine to iodine is not unexpected because the atomic size decreases down then group. Surprisingly the negative electron gain enthalpy of F (- 327. 9 kJ mol^{-1}) is slightly less than that of Cl (- 348. 8 kJ mol^{-1}). This may be explained in this way: the very small size of fluorine atom with its high electron density makes the addition of an electron slightly less favourable energetically than for Cl.

Table 5: Electron Gain Enthalpies* / (kJ mol^{-1}) of some main group Elements

H	-73					He	+ 46
Li	-60	0	-141	F	-328	Ne	+116
Na	-53	S	-200	Cl	-349	Ar	+ 96
K	-48	Se	-195	Br	-325	Kr	+ 96
Rb	-47	Te	-190	1	-295	Xe	+ 77
Cs	-46	Po	-190	At	-270	Rn	+ 68

*In many earlier books the opposite of the negative electron gain enthalpy is defined as Electron Affinity (A) of the atom under consideration. If energy is released on the addition of an electron to an atom, the electron affinity is taken to be positive, contrary to thermodynamic convention. If energy has to be supplied to add an electron to an atom then the electron affinity of the atom is assigned a negative sign.

Note: Electron gain enthalpies of only a few elements have been determined directly. For many elements these values can be derived by indirect means involving Born – Haber cycle.

v. Electronegativity: When a covalent bond exists between two atoms of different elements, it is very likely that one of the atoms will attract the electron pair more strongly than the other. The atom which exerts the stronger attraction is said to be more electronegative. The relative tendency of a bonded atom in a molecule to attract electrons is expressed by the term electronegativity. Earlier this concept was associated with the rather crude concept of metallic and nonmetallic character of elements; the more nonmetallic character of an element, the greater its electronegativity. Although it is not possible to determine absolute values of electro- negativities of atoms, relative tendencies of bonded atoms to attract electrons can be estimated. A number of electronegativity scales have thus been suggested. The most widely used scale is that proposed by L. Pauling and subsequently modified by others. We shall first consider the scale suggested by Pauling and then other scales of electenegativity.

Pauling's Scales of Electronegativity

Pauling based his electronegativity scale on 'excess' bond energies. The relation between bond energy and electronegativity can be understood from the following example. It is found that 430 kJ of heat is required to break a mole of H_2 molecules. Thus, the bond energy of H_2 per bond is 430 /6.02 $\times 10^{23}$ kJ = 71. 5 $\times$ 10^{-23}kJ. Because the sharing of electron pair is equal it may be assumed that each bonded atom contributes half of the bond energy i.e. 35.7 $\times$ 10^{-23} kJ.

Furthermore, it would be reasonable to assume that in any bond in which hydrogen shares an electron pair equally with another, the contribution by H to the bond energy should be 35.7×10^{-23} kJ. Similarly from the bond energy of Cl_2, 239 kJ, it may be deduced that a chlorine atom should contribute 19.8×10^{-23} kJ to any bond in which the sharing of electron pair is equal.

Now, if we consider HCl and supposing that the electron pair in it is shared equally, the expected bond energy of HCl should be the sum of the contributions of H and Cl atoms, 55.5×1^{-23} kJ. Actually the bond energy of HCl is 426 kJ / mole or 70.6×10^{-23} kJ per bond. The fact that the observed bond energy is significantly greater than the calculated value, 55.5×10^{-23} kJ., suggests that electrons are not equally shared in HCl. The bond in question is actually more stable (requires more energy to break) than would be predicted by equal sharing.

The enhanced stability of HCl can be attributed to unequal sharing of the electron pair which makes the chlorine end of the molecule negative and hydrogen end positive. Since the positive and negative ends would attract each other, there would be addition to bond energy. The amount of additional bond energy would depend on the relative attracting tendency of the bonded atoms since greater the charge difference between the two ends greater will be the additional bond energy. Thus it should be possible to estimate the relative electronegativities from the difference between observed bond energies and those calculated on the assumption of equal sharing. This additional bond energy was called *ionic – covalent resonance energy* by Pauling. Experimental and calculated bond energies of hydrogen halides are given in Table 6. It is evident that discrepancy is maximum in HF and least in HI. This implies that the sharing of electrons between H and F is more unequal than between H and I. Pauling made use of 'ionic covalent resonance energy' for suggesting a scale of electronegativities.

