587Periodic Classification of Elements

Concept

Early attempts of classification

Introduction

You must have visited a library. There are thousands of books in a large library. In spite of this, if you ask for a particular book, the library staff can locate it easily. How is it possible? In library, the books are classified into various categories and sub-categories. They are arranged on shelves accordingly. Therefore, location of books becomes easy.

Same is the story with chemical elements. A large number of elements and compounds are known today. But a systematic classification of these elements has made their study possible and easy. The well organized and tabulated classification of elements, as we know it today, is called the **periodic table.** It not only helps to locate, identify and characterize the element and its properties but also points out the directions in which new investigations is made.

Early Attempts

In 18th century, the number of elements was limited. In 19th century, scientists began to seek