

## CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

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### 3.1 Introduction

You have got preliminary knowledge about periodic table for the classification of elements in Standard 9. According to scientist Glenn T. Seaborg, the periodic table is the most important concept in chemistry. It always supports the students, suggests new avenues of research and presents the entire branch of chemistry in brief. In this unit, we will study the history of periodic table and modern periodic law. Moreover, we will understand the periodic classification and some of the periodic trends in the physical and chemical properties of the elements.

### 3.2 Necessity of Classification of Elements

In 1800, only 31 elements were known. In 1865, number of known elements became 63. Due to increasing number of elements, it became difficult to study the chemical properties of elements and their innumerable compounds individually. To solve this problem, scientists classified the elements. This classification was not only for understanding of chemical properties of elements but was also helpful in prediction of new elements. On account of this 114 elements are known to us today. Out of these, 92 elements are found in nature and rest are synthetic elements which are man-made. Efforts to synthesise new elements are continuing even today.

### 3.3 History of the Development of Periodic Table

The German chemist, Johann Dobereiner, in early 1800, was the first to consider the idea of trends among properties of elements. In 1829, he arranged many of the elements into triads (group of three) based on their physical and chemical properties. For example, lithium, sodium and potassium were grouped together as being soft and reactive metals. Dobereiner also observed that when these elements were arranged according to their atomic weights, the atomic weight of middle element was roughly the average of the atomic weights of first and

third (table 3.1). This relationship was known as the Law of Triads, but this was found in few elements and so it was rejected by considering it as a coincidence.

**Table 3.1 Dobereiner's Triads**  
(For information only)

Element	Atomic Weight	Element	Atomic Weight	Element	Atomic Weight
Li	7	Ca	40	Cl	35.5
Na	23	Sr	88	Br	80
K	39	Ba	137	I	127

In 1862, French geologist, A.E.B. de Chancourtois published first probable periodic table. He arranged known elements in cylindrical form (chart) in terms of increasing order of their atomic weights. He was first to find out that the properties of an element repeat after every seventh element. He was able to predict stoichiometry of some metal oxides using this chart. This work has also not attracted attention of anyone. The English chemist, John Newlands presented a research paper, in 1863, in which he classified 56 elements in 11 groups based on their physical properties. In 1865, Newlands published his version of periodic table and presented Law of Octaves. This law suggests that any element in periodic table shows similar behaviour with its eighth element (Table 3.2). He linked this law to the octaves of music (Sa, Re, Ga, Ma, Pa, Dha, Nee,... Sa...). This law of octaves proposed by Newlands was found to be true up to calcium. Even at that time this concept was also not accepted widely. Dmitri Mendeleev and Lothar Meyer published their periodic tables independently in 1869 and 1870, respectively. They both constructed their tables in a similar manner, by arranging the elements in a row or column in order of atomic weight and started with a new row or column when the characteristics of elements began to repeat. The credit for development of periodic law, that we know today, goes to the Russian chemist, Dmitri Mendeleev (1834-1907) and the German chemist Lothar Meyer (1830-1895). But Mendeleev was first to publish the periodic law. It can be stated as follows :

“The properties of the elements are periodic function of their atomic weights.”

Mandeleev was more courageous than Meyer. He assumed that if a measured atomic weight of an element is placed in the wrong

place in the periodic table, the atomic weight must be wrong. In some cases this was true. For example, previously assigned atomic weight of indium was nearly 76. But this could not be placed in the periodic table between arsenic (atomic weight 75) and selenium (atomic weight 79). Mendeleev suggested that its atomic weight should be 114 instead of 76. It was nearer to the currently accepted value 114.82. Besides chemical properties of indium it was inconsistency with the group pattern defined by the known properties of aluminium and thallium. It supported Mendeleev. In the same way though the atomic weight of K(39.10) was less than atomic weight of Ar (39.95), Mendeleev placed Ar before K, because Ar possessed properties similar to elements of group 0 and K possessed properties similar to elements of group I.

**Table 3.2 Newlands' Octaves**

Elements	Li	Be	B	C	N	O	F
Atomic weight	7	9	11	12	14	16	19
Elements	Na	Mg	Al	Si	P	S	Cl
Atomic weight	23	24	27	29	31	32	35.5
Elements	K	Ca					
Atomic weight	39	40					

Subsequent studies about periodic system have shown that the arrangement of elements is not in order strictly according to atomic weight. For example, atomic weight of iodine is less than that of tellurium (group VI). It was placed along with fluorine, chlorine and bromine in group VII because they showed similar properties (Fig. 3.1). Thus at that time, the basic concept of Mendeleev was to arrange the elements having similar properties in the same group. When Mendeleev published periodic table, gallium and germanium were not known. He left the gaps for unknown elements below the aluminum and the silicon. Mendeleev named these elements as “Eka-Aluminium” and “Eka-Silicon” (in Sanskrit ‘eka’ means ‘next’), respectively. Quite similarities in some properties were predicted by Mendeleev for these elements and the properties experimentally observed were found (Table 3.3). On these bases Mendeleev named “gallium” to “Eka-Aluminium” and “germanium” to “Eka-silicon”. Thus development of periodic table took place. Mendeleev's periodic table published in 1905 is shown in Fig. 3.1.

**Table 3.3 Mendeleev's predictions about properties of Eka-aluminium and Eka-silicon elements and experimental results**

(For information only)

Property	Eka-aluminium (Predicted)	Gallium (found)	Eka-Silicon (Predicted)	Germanium (found)
Atomic weight	68	70	72	72.6
Density (gcm <sup>-3</sup> )	5.9	5.94	5.5	5.36
Melting point (K)	Low	302.93	High	1231
Formula of oxide	E <sub>2</sub> O <sub>3</sub>	Ga <sub>2</sub> O <sub>3</sub>	EO <sub>2</sub>	GeO <sub>2</sub>
Formula of chloride	ECl <sub>3</sub>	GaCl <sub>3</sub>	ECl <sub>4</sub>	GeCl <sub>4</sub>

**Periodic system of the elements in groups and series (For Information Only)**

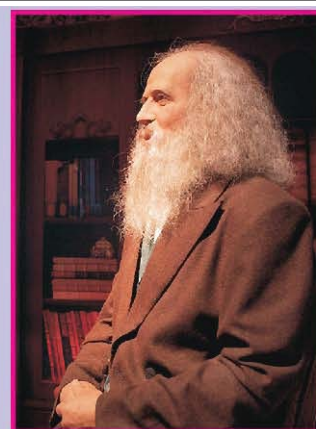
	I	II	III	IV	V	VI	VII	VIII	
1	Helium He 4.0	Lithium Li 7.03	Boron B 11.1	Carbon C 12.0	Nitrogen N 14.04	Oxygen O 16.00	Fluorine F 19.0		
2	Neon Ne 19.9	Sodium Na 23.5	Aluminium Al 27.0	Silicon Si 28.4	Phosphorus P 31.0	Sulphur S 32.06	Chlorine Cl 35.45		
3	Argon Ar 38	Potassium K 39.1	Calcium Ca 40.1	Thallium Tl 204.1	Vanadium V 51.4	Chromium Cr 52.1	Manganese Mn 55.0	Iron Fe 55.9	
4		Copper Cu 63.6	Scandium Sc 44.1	Germanium Ge 72.3	Arsenic As 75	Selenium Se 79	Bromine Br 79.95	Nickel Ni 59	
5			Zinc Zn 65.4					Cobalt Co 59	
6		Rubidium Rb 85.4	Strontium Sr 87.6	Zirconium Zr 90.6	Niobium Nb 94.0	Molybdenum Mo 96.0		Ruthenium Ru 101.7	
7	Krypton Kr 81.8	Silver Ag 107.9	Cadmium Cd 112.4	Tin Sn 119.0	Antimony Sb 120.0	Tellurium Te 127.6	Iodine I 126.9	Rhodium Rh 103.5	
8	Xenon Xe 128	Cesium Cs 132.9	Barium Ba 137.4	Cerium Ce 140				Palladium Pd (Ag) 106.5	
9			Lanthanum La 139						
10			Ytterbium Yb 173	Lead Pb 206.9	Tantalum Ta 183	Tungsten W 184		Osmium Os 191	
11		Gold Au 197.2	Thallium Tl 204.1		Bi 208			Iridium Ir 193	
12			Radium Ra 224	Thorium Th 232	Uranium U 239			Platinum Pt 194.9	
	R	R <sub>2</sub> O	RO	RO <sub>2</sub> RH <sub>4</sub>	HIGHER SALINE OXIDES R <sub>2</sub> O <sub>3</sub> RH <sub>3</sub>		HIGHER GASEOUS HYDROGEN COMPOUNDS RH		RO <sub>4</sub>

Mendeleev's periodic table published earlier (1905)

Figure 3.1

Dmitri Mendeleev was born in Tobolsk of Siberia in Russia. He received his Master's degree (M.Sc.) in chemistry in 1856 and the Doctoral degree (Ph.D.) in 1865.