Table 6: Bond Energy of Hydrogen Halides in kJ mol^{-1}

Bond	X = F	X= Cl	X= Br	X=I
H-H	431.4	431.4	431.4	431.4
X-X	159.5	239.1	189.8	148.4
X-X(calc)	292.6	335	310.9	290.1
H-X (obs)	504.3	427	395.1	294.3
Difference	271.7	91.5	48.5	4.18

In general, the ionic-covalent resonance energy, Δ(A-B), can be defined as:-.

$$\Delta(A\text{-}B) = D(A\text{-}B) - \tfrac{1}{2}[D(A\text{-}A) + D(B\text{-}B)]$$

where D(A-B) is the bond energy observed experimentally , and D(A-A) and D(B-B) are the covalent bond energies for the bonds A-A and B-B respectively. The ionic covalent resonance energy, Δ(A-B) , values are found to be negative for some of the active metal hydrides but the negative value can be avoided by using the geometric mean of the bond energies D(A-A) and D(B-B) instead of arithmetic mean, thus the ionic covalent resonance energy using geometric mean is

$$\Delta' (A\text{-}B) = D(A\text{-}B) - [\, D(A\text{-}A) \times D(B\text{-}B)\,]^{1/2}$$

The two methods do not differ appreciably for most bonds but geometric means are always smaller than arithmetic means. Table 7 lists the ionic covalent resonance energies calculated by the two methods.

Table 7: Ionic Covalent Resonance Energies

	Arithmetic means	Geometric means
Bond energy in kJ /mol , Cl_2	D(Cl-Cl) = 2390.5	D(Cl-Cl) = 239.5
Bond energy in kJ / mol. F_2	D(F-F) = 154.7	D(F-F) = 154.7
Average in kJ / mol, ClF	$\frac{D(Cl-Cl)+D(F-F)}{2}$	$[D(Cl\text{–}Cl) \times D(F\text{-}F)^{12}]$
	= 197.1	= 193.5
Experimental Bond energy In kJ / mol, ClF	D(Cl –F) = 248.78	D(Cl – F) = 248.7
Ionic covalent resonance energy	(Cl-F) = 248.7-197.1 = 51.6	(Cl-F) = 248.7 – 193.5 = 55.2

In order to assign the electronegativity values of elements, Pauling used the relationship:

$$\Delta = 23.06\, (X_A - X_B)^2$$

or $$X_A - X_B = 0.208\sqrt{\Delta}$$

where X_A and X_B are the numerical values of the electronegativity of elements A and B respectively, and the factor 0.2208, arises from the conversion of kcal to electron volt (1 eV = 23.06 kcal / mol).

The method gives values for the difference in electronegativity but does not give absolute values. Thus one element must be chosen and arbitrary value assigned to it as a standard for deriving values for others. To give the values a suitable range, hydrogen is given the value 2.1 arbitrarily. Pauling thus set up the scale of electronegativity for other elements relative to hydrogen 2.1. The Pauling's values resulting from improved energy data are listed in Table 8.

Other Methods of Estimating Electronegativity:

Many other methods of determining electronegativity values of elements have been suggested. Only two general methods considered to be important are described below:

Mulliken scale of electronegativity: Mulliken, shortly after Pauling, suggested that attraction of an atom for electrons should be an average of the ionization energy and electron affinity. Electronegativity is then estimated as the average of ionization potential and the electron affinity (both in electron valts, 1eV/ molecule = 96.36 kJ mol^{-1}) of the element.

$$\text{Electronegativity} = \frac{IE + EA}{2}$$

where *IE* and *EA* are the valence state ionization energy and electron affinity respectively. These are not experimentally observed values but those calculated for the atom in its valence state as it exists in a molecule.

Mulliken's values are about 2.8 times as large as Pauling's, Mulliken's electronegativity has the advantage that different values can be obtained for elements taking into account different ionic states and the change in hybrid orbitals. The electronegativity of carbon, for example, in *sp* hybridization is 0.6 unit higher than for sp^3 hybridization.
One reasonably accurate conversions between Pauling and Mulliken electronegativity scale is

$$X_p = 1.35\, X_m^{\frac{1}{2}} - 1.35$$

where X_p denotes Pauling electronegativity and X_m a Mulliken electronegativity.