He joined as professor of chemistry at the university of St. Petersburg in 1867. Preliminary work for his great text book *Principles of Chemistry* led Mendeleev to propose the periodic law and to construct his periodic table of elements. Mendeleev's periodic law gave encouragement to several areas of research during the subsequent decades. Periodic law and periodic table became the key to the discovery of uranium, thorium, noble gases and transuranium elements. Mendeleev invented an accurate barometer. He resigned from the professorship in 1890. Mendeleev passed away in 1907. To appreciate the work of Mendeleev, G. T. Seaborg suggested to give name Mendelevium to the element 101 discovered by him.



Dmitri Evgenovich Mendeleev  
(1834-1907)

### 3.4 Modern Periodic Law and Modern Periodic Table

When Mendeleev developed periodic table, chemists had no knowledge about internal structure of atom. However, principles about subatomic particles were developed in the beginning of 20th century. In 1913, the English physicist, Henry Moseley observed regularity in the characteristics of X-ray spectra of the elements. He plotted two types of graphs  $\sqrt{\nu}$  (where  $\nu$  is frequency of X-rays emitted) against atomic number and against atomic weight. First graph was found to be a straight line but second graph was not found to be a straight line. It indicated that atomic number was the fundamental property and not the atomic weight. As a result, periodic law of Mendeleev was corrected by putting the word 'atomic number' instead of 'atomic weight'. It is known as modern periodic law. It can be stated that

"The physical and chemical properties of elements are periodic function of their atomic numbers."

We know that atomic number is equal to the nuclear charge (i.e., number of protons) or number of electrons in a neutral atom. Many forms of periodic table have been considered from time to time. Some forms emphasise the electronic configuration of elements, whereas the other forms emphasise the chemical reactions and valence. "Long form" of periodic table of elements is known as modern periodic

table (Fig. 3.2) which is used widely. The elements having similar electronic configuration in their outermost orbit, are arranged in vertical column of periodic table known as group or family. According to the recommendation of International Union of Pure and Applied Chemistry (IUPAC), the numbering of groups is from 1 to 18 instead of numbers which are given by old method IA..... VII A, VIII, IB..... VII B and 0.

The horizontal rows in periodic table are called periods. There are altogether total seven periods. To identify the periods, numbers are given from the top of periodic table. Thus, period 3 is the third row of elements and period 4 is the fourth row of elements. First period has two elements. Following periods have 8, 8, 18, 18 and 32 elements respectively, whereas seventh period is the incomplete period. Period 1 is known as a very small period, period 2 and 3 are known as first and second small periods respectively, periods 4 and 5 are known as first and second long periods respectively, periods 6 and 7 are known as first and second very long periods respectively. List of 14 elements of both sixth and seventh periods are mentioned separately at the bottom in this form of periodic table.

Thus, we have seen the classification of elements in group, development of periodic law and periodic table are results of sincere efforts of many scientists.

PERIOD NUMBER	Representative elements																		Noble gases
	GROUP NUMBER																		
1	1 H 1s <sup>1</sup>																		2 He 1s <sup>2</sup>
2	3 IA 4 IIA																		9 VII B
3	5 VA 6 VIA 7 VII A 8 VIII 9 VIII 10 VIII 11 IB 12 IIB																		16 VI B 17 VII B
4	19 K 20 Ca 37 Rb 38 Sr 55 Cs 56 Ba 81 Tl 82 Pb 83 Bi 84 Po 85 At 86 Rn																		35 Kr 36 Kr
5	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																		53 I 54 Xe
6	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																		81 Tl 82 Pb 83 Bi 84 Po 85 At 86 Rn
7	57 La* 58 Ce 59 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																		87 Fr 88 Ra 89 Ac**

f - Inner transition elements

58 Ce 4f <sup>0</sup> 5d <sup>1</sup> 6s <sup>2</sup>	59 Pr 4f <sup>1</sup> 5d <sup>0</sup> 6s <sup>2</sup>	60 Nd 4f <sup>2</sup> 5d <sup>0</sup> 6s <sup>2</sup>	61 Pm 4f <sup>5</sup> 5d <sup>0</sup> 6s <sup>2</sup>	62 Sm 4f <sup>6</sup> 5d <sup>0</sup> 6s <sup>2</sup>	63 Eu 4f <sup>7</sup> 5d <sup>0</sup> 6s <sup>2</sup>	64 Gd 4f <sup>7</sup> 5d <sup>1</sup> 6s <sup>2</sup>	65 Tb 4f <sup>9</sup> 5d <sup>0</sup> 6s <sup>2</sup>	66 Dy 4f <sup>10</sup> 5d <sup>0</sup> 6s <sup>2</sup>	67 Ho 4f <sup>11</sup> 5d <sup>0</sup> 6s <sup>2</sup>	68 Er 4f <sup>12</sup> 5d <sup>0</sup> 6s <sup>2</sup>	69 Tm 4f <sup>13</sup> 5d <sup>0</sup> 6s <sup>2</sup>	70 Yb 4f <sup>14</sup> 5d <sup>0</sup> 6s <sup>2</sup>	71 Lu 4f <sup>14</sup> 5d <sup>1</sup> 6s <sup>2</sup>
90 Th 5f <sup>0</sup> 6d <sup>2</sup> 7s <sup>2</sup>	91 Pa 5f <sup>2</sup> 6d <sup>1</sup> 7s <sup>2</sup>	92 U 5f <sup>3</sup> 6d <sup>1</sup> 7s <sup>2</sup>	93 Np 5f <sup>4</sup> 6d <sup>1</sup> 7s <sup>2</sup>	94 Pu 5f <sup>6</sup> 6d <sup>0</sup> 7s <sup>2</sup>	95 Am 5f <sup>7</sup> 6d <sup>0</sup> 7s <sup>2</sup>	96 Cm 5f <sup>7</sup> 6d <sup>1</sup> 7s <sup>2</sup>	97 Bk 5f <sup>9</sup> 6d <sup>0</sup> 7s <sup>2</sup>	98 Cf 5f <sup>10</sup> 6d <sup>0</sup> 7s <sup>2</sup>	99 Es 5f <sup>11</sup> 6d <sup>0</sup> 7s <sup>2</sup>	100 Fm 5f <sup>12</sup> 6d <sup>0</sup> 7s <sup>2</sup>	101 Md 5f <sup>13</sup> 6d <sup>0</sup> 7s <sup>2</sup>	102 No 5f <sup>14</sup> 6d <sup>0</sup> 7s <sup>2</sup>	103 Lr 5f <sup>14</sup> 6d <sup>1</sup> 7s <sup>2</sup>

Long form of periodic table (modern periodic table.)  
Figure 3.2

\* Lanthanoids  
4f<sup>n</sup>5d<sup>0-1</sup>6s<sup>2</sup>

\*\* Actinoids  
5f<sup>n</sup>6d<sup>0-2</sup>7s<sup>2</sup>

### 3.5 Nomenclature of Elements with Atomic Numbers > 100

Scientist who discovers element can give the name to the newly discovered element by traditional method. This suggested name is approved by IUPAC. But nomenclature by this method created some problems. Because of the competition for discovery of new elements, scientists, before collecting the reliable data about new element, got tempted to claim for its discovery. For example, Americans and Soviets claimed for discovery of element 104. Americans named it as Rutherfordium whereas Soviets named it as Kurchotavium. The systematic nomenclature is suggested to avoid such types of problems. To use the numerals for number 0 and 1 to 9 in systematic nomenclature are shown in Table 3.4. The numerals are put together in order of digits which make up the atomic number and 'ium' is added at the end. This creates the name of an element. The nomenclature of elements with atomic number above 100 are shown in Table 3.5.

**Table 3.4**  
Notation for IUPAC nomenclature of elements

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	p
6	hex	h
7	sept	s
8	oct	o
9	enn	e

**Table 3.5**  
Nomenclature of elements with atomic numbers above 100

Atomic Number	Name	Sym- bol	IUPAC Official Name	IUPAC Symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf

105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununmillium	Uun	Darmstadtium	Ds
111	Unununnium	Uuu	Rontgenium	Rg*
112	Ununbium	Uub	*	*
113	Ununtrium	Uut	+	
114	Ununquadium	Uuq	*	*
115	Ununpentium	Uup	*	
116	Ununhexium	Uuh	*	*
117	Ununseptium	Uus	+	
118	Ununoctium	Uuo	+	

\* Official IUPAC name yet to be announced  
+ Elements yet to be discovered

Let us think about IUPAC name and symbol of element 119, according to Table 3.4, numerals of digit 1 and 9 are un and enn respectively. Therefore, the symbol of element (119) will be Uue and name will be Ununennium. Thus three letters consisting a symbol and a temporary name are given to the new element. Afterwards the permanent name and symbol are given to the element according to opinion of IUPAC representatives from each country. Permanent name is given considering the contribution of well known scientist or the country in which the element is discovered. For example, name Mendelevium is given to element with atomic number 101 for appreciation of Mendeleev's work. This name was suggested by American scientist Glenn T. Seaborg who discovered this element. In this manner naming the element with atomic number 106 as Seaborgium is to honour the Nobel Prize winner scientist Glenn T. Seaborg.

### 3.6 Electronic Configuration of Elements and the Periodic Table :

We have learnt in previous unit about four quantum numbers which explain the characteristics of electrons, energy levels (shells) (K, L, M...), atomic orbitals (s, p, d, f) and electronic configuration of elements. Now we shall study about relation between electronic configuration of elements and long form of periodic table.

#### (a) Electronic configuration in periods :

The period number indicates value of n for outer orbit or valence shell. This can be understood from Table 3.6.

**Table 3.6 Electronic configuration of elements in periods**

Period number	Elements of the period and electron configuration in their outermost orbit	Number of outermost orbits
First	Hydrogen ( $1s^1$ ) to Helium ( $1s^2$ ) ${}_1\text{H}$ to ${}_2\text{He}$	1
Second	Lithium ( $2s^1$ ) to Neon ( $2s^2 2p^6$ ) ${}_3\text{Li}$ to ${}_{10}\text{Ne}$	2
Third	Sodium ( $3s^1$ ) to Argon ( $3s^2 3p^6$ ) ${}_{11}\text{Na}$ to ${}_{18}\text{Ar}$	3
Fourth	Potassium ( $4s^1$ ) to Krypton ( $4s^2 4p^6$ ) ${}_{19}\text{K}$ to ${}_{36}\text{Kr}$	4
Fifth	Rubidium ( $5s^1$ ) to Xenon ( $5s^2 5p^6$ ) ${}_{37}\text{Rb}$ to ${}_{54}\text{Xe}$	5
Sixth	Cesium ( $6s^1$ ) to Radon ( $6s^2 6p^6$ ) ${}_{55}\text{Cs}$ to ${}_{86}\text{Rn}$	6
Seventh	Francium ( $7s^1$ ) to incomplete period ${}_{87}\text{Fr}$ to incomplete period	7

As shown in Table 3.6, the electronic configuration of the outermost orbit of the atom of every first element of each period is  $ns^1$  where as the electronic configuration of the outer most orbit of the last element is  $ns^2 np^6$  except that of the first period.

**(b) Electronic configuration in groups :** Electronic configuration in outermost orbit is similar in all elements of any group. Therefore, these elements show similar properties. For example, electronic configuration of outermost orbit of all elements of group 1 is  $ns^1$  as mentioned in Table 3.7. In same way electronic configuration of outermost orbit of all elements of halogen group is  $ns^2 np^5$ .

**Table 3.7 Electronic configuration of elements of group 1**

Atomic number	Symbol	Electronic configuration
3	Li	$1s^2 2s^1$ or $[\text{He}] 2s^1$
11	Na	$1s^2 2s^2 2p^6 3s^1$ or $[\text{Ne}] 3s^1$
19	K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ or $[\text{Ar}] 4s^1$
37	Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$ or $[\text{Kr}] 5s^1$
55	Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^1$ or $[\text{Xe}] 6s^1$
87	Fr	$[\text{Rn}] 7s^1$

### 3.7 Electronic Configuration and Types of Elements : s-, p-, d-, f- blocks

Classification of elements into different blocks in periodic table is made easy due to electronic configuration of atom. It can be classified as elements in s-block, p-block, d-block, f-block based on the orbital of element filled with electron lastly as mentioned in Fig 3.3. In having such types, hydrogen and helium are found to be exceptions. Due to electronic configuration of helium it will be the element of s-block, but its place may be considered in ground 18 of p-block in periodic table. This is because it has properties

similar to the other elements (noble gases) of group 18 due to its completely filled valence shell. Electronic configuration of hydrogen is  $1s^1$ . So it can be placed in group 1 but it can achieve the electronic configuration of noble gas helium by gaining one electron. Thus, its behaviour is similar to elements of group 17 (the halogen group). So H can be placed in group 1 or in group 17 but hydrogen is placed separately in the middle at the top of the periodic table as shown in Fig 3.2 and Fig 3.3. Now we will classify the elements in s-, p-, d- and f- blocks.

s - BLOCK		p - BLOCK												
		13	14	15	16	17	18							
1s	1	2												
2s	Li	Be												
3s	Na	Mg												
4s	K	Ca												
5s	Rb	Mg												
6s	Cs	Ba												
7s	Fr	Ra												

d - BLOCK											
	3	4	5	6	7	8	9	10	11	12	
3d	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	
4d	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	
5d	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	
6d	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Uuu	Uub	






  

f - BLOCK														
	3	4	5	6	7	8	9	10	11	12				
Lanthanoids 4f	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Actinoids 5f	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

p - BLOCK												
	13	14	15	16	17	18						
2p	B	C	N	O	F	He						
3p	Al	Si	P	S	Cl	Ar						
4p	Ga	Ge	As	Se	Br	Kr						
5p	In	Sn	Sb	Te	I	Xe						
6p	Tl	Pb	Bi	Po	At	Rn						
7p	-	Uuq	-	Uuh	-	-						

METALS (  ,  ,  , NON METALS (  ) and METALLOIDS (  )											
<p>The types of elements in the periodic table based on the orbitals that are being filled</p> <p>Figure 3.3</p>											



**3.7.1 The s-block elements :** The electrons are filled at last in s orbital in elements of group 1 (alkali metals) and group 2 (alkaline earth metals). Electronic configuration of these orbitals are  $ns^1$  and  $ns^2$ , respectively. Therefore, these elements are known as s-block elements. They are reactive metals with low ionization enthalpy.