Allred and Rochow scale: Allred and Rochow defined electronegativity as the electrostatic force exerted by the nucleus on the valence electrons. Thus

$$X_A = 0.359\frac{Z^*}{r^2_A} + 0.744$$

Where Z* is the effective nuclear charge and r^2_A the covalent radius of the atom, expressed in Angstrom's unit. Electronegativities of elements based on Allred and Rochow have been calculated for most of the elements and are accepted as an alternative to Pauling's thermochemical method for estimating electronegativities. See Table 8. for Allred – Rochow scale of electronegativity.

With Allred – Rochow scale we find another insight into the variation of *X* through the periodic table since we know how Z_{eff} and *r* vary from element to element. Elements with the highest electronegativity can now be seen to be those with the highest effective nuclear charge and the smallest radii; these lie close to F.

Electronegativity Variations:
Although electronegativity is generally considered as an invariant property of atom, it actually depends on the oxidation state and the types of hybrid orbitals of the atom involved in the bond formation. This can be easily understood from the fact that in any calculation of electronegativity, the thermochemical quantities such as ionization enthalpy, bond energy, electron gain enthalpy, etc., are taken into considerations and all these quantities for an atom vary under different situations, so also there is variation in electronegativity. Consider for example, carbon atom in different hybridized states. The electronegativity of tetrahedral carbon (with sp^3 hybrization) as in methane is nearly the same as that of hydrogen. On the other hand the sp^2 hybridized carbon is found to be more electronegative as is evident from its greater reactivity than hydrogen. Finally with sp hybridization (as in acetylene) carbon has almost the same electronegativity as chlorine; hydrogen atoms in acetylene are definitely acidic. On Pauling's scale the electronegativities of carbon and nitrogen in different hybridized stats are given below:

	sp^3	sp^2	*sp*
C	2.48	2.75	3.27
N	3.88	3.94	4.67

Electronegativity values of most of the elements are given in Table 8.

Table 8: The Electronegativity Values of Elements*

1 I	2 II	3 IIII	4 IV	5 V	6 II	7 II	8 II	9 II	10 II	11 II	12 II	13 III	14 IV	15 III	16 II	17 I
H																
2.1																
2.2																
Li	Be											B	C	N	O	F
1.0	1.5											2.0	2.5	3.0	3.5	4.0
1.0	*1.6*											*2.0*	*2.6*	*3.0*	*3.4*	*3.0*
Na	Mg											Al	Si	P	S	Cl
0.9	102											1.5	1.8	2.1	2.5	3.0
0.9	*1.3*											*1.6*	*1.9*	*2.2*	*2.6*	*3.2*
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br
0.8	1.0	1.3	1.5	1.6	1.6	1.5	1.8	1.8	1.8	1.9	1.6	1.6	1.8	2.0	2.4	2.8
0.8	*1.8*	*1.4*	*1.5*	*1.6*	*1.7*	*1.6*	*1.8*	*1.9*	*1.9*	*1.9*	*1.7*	*1.8*	*2.0*	*1.0*	*2.6*	*3.0*
Rb	Sr.	Y	Zr	Nb	Mo	Tc	Ru	Rb	Pd	Ag	Cd	In	Sn	Sv	Te	I
0.8	1.0	1.2	1.4	1.6	1.8	1.9	2.2	2.2	2.2	1.9	1.7	1.7	1.8	1.9	2.1	2.5
0.8	*1.0*	*1.2*	*1.3*	*2.2*				*2.3*	*2.2*	*1.9*	*1.7*	*1.8*	*2.0*	*2.1*		*2.7*
Cs	Ba	La- Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At
0.7	0.9	1.1- 1.2	1.3	1.5	1.7	1.9	2.2	2.2	2.2	2.4	1.9	1.8	1.8	1.9	2.0	2.2
0.8	*0.9*	*1.1- 1.3*		*2.4*				*2.2*	*2.3*	*2.5*	*2.0*	*2.0*	2.3	1.0		
		Ac	Th	Pa	U	Np- No	Pu									
		1.1	1.3	1.5	1.7	1.3										
					1.4		1.3									

* The first values given for each element is Pauling's. The second (Italics) is Allred's. The oxidation state is specified below the group number at the top of each group.