**3.7.2 The p-block elements :** The elements in which the p-orbital is filled at last are called p-block elements. In elements of groups 13 to 18 the last electron is filled in p orbital so, they are included in p-block. Electronic configuration of outermost orbit of elements of each period varies from  $ns^2 np^1$  to  $ns^2 np^6$ . All orbitals of valence shell of elements of group 18 are completely filled with electrons. It is stable configuration. These elements do not gain or lose electrons. These gaseous elements, are known as noble gases. These elements do not participate in chemical reaction. It means that they remain inert so they are known as inert gases.

Groups 16 and 17 are known as chalcogens and halogens, respectively. Elements of s and p-block are called representative elements and main group elements.

**3.7.3 The d-block elements : Transition elements :** Elements in which the last electron is filled in d-orbital are known as d-block elements. In elements of groups 3 to 12 the last electron is filled in d-orbital so these elements are called d-block elements. General outer electronic configuration of these elements is  $(n-1)d^{1-10} ns^{1-2}$ . These elements are transition metals.

**3.7.4 The f-block elements : Inner-transition elements :** Elements in which the last electron is filled in f-orbital are called f-block elements. Elements of two horizontal rows i.e. Ce(Z = 58)–Lu(Z = 71) (Lanthanoids) and Th(Z = 90)–Lr(Z = 103) (Actinoids) which are kept at bottom of periodic table, have electronic configuration as  $(n-2)f^{1-14} (n-1)d^{0-1} ns^2$  in their outermost orbit. In every element of these two series, the last electron is filled in f-orbital so these elements are called f-block elements. They are also known as innertransition elements. They all are metals.

Metals appear on the left side of periodic table whereas non metals appear at the top right hand side of the periodic table. The change from metallic to non metallic character is not abrupt as shown by the thick zig-zag line in Fig : 3.3.

These elements (e.g., silicon, germanium, arsenic, antimony and tellurium) bordering this line and running diagonally across the periodic table show properties of both metals and nonmetals, so these elements are called semimetals or metalloids.

### 3.8 Periodic Trends in Properties of Elements

Many patterns in physical and chemical properties of elements are observed in periodic table as we move from left to right in a period or from top to bottom in a group. For example, in same period, chemical reactivity trend is high in metals of group 1, low in middle elements of periodic table and increases to maximum in nonmetals of group 17. In the same way reactivity increases when we go down in group 1 whereas reactivity decreases as we go down in group 17. Why do chemical and physical properties of elements show this trend ? How can we explain the periodic trend in properties ? Here we will try to get answers of these questions.

**3.8.1 Periodic Trends in Physical properties :** We find periodic deviation in many physical properties of elements. Here we will discuss about atomic radius, ionic radius, ionization enthalpy, electron gain enthalpy and electronegativity.

**3.8.1.1 Atomic radius :** As we know atom is a very minute particle, therefore, it is very difficult to measure the radius of individual atom. Nevertheless, an estimate for the atomic radius can be given by knowing the distance between the atoms in a molecule. Atomic radius can be measured by X-rays or other spectroscopic methods. Average half value of distance between nuclei of two atoms in covalent molecule is considered as atomic radius. It is known as covalent radius of element. For example, the bond distance in chlorine molecule ( $Cl_2$ ) is 198 pm. The half value (99 pm) of this distance is taken as atomic radius of chlorine. Average half value of distance of nucleus between two adjacent atoms in metallic crystal is called radius of metal element. It is known as metallic radius. For example, the distance between two adjacent copper atoms in solid copper is 256 pm so, metallic radius of copper will be 128 pm.

Atomic radius is related with electronic configuration of atoms. As shown in Table 3.8 (a) the positive charge increases with increase in atomic number of elements when we move from left to right in the period.

On the other hand number of electrons increases in outermost orbit. The number of outermost orbit (principal quantum number  $n$ ) does not increase so the distance of electron from nucleus does not increase. As a result the force of attraction acting on the electron towards the nucleus increases which causes shrinking in the atomic orbital, hence the atomic radius decreases. It is shown in Fig. 3.4 (a) for the elements of the second period. There is a negligible change in the atomic radius of d-block and f-block elements present in the middle of the long period.

**Table 3.8 (a) Atomic radius in periods (pm)\***  
(For information only)

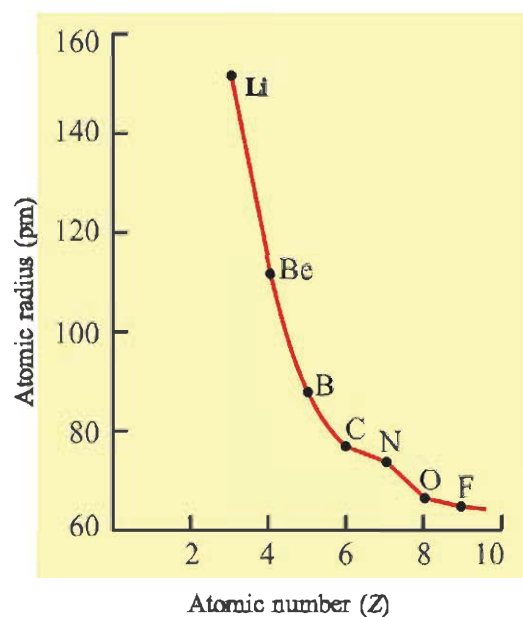
<b>Atom (Period II)</b>	Li	Be	B	C	N	O	F
<b>Atomic radius</b>	152	111	88	77	74	66	64
<b>Atom (Period III)</b>	Na	Mg	Al	Si	P	S	Cl
<b>Atomic radius</b>	186	160	143	117	110	104	99

pm\* = Picometer =  $10^{-12}$  meter

As shown in table 3.8(b) when we move from top to bottom in a group, principal quantum number ( $n$ ) increases with increase in atomic number, so valence electrons are farther away from the nucleus. Therefore, the positive charge of nucleus increases despite of decreasing of the attraction on electrons towards nucleus. Thus orbits expand and the atomic radius increases. It is explained in Fig. 3.4 (b) for alkali metals and halogen group.

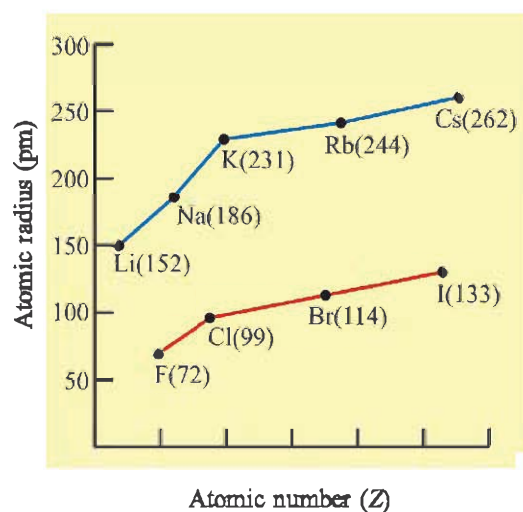
**Table 3.8 (b) Atomic radius (pm) in group**  
(For information only)

<b>Atom (Group 1)</b>	<b>Atomic radius</b>	<b>Atom (Group 17)</b>	<b>Atomic radius</b>
Li	152	F	64
Na	186	Cl	99
K	231	Br	114
Rb	244	I	133
Cs	262	At	140



Variation of atomic radius with atomic number across the second period

Figure 3.4 (a)



Variation of atomic radius with atomic number for alkali metals and halogens.