Electronegativity and per cent Ionic Character:

An interesting feature of the electronegativities is that the difference, $X_A - X_B$ (on Pauling's scale) is almost equal to the dipole moment of the bond A- B. Several attempts have, therefore, been made to derive an equation for per cent ionic character in terms of the electronegativity difference, $X_A - X_B$. The most successful is

$$\text{Per cent ionic character} = 16\,(X_A - X_B) + 3.5\,(X_A - X_B)^2$$

Table 9: Comparison of Dipole Moment, μ with ($X_x - X_H$) for Hydrogen Halides

	H – F	H-Cl	H- Br	H- I
μ	1.98	1.03	0.79	0.38
$(X_x - X_H)$	1.9	0.9	0.7	0.4

On this basis the per cent ionic character of some typical bonds are given below:

C-H	N- H	O- H	F – H
4%	19%	39%	60%
C-F	C-Cl	C-Br	C-I
43%	11%	3%	1%

The electronegativity values for calculating the per cent ionic character should, however, be not taken too rigidly but for a rough guide for estimating it.

Ionic Character and Dipole Moment:
The relation between ionic character and dipole moment in a molecule such as HCl appears to be straightforward if the bond length (127 pm) and dipole moment (3.44×10^{-30} C m) are known:
H – Cl $\mu = qr = 3.44 \times 10^{-30}$ C m

$$q = \frac{3.44 \times 10^{-30} Cm}{127 \times 10^{-12} \ Cm} \times \frac{1.electron}{1.6 \times 10^{-19} C} = 0.17 \text{ electron}$$

Thus $H^{\delta +} = 0.17$ and $Cl^{\delta -} = - 0.17$

The simplicity of the method has made it popular in textbooks, though the situation is unfortunately not so simple.

Electronegativity Values and Periodic Table
Numerical values assigned for the electronegativities of the various elements are shown in Table 8. Fluorine has the highest electronegativity value, 4.0 of any element in the periodic table. The noble gas elements have only recently been found to form chemical bonds and their values have not yet been evaluated correctly. Otherwise, as we go from left to right across a period, the electronegativity increases. The elements as the far left side of the table have low electronegativity. In a group the electronegativity decreases with increasing atomic number; for example, the group 17 elements are assigned the values F, 4.0 ; Cl, 3.0; Br, 2.8; and I 2.5 ; the decrease being a regular.

The periodic trend in electronegativity value is important in the understanding of chemical behaviour of the elements. Another application of electronegativity values is that their differences can be related, semi-quantitatively to properties of bond energies and dipole moments.

Electronegativity Equalization:
When a covalent bond forms between two different atoms electron density shifts from one atom to the other until the electronegativities got equalized. Initially the more electronegative element will have a greater attraction for electrons but as the electron density shifts towards the atom it will become negative and tend to attract electrons less. Conversely, the atom which is losing electron becomes somewhat positive and attracts electrons better than it did when neutral. This process will continue until the two atoms attract electrons equally at which point the electronegativities will have been equalized and charge transfer will cease. Thus, a covalent bond formation is somewhat more complicated than an ionic bond formation. Although the latter can be treated to a first approximation by classical electrostatics, treatment of covalent bonds requires a quantum mechanical approach at the outset.

vi. Periodicity of Valence: We have seen that members of the same family of elements have identical number of electrons in the outermost orbitals. The electrons in the outermost orbitals which take part in chemical combinations are referred to as valence electrons. The noble gases are all relatively stable and chemically unreactive and all have completed *s* and *p* orbitals in the outermost shell (s^2 p^6, except for He $1s^2$). The valence of representative elements is usually equal to the number of electrons in the outermost orbitals and / or equal to the outermost electrons. Some periodic trends observed in the valence of elements (hydrides and oxides) are shown in Table10. Other such periodic trends occur in many other chemical properties of elements which are usually considered as and when the need arises. There are many elements which exhibit variable valence. This is particularly characteristic of the transition elements.

Table 10: Periodic trends in Valence of Representative Elements as shown by the Formulae of their Compounds.

Group	1	2	13	14	15	16	17
Formulae of hydrides	LiH NaH KH			CH_4 SiH_4 GeH_4 SnH_4	NH_3 PH_3 AsH_3 SbH_3	H_2O H_2S H_2Se	HF HCl HBr HI
Formulae of oxides	Li_2O Na_2O K_2O	MgO CaO SrO BaO	B_2O_3 Al_2O_3 Ga_2O_3 In_2O_3	CO_2 SiO_2 GeO_2 SnO_2 PbO_2			

vii. Periodicity in Some Chemical Properties: Many of the periodic properties considered above, viz electronic configuration, atomic size, ionization enthalpy, etc, determine the chemical properties of the elements. The two structural properties, i.e, the size of atom and the ease of removal of electron are the most important among them. Considering the periodic table as a whole the following conclusions can be reached.