Figure 3.4 (b)

In the end of above discussion we will try to arrange three elements viz. Al, C and Si in ascending order of their atomic radii. The sequence becomes  $C < Si < Al$ . It is because atomic radius decreases when we move from left to right in same period. Al ( $Z=13$ ), Si ( $Z=14$ ) belong to same period (third period), Therefore sequence of their atomic radius is  $Si < Al$ ; where as C ( $Z=6$ ) and Si ( $Z=14$ ) belong to group 14. We know that the atomic radius increases when we move from top to bottom in a group. Hence the sequence becomes  $C < Si$ ,

**3.8.1.2 Ionic radius :** Generally atom forms positive ion by loss of electron and negative ion by gain of electron. Ionic radius can be measured by measuring the distance between positive ion and negative ion in ionic crystals. The positive ion is smaller than its parent atom because it has less electrons, whereas their positive charges are similar. This means that effective nuclear charge is higher in positive ion than its parent atom. For example, 11 protons attract 11 electrons in Na, whereas 11 protons attract 10 electrons in  $\text{Na}^+$ , so ionic radius of  $\text{Na}^+$  is 95 pm and atomic radius of Na is 186 pm. The size of negative ion is bigger than its parent atom because it has more electrons whereas their positive charges are same. Hence the effective nuclear charge is lower in negative ion than its parent atom. For example, 9 protons attract 9 electrons in F. Whereas 9 protons attract 10 electrons in  $\text{F}^-$ . Ionic radius of  $\text{F}^-$  is 136 pm and atomic radius of F is only 64 pm. Ionic radii of elements show similar trend to the atomic radii in periodic table.

We also find equal number of electrons in some atoms and ions. They are known as isoelectronic species. For example,  $\text{O}^{2-}$ ,  $\text{F}^-$ ,  $\text{Na}^+$  and  $\text{Mg}^{2+}$  all have equal number of electrons (10). Their radii are different because they have different nuclear charges. Effective nuclear charge is increased due to higher positive charge in a positive ion so its radius will be small, the effective nuclear charge is decreased due to higher negative charge in a negative ion which results in bigger radius.

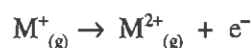
Now, we will think about size of Mg,  $\text{Mg}^{2+}$ , Al and  $\text{Al}^{3+}$  species. The Mg has the largest size and  $\text{Al}^{3+}$  has the smallest size. It is so because Mg and Al belong to the same period (period 3) and in the same period atomic radius decreases when we move from left to right so that  $\text{Al} < \text{Mg}$ . In isoelectronic species when the positive charge increases then the radius decreases. Hence  $\text{Al}^{3+} < \text{Mg}^{2+}$ . Moreover size

of positive ion is smaller than its parent atom. Therefore,  $\text{Mg}^{2+} < \text{Mg}$  and  $\text{Al}^{3+} < \text{Al}$ . Besides this, because of higher effective nuclear charge in  $\text{Mg}^{2+}$  than Al, it becomes  $\text{Mg}^{2+} < \text{Al}$ .

**3.8.1.3 Ionization enthalpy :** The minimum energy required for formation of positive ion from gaseous neutral atom (M) by removing electron, is called ionization enthalpy (energy). First ionization enthalpy is the energy required to remove first electron. In other words first ionization enthalpy of an element M is the difference of enthalpy ( $\Delta_i H_1$ ) for the reaction shown by the following equation

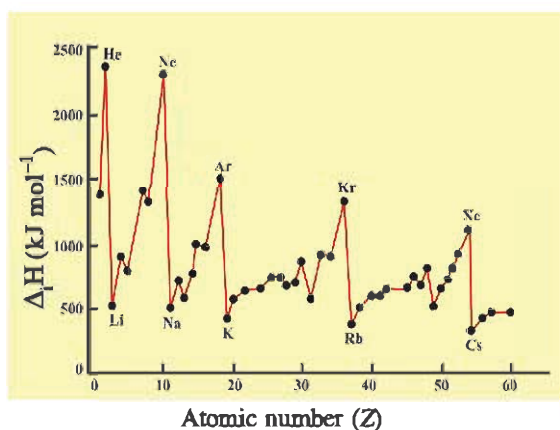


The ionization enthalpy is expressed in units of  $\text{kJ mol}^{-1}$ . Energy required to remove the second electron from element M is called second ionization enthalpy ( $\Delta_i H_2$ ).



The second ionization enthalpy will be higher than the first ionization enthalpy because of increased effective nuclear charge. The force of attraction of outermost electron towards the nucleus increases with increase in effective nuclear charge. Removal of electrons having higher force of attraction towards the nucleus requires more energy. Thus, ionization enthalpy also increases with increased effective nuclear charge. The energy required to remove the third electron is called third ionization enthalpy ( $\Delta_i H_3$ ). Its value is found higher than the second ionization enthalpy. Thus, the value of next ionization enthalpy increases gradually as  $\Delta_i H_1 < \Delta_i H_2 < \Delta_i H_3, \dots$

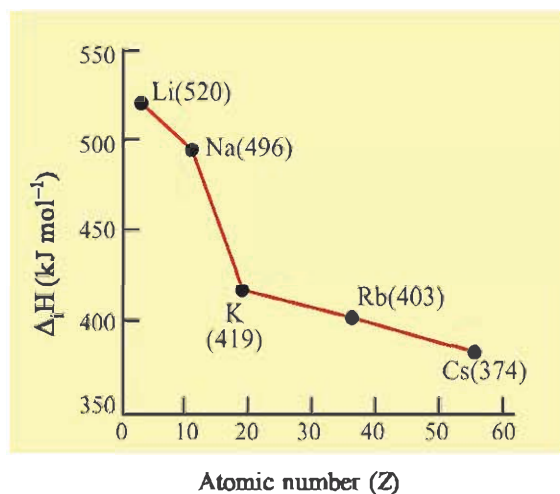
Relation between atomic number and first ionization enthalpy of elements having atomic numbers 1 to 60 is shown in Fig. 3.5. It indicates that values of first ionization enthalpy of noble gases are maximum. Whereas these values are minimum for alkali metals because the outermost orbit of noble gases is completely filled which gives higher stability. It is more difficult to remove electron from noble gases, whereas less energy is required to remove electrons from alkali metals due to their high reactivity.



Relation between first ionization enthalpies and atomic number with  $Z = 1$  to 60

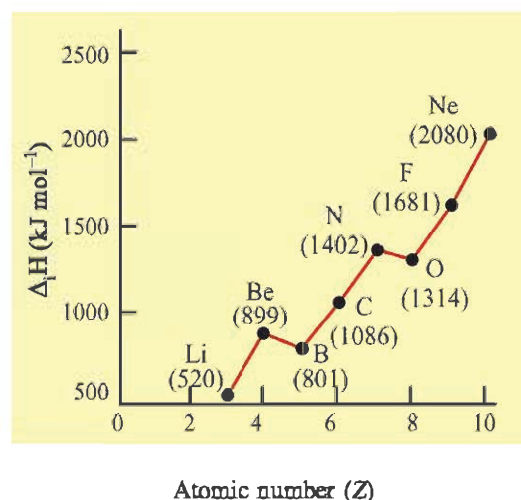
Figure 3.5

The inner orbital electrons of an element fulfil the need to shield or screen valence electrons and causes decreased attraction of valence electrons towards the nucleus. It results in decreased ionization enthalpy of valence electrons. This effect is called shielding or screening effect. When shielding effect increases then ionization enthalpy decreases. So, in any group when we move from top to bottom then ionization enthalpy decreases with increase in atomic number. This type of trend is clearly observed in representative elements. This can be seen in the graph of first ionization enthalpy of alkali metal elements and their atomic numbers as shown in Fig. 3.6 (a).



Relation between first ionization enthalpies of alkali metals elements and their atomic number

Figure 3.6 (a)



Relation between first ionization enthalpies of second period elements and their atomic number

Figure 3.6 (b)

In a period, when we move from left to right then ionization enthalpy increases with increase in atomic number. As shown Fig 3.6 (b) the ionization enthalpy increases from Li to Ne. In elements of same period as the atomic number increases positive charge of nucleus also increases. The outermost orbit remains the same for elements of the same period. This means that the distance of electrons of outermost orbit from nucleus does not increase, so the shielding effect for outer electrons also does not increase. That is why the force of attraction on electrons towards the nucleus increases due to increase in positive charge of nucleus. It results in the increase of enthalpy required to remove an electron from atom.