On progressing in horizontal rows from left to right, the metallic character of elements decreases. The gradual increase in nuclear charge and decrease in size of elements result in more tightly bound electrons in their atoms. Thus atoms of smaller size constitute elements which are non-metallic, and large atoms form elements which are metals. Application of these considerations shows as to why elements on the left side of the periodic table form more basic oxides and the oxides of the elements on the right hand side are more acidic. In groups of *s*- and *p*-blocks the metallic character increases with atomic number from top to bottom. The best illustration is offered by group 15 in which non-metal nitrogen and metal bismuth occupy top and bottom positions respectively. A notable feature of the periodic table is that the first element in a group shows striking resemblance to the second element of the next higher group. This phenomenon is called *diagonal relationship*.

This is again because of the size and the ease of removal of electron. For example, on moving from lithium to beryllium the ionic size has decreased from 60 pm for Li to 31 pm for Be^{2+}. But from Be^{2+} to Mg^{2+} it has again increased to 0.65 for Mg^{2+}. Similar consideration can explain the resemblances between beryllium and aluminum and between boron and silicon.

When we consider the transition elements, nothing specifically can be stated unless one considers a particular oxidation state. Normally a transition element in lower oxidation state forms more basic oxide than the one in higher oxidation state.

Suggested Readings

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Appendix

Ionization Enthalpies of Gaseous Atoms (kJ mol^{-1} 0 K) *

Atomic No.	Element	I	II	III	IV
1.	H	1310			
2.	He	2369	5242		
3.	Li	519	7290	11800	
4.	Be	898	1755	14830	24996
5.	B	799	2424	3659.9	2098
6.	C	1085	3250	4615	6316
7.	N	1400	2854	4573	7465
8.	O	1312	3388	5296	7461
9.	F	1679	3372	6040	8410
10.	Ne	2078	3959	6270	9367
11.	Na	495	4560	6905	9530
12.	Mg	737	1449	7725	10538
13.	Al	577	1814	2742	11566
14.	Si	785	1575	3226	4351
15.	P	1061	1894	2742	11566
16.	S	999	2257	3373	4566
17.	CL	1254	2294	3846	5141
18.	Ar	1519	2663	3943	5764
19.	K	418	3066	4598	5872
20.	Ca	589	1144	4937	6479
21.	Sc	632	1243	2386	7106
22.	Ti	660	1314	2713	4168
23.	V	648	1371	2863	4598
24.	Cr	652	1602	2984	4765
25.	Mn	716	1508	3248	
26.	Fe	761	1560	2954	
27	Co	757	1643	3229	
28.	Ni	735	1750	3389	
29	Cu	744	1956	3550	
30.	Zn	905	1731	3287	
31.	Ga	577	1977	2959	6186
32.	Ge	761	1536	3298	4389
33.	As	945	1948	2729	4849
34.	Se	940	2073	3085	4134
35	Br	1141	2082	3461	
36	Kr	1349	2367	3557	
37	Rb	402	2654	3846	
38	Sr.	549	1063		5434
39	Y	614	1179	977	

40	Zr	660	1266	2215	3310
41	Nb	662	1381	2423	3691
42	Mo	685	1557	2615	4473
43	Tc	702	1471		
44	Ru	710	1615	2744	
45	Rh	719	1442	2993	
46	Pd	802	1872	3173	
47	Ag	730	2071	3357	
48	Cd	867	1630	3612	
49	In	558	1818	2702	5515
50	Sn	708	1410	2939	3925
51	Sb	833	1588	2437	4264
52	Te	869	1793	3010	3678
53	I	1008	1839		
54	Xe	1169	2044	3093	
55	Cs	375	2420		
56	Ba	502	964		
57	La	539	1102	1102	
72.	Hf	669	1438		
73	Ta	761	1563		
74	W	769	1705		
75	Re	761	1601		
76	Os	836	1630		
77	Ir	836			
78	Pt	878	1789		
79	Au	890	1977		
80	Hg	1005	1808	3298	
81	Ti	588	1968	2872	4890
82	Pb	715	1449	3078	4079
83	Bi	703	1608	2464	4347
84	Po	811			
85	At				
86	Rn			1036	
87	Fr				
88	Ra			509	978
89	Ac			669	1166

* To a good approximation, the values at 298 K may be obtained by adding 6.2 kJ per electron removed to the listed values.