It can be seen from Fig 3.6(b) that boron ( $Z = 5$ ) has higher nuclear charge than beryllium ( $Z = 4$ ), even then its value of ionization enthalpy is low. What can be the reason for this? Let us think about it. Electronic configurations of beryllium and boron are  $1s^2 2s^2$  and  $1s^2 2s^2 2p^1$ , respectively. The electron to be removed is in  $2s$  orbital in beryllium and in  $2p$  orbital in boron. Therefore, more energy is required to remove the electron from beryllium than boron. The  $s$  type electron is in spherical orbital, hence it is nearer to nucleus than the electron of  $p$ ,  $d$  and  $f$  orbital of same orbit. The  $s$ -type electron which is nearer to nucleus remains highly attracted towards nuclear charge and so it is more difficult to remove it. If all the factors remain same then the ionization enthalpy decreases as we go from  $s \rightarrow p \rightarrow d \rightarrow f$  orbital with same principal quantum number.

Like boron and beryllium, the anomaly is also found in nitrogen and oxygen. The first ionization enthalpy is lower in oxygen than that of nitrogen as shown in Fig 3.6 (b). Electronic configuration of nitrogen ( $Z = 7$ ) and oxygen ( $Z = 8$ ) are  $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$  and  $1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$ , respectively. In oxygen the electron-electron repulsion value increases due to two electrons filled in  $2p_x$  orbital. So it is easy to remove the fourth electron from the  $2p$  orbital in oxygen than the removal of any one electron out of three electrons from  $2p$  orbital in nitrogen.

The first ionization enthalpies of the elements Na, Mg and Si of third period are 496, 737, and 786  $\text{kJ mol}^{-1}$ , respectively. Will the first ionization enthalpy of Al be closer to 575 or 760  $\text{kJ mol}^{-1}$ ? Let us think about it.

Electronic configuration of Na ( $Z = 11$ ) is  $[\text{Ne}] 3s^1$

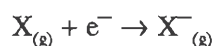
Electronic configuration of Mg ( $Z = 12$ ) is  $[\text{Ne}] 3s^2$

Electronic configuration of Al ( $Z = 13$ ) is  $[\text{Ne}] 3s^2 3p^1$

Electronic configuration of Si ( $Z = 14$ ) is  $[\text{Ne}] 3s^2 3p^2$

As mentioned in the question, probable value of ionization enthalpy of Al (575 or 760  $\text{kJ mol}^{-1}$ ) is nearer to the value of ionization enthalpy of Mg (737  $\text{kJ mol}^{-1}$ ). Thus we will have to study carefully about the electron configuration of Al and Mg. This suggests that less energy is required to remove an electron from  $3p$  orbital of Al than to remove an electron from  $3s$  orbital of Mg, so the value of first ionization enthalpy of Al will be 575  $\text{kJ mol}^{-1}$ .

**3.8.1.4 Electron gain enthalpy :** Addition of an electron to neutral gaseous atom (X) converts it into negative ion which results in change of enthalpy. This is known as electron gain enthalpy ( $\Delta_{\text{eg}}H$ ).



The reaction of adding an electron to the atom may be either exothermic or endothermic; it depends on corresponding element. When an electron is added to the atom then energy is released, so the value of electron gain enthalpy is negative. For example, elements of group 17 (halogens) have very high negative electron gain enthalpy. It is because they accept an electron and achieve the stable

electronic configuration similar to nearest noble gas element. The noble gas elements have more positive values of electron gain enthalpy because the added new electron enters in the next higher energy level orbit. Therefore, it becomes very unstable configuration. Thus, when the value of electron gain enthalpy is more negative than negative ion is easily formed which is more stable.

Generally if effective nuclear charge increases then trend of attraction of added electron towards nucleus increases. The elements with small size can easily show this trend. Therefore, in a period when we move from left to right the effective nuclear charge increases as the atomic number increases and the value of electron gain enthalpy becomes more negative. This trend is not exactly so. For example when  $\text{Be}^-$  is formed then added electron enters the energy orbital ( $2p$ ) which is higher than the valence orbital ( $2s$ ). This is not a stable state because the value of electron gain enthalpy becomes less negative. In the same way, generally the nitrogen atom also does not show the trend to accept electron because when an electron is added to it then  $2p_x$  orbital gains two electrons. This causes increased electron-electron repulsion due to which it becomes unstable.

In a group when we move from top to bottom then the added electron is found farther away from the nucleus due to increased atomic size. This results in decreased attraction of added electron towards the nucleus. Hence, the trend of atom to become negative ion by accepting electron, decreases gradually. Thus, generally, when we go from top to bottom the value of electron gain enthalpy becomes more negative (Table 3.9) but the electron gain enthalpy of O or F is less negative than that of the succeeding elements of the same group (S or Cl respectively) (Table 3.9). This is because in O or F, the added electron takes place in lower energy level  $n = 2$  which is already having 6 or 7 electrons. This results in more electron-electron repulsion which decreases the stability of negative ion, whereas in S and Cl added electron takes the place at higher energy level  $n = 3$ . Therefore, less electron-electron repulsion occurs.

**Table 3.9 Values of electron gain enthalpy ( $\text{kJ mol}^{-1}$ ) of some representative elements (For information only)**

Group 1	( $\Delta_{\text{eg}}\text{H}$ )	Group 16	( $\Delta_{\text{eg}}\text{H}$ )	Group 17	( $\Delta_{\text{eg}}\text{H}$ )	Group 18	( $\Delta_{\text{eg}}\text{H}$ )
						He	+48
Li	-60	O	-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	I	-295	Xe	+77
Cs	-46	Po	-174	At	-270	Rn	+68

Now to make the firm discussion about electron gain enthalpy stronger we have to get solution of the question- 'among P, S, Cl and F which one has the most negative and which one has the least negative electron gain enthalpy?' Generally, in a period, when we move from left to right then the electron gain enthalpy becomes more negative. So the order of negative value of electron gain enthalpy for P, S and Cl elements of third period is  $\text{P} < \text{S} < \text{Cl}$ . F and Cl are members of group 17. Added new electron takes place in 2p orbital in F and in 3p orbital in Cl. As we know, repulsion between electron-electron in 2p orbital is higher than in 3p orbital. So the electron can enter easily in 3p orbital. Therefore, order of negative value of electron gain enthalpy is  $\text{F} < \text{Cl}$ . Thus, Cl has the most negative electron gain enthalpy and P has the least negative electron gain enthalpy.

**3.8.1.5 Electronegativity :** The ability of an atom in a chemical compound to attract shared electrons of bond towards itself is called electronegativity. It is not possible to measure its absolute value, unlike ionization enthalpy and electron gain enthalpy. Although many numerical scales of electronegativity of elements have been developed such as Pauling Scale, Mulliken-Jaffe Scale and Alfred Rochow Scale. Out of these, the Pauling scale is widely used. In 1922, an American scientist Linus Pauling assigned arbitrarily the value for electronegativity. He considered to have studied the concept of

ability to attract electrons for fluorine and suggested the value of 4.0 to fluorine. As an another option, accepting the electronegativity of lithium atom as unity, the value of electronegativity of F becomes 4.0 which supports the electronegativity concept of Pauling. As we have understood previously, when the atomic radius of atom increases then, attraction of electrons of outermost orbit towards nucleus decreases. On the contrary when the atomic radius decreases then the attraction of electrons of outermost orbit towards nucleus increases. This means that the electronegativity of atom (small size) having small atomic radius is higher than that of the atom (big size) having more atomic radius. Thus in a period when we move from left to right then the electronegativity increases as the atomic radius of atom decreases (Table 3.10 (a)).

**Table 3.10 (a) Electronegativity values (on Pauling scale) in periods**

<b>Atom (Period II)</b>	Li	Be	B	C	N	O	F
<b>Electro-negativity</b>	1.0	1.5	2.0	2.5	3.0	3.5	4.0
<b>Atom (Period III)</b>	Na	Mg	Al	Si	P	S	Cl
<b>Electro-negativity</b>	0.9	1.2	1.5	1.8	2.1	2.5	3.0

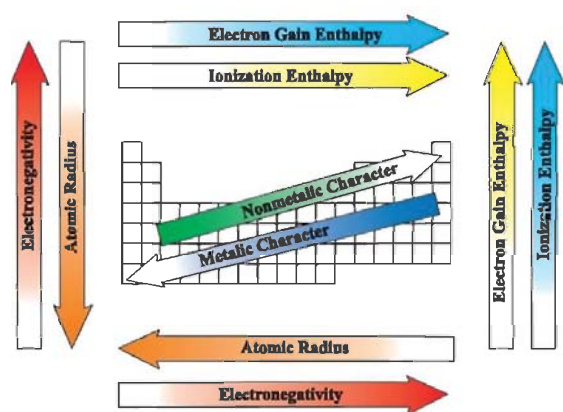
In a group when we move from top to bottom the electronegativity decreases because atomic radius of the atom increases.

**Table 3.10 (b) Electronegativity values (on Pauling scale) in group**

Atom (Group 1)	Electro-negativity	Atom (Group 17)	Electro-negativity
Li	1.0	F	4.0
Na	0.9	Cl	3.0
K	0.8	Br	2.8
Rb	0.8	I	2.5
Cs	0.7	At	2.2

We know that the non-metallic elements have strong tendency to gain electron. Thus, there is a direct relation of electronegativity with the non-metallic properties. In other words, there is an inverse relation of electronegativity with the metallic properties. Thus, in a period when we move from left to right the non-metallic character increases and the electronegativity also increases. In a group when we move from top to bottom the electronegativity and non-metallic properties decrease. So, in a period when we move from left to right the electropositivity decreases and electronegativity increases.

The periodic trends in physical properties of elements are shown in brief in Fig. 3.7.



The periodic trends in physical properties of elements in the periodic table

Figure 3.7

### 3.8.2 Periodic Trends in Chemical

**Properties :** Most of the trends in chemical properties of elements such as diagonal relationship, inert pair effect, effect of lanthanoid contraction etc. will be explained along with the discussion of each group in later units. Here,

we will study about the valence or state of oxidation elements.

**3.8.2.1 Periodicity of Oxidation State or Valence :** The valence is the important property of an element. It can be easily understood by electronic configuration of an element. The valence of representative elements is usually equal to the number of electrons in the outermost orbitals and/or equal to eight minus the number of outermost electrons. This can be mentioned below :

Group	1	2	13	14	15	16	17	18
Number of valence electron	1	2	3	4	5	6	7	8
valence	1	2	3	4	3,5	2,6	1,7	0,8

The word oxidation state is also used for valence. Let us consider two compounds containing oxygen- $\text{OF}_2$  and  $\text{Na}_2\text{O}$ . The order of electronegativity of elements present in these compounds is  $\text{F} > \text{O} > \text{Na}$ . Electronic configurations of outermost orbits of O, F, and Na are  $2s^2 2p^4$ ,  $2s^2 2p^5$  and  $3s^1$  respectively. In  $\text{OF}_2$ , fluorine atom shares one electron with oxygen atom, fluorine being more electronegative element contains  $-1$  oxidation state whereas O atom shares two electrons with fluorine atoms, so oxidation state of O is  $(+2)$ . In  $\text{Na}_2\text{O}$  oxygen being more electronegative accepts two electrons from two sodium atoms. Because of this its oxidation state is  $(-2)$ . On the other hand sodium atom loses one electron from 3s orbital which shows its  $(+1)$  oxidation state. Thus the oxidation state of an element in a particular compound is the charge acquired by its atoms from other atoms on the basis of electronegativity consideration in molecule. Table 3.11 shows some periodic trends in valence of elements.

**Table 3.11 Periodic trends in valence of elements  
(as formulae of compounds of elements)**

Group	1	2	13	14	15	16	17
Formula of hydride	LiH		$B_2H_6$	$CH_4$	$NH_3$	$H_2O$	HF
	NaH	$CaH_2$	$AlH_3$	$SiH_4$	$PH_3$	$H_2S$	HCl
	KH			$GeH_4$	$AsH_3$	$H_2Se$	HBr
				$SnH_4$	$SbH_3$	$H_2Te$	HI
Formula of oxide	$Li_2O$	MgO	$B_2O_3$	$CO_2$	$N_2O_3, N_2O_5$	$SO_3$	
	$Na_2O$	CaO	$Al_2O_3$	$SiO_2$	$P_4O_6, P_4O_{10}$	$SeO_3$	$Cl_2O_7$
	$K_2O$	SrO	$Ga_2O_3$	$GeO_2$	$As_2O_3, As_2O_5$	$TeO_3$	
		BaO	$In_2O_3$	$SnO_2$	$Sb_2O_3, Sb_2O_5$		
				$PbO_2$	$Bi_2O_3$		

Now let us see, what would be the formula of compound formed from Si and Br ? The answer is  $SiBr_4$  because Si is the member of group 14 so its valence becomes (+4) and Br is the member of group 17 so its valence becomes (-1). Similarly the

formula of compound formed from Al and S becomes  $Al_2S_3$  because the valence of Al is (+3) and that of S is (-2).

We shall study later on about the transition and innertransition elements containing different valence or oxidation states.

### SUMMARY

In early 1800 several efforts were made to classify newly invented and existing elements of the world. In which Dobereiner, A.E.B. De Chancourtois, John Newlands, Lothar Meyer and Mendeleev contributed elegantly. Modern periodic table exists on the basis of atomic number which is developed from the Mendeleev's old periodic table considering the atomic weight as the fundamental property. Modern periodic table is classified into 18 groups and 7 periods. Because of the same number of electrons in outermost orbit of element of the same group, they show similar chemical properties. Elements of same period have same quantum number (n) of outermost orbit. Classification of elements of periodic table in s, p, d, and f- blocks is based on filling of last electron in that orbital of element. Elements of s and p blocks are known as representative elements or main group elements; d-block elements as transition elements and; f-block elements as innertransition elements. When we move from left to right in a period in the periodic table, then, due to increase in atomic number of element, ionization enthalpy, negative value of electron gain enthalpy and value of electronegativity increase. On the other hand atomic radius and metallic property decrease. When we move from top to bottom in a group, then, due to increase in atomic number of element, ionization enthalpy, negative value of electron gain enthalpy and value of electronegativity decrease. On the other hand atomic radius and metallic property increase.

### EXERCISE

#### 1. Select the proper choice from the given multiple choices :

- (1) By which name the Eka-silicon element is known after its discovery ?
- (A) Gallium (B) Germanium  
(C) Tellurium (D) Indium



- (2) Uub means.....  
 (A) element having atomic number 111 (B) element having atomic number 112  
 (C) element having atomic number 113 (D) element having atomic number 114
- (3) Who suggested "Mendelevium" name for the element having atomic number 101 ?  
 (A) Mendeleev (B) Dobereiner  
 (C) Seaborg (D) Newlands.
- (4) Which electronic configuration from the following would be the element of third period ?  
 (A)  $[\text{Ar}] 3d^7 4s^2$  (B)  $[\text{Ne}] 3s^2 3p^4$   
 (C)  $[\text{Ar}] 3d^1 4s^2$  (D)  $[\text{Ar}] 3d^3 4s^2$
- (5) Which group contains the element having electronic configuration  $[\text{Ne}] 3s^2 3p^2$  ?  
 (A) 14 (B) 15  
 (C) 16 (D) 17
- (6) Which electronic configuration from the following is of s- block element ?  
 (A)  $[\text{Ar}] 3d^2 4s^2$  (B)  $[\text{Ar}] 4s^2$   
 (C)  $[\text{Ar}] 3d^3 4s^2$  (D)  $[\text{Xe}] 4f^4 5d^0 6s^2$
- (7) Which elements have the electronic configuration  $ns^2 np^6$  in their outermost orbit ?  
 (A) Alkali metals (B) Transition metals  
 (C) Noble gases (D) Innertransition metals
- (8) What is not consistent about atomic radius ?  
 (A) Decreases when we move from left to right  
 (B) Increases when we go from top to bottom  
 (C) Increases with increase in principal quantum number  
 (D) Increases with increase in positive charge of nucleus
- (9) Which is true from the following ?  
 (A) Size of  $\text{Al}^{3+} <$  Size of Al (B) Size of  $\text{Al}^{3+} >$  Size of Al  
 (C) Size of  $\text{F}^- <$  Size of F (D) Size of  $\text{Na}^+ =$  Size of Na
- (10) For which element the highest shielding effect for outermost electron is observed ?  
 (A) Element of group 13 and period 2 (B) Element of group 13 and period 3  
 (C) Element of group 13 and period 4 (D) Element of group 13 and period 5
- (11) Which one is the descending order of atomic radius of elements of third period.  $\text{Na}(Z = 11)$ ,  $\text{Mg}(Z = 12)$ ,  $\text{Al}(Z = 13)$  and  $\text{Si}(Z = 14)$  ?  
 (A)  $\text{Si} > \text{Al} > \text{Mg} > \text{Na}$  (B)  $\text{Na} > \text{Mg} > \text{Al} > \text{Si}$   
 (C)  $\text{Na} < \text{Mg} < \text{Al} < \text{Si}$  (D)  $\text{Na} > \text{Al} > \text{Mg} > \text{Si}$
- (12) Which order is true with reference to size of species ?  
 (A)  $\text{Pb} < \text{Pb}^{2+} < \text{Pb}^{4+}$  (B)  $\text{Pb}^{4+} > \text{Pb}^{2+} > \text{Pb}$   
 (C)  $\text{Pb} > \text{Pb}^{2+} > \text{Pb}^{4+}$  (D)  $\text{Pb}^{2+} < \text{Pb} < \text{Pb}^{4+}$

**2. Write the answers of the following questions in brief :**

- (1) Who gave the law of triads ?
- (2) Mention the temporary name of element having atomic number 114.
- (3) Write the general electronic configuration of f-block elements.
- (4) Which periods are called very long periods ?
- (5) Which electronic configuration occurs in outermost orbit of element of group 17 and period 3 ?
- (6) Which one has the largest size among  $\text{Cr}^{2+}$ ,  $\text{Cr}^{3+}$ ,  $\text{Cr}^{6+}$ , and Cr ?

**3. Write the answers of the following questions :****(1) Give definitions :**

- |                         |                             |
|-------------------------|-----------------------------|
| (1) Modern periodic law | (8) Semimetal               |
| (2) Group               | (9) Covalent radius         |
| (3) Period              | (10) Metallic radius        |
| (4) Elements of s-block | (11) Ionization enthalpy    |
| (5) Elements of p-block | (12) Shielding effect       |
| (6) Elements of d-block | (13) Electron gain enthalpy |
| (7) Elements of f-block | (14) Electronegativity      |

**(2) Describe the efforts of following scientists about the classification of elements :**

- |                             |                      |
|-----------------------------|----------------------|
| (1) Johann Dobereiner       | (4) Lothar Meyer     |
| (2) A. E. B.de Chancourtois | (5) Dmitri Mendeleev |
| (3) John Newlands           |                      |

- (3) Why is the size of positive ion smaller than its parent atom ?
- (4) Why is the size of negative ion larger than its parent atom ?
- (5) Mention the oxidation state of Ba and O in the compound BaO formed from Ba(group 2) and (group 16).
- (6) Arrange in decreasing order of size of  $K^+$ ,  $Cl^-$ ,  $S^{2-}$ ,  $Ca^{2+}$ . Give reason.
- (7) Arrange in increasing order of size :  $Na^+$ ,  $Mg^{2+}$ ,  $Al^{3+}$ ,  $F^-$ ,  $O^{2-}$ ,  $C^{4-}$ . Give reason.
- (8) Out of Na and K which one has lower value of ionization enthalpy ? Why ?
- (9) Which element from Na, I, Cl and Ar can form stable negative ion by easily gaining one electron ?
- (10) Which element from Li, B, N, O, and C has the highest electronegativity ? Why ?

**4. Write the answers of the following questions in detail :**

- (1) Describe the history of development of periodic table.
- (2) Mention the outline of modern periodic table.
- (3) Discuss the classification of elements in s, p, d, and f- block.
- (4) Explain the change in atomic radii when we move from left to right in a period.
- (5) In a group when we move from top to bottom the atomic radii of elements increase. Explain giving reasons.
- (6) What is ionization enthalpy ? In a group when we move from top to bottom, the values of ionization enthalpies decrease. Explain giving reasons.
- (7) In a period, when we move from left to right, the values of ionization enthalpies increase. Explain giving reasons.
- (8) Why the value of third ionization enthalpy of element is more than that of the second ionization enthalpy which is more than that of first ionization enthalpy ?
- (9) What is electron gain enthalpy ? The values of electron gain enthalpy become more negative as we move from left to right in the period ? Explain giving reasons.
- (10) In the same group with the increase in atomic number the values of electron gain enthalpies are found to be less negative ? Why ?
- (11) Explain the changes in electronegativity when we move from top to bottom in the same group and when we move from left to right in the same period.
- (12) What is an oxidation state ? Explain by giving an example.



## REDOX REACTIONS

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### 4.1 Introduction

#### 4.2.1 Oxidation and Reduction

#### 4.2.2 Oxidation - Reduction and Electron Transfer Methods

### 4.3 Competitive Electron Transfer Reaction

### 4.4 Oxidation Number

#### 4.4.1 Rules for the calculation of Oxidation Number

#### 4.4.2 Calculation of Oxidation Number

#### 4.4.3 Oxidation Number and Redox Reaction

### 4.5 Oxidation Number and Nomenclature

### 4.6 Equation of Redox Reaction

### 4.7 Balancing of Redox Reaction Equation

#### 4.7.1 Balancing of Redox Reaction equation using Oxidation Number Method.

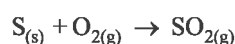
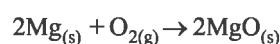
#### 4.7.2 Balancing of Redox Reaction equation using Half Reaction equation Method

### 4.1 Introduction

In chemical reaction when there is oxidation reaction, there is always reduction reaction. Hence, study of redox reactions are essential. Many phenomena are included in chemistry. One of them is the change from one phenomenon to the other one, substance is converted to the other substance at the end of chemical reaction. One of them is redox reaction. Redox reactions are related to physical and biological phenomena. The wider use of this reaction is found in different fields like pharmaceutical science, biology, industrial metallurgy, agriculture etc. The redox reactions are included in obtaining household energy in the use of different types of fuels for commercial purpose, production of compounds like caustic soda. In recent time hydrogen as an economy energy source and also holes in ozone layer are notable redox reactions.

#### 4.2.1 Oxidation and Reduction :

Basically the word oxidation is used to describe addition of oxygen to the substance. Many substances combine with oxygen because of its presence in atmosphere (nearly 20%). Because of this reason generally they are found in the form of oxides, the following reactions indicate oxidation :



In the above reactions, by addition of oxygen to  $\text{Mg}_{(s)}$  and  $\text{S}_{(s)}$  oxidation reactions take place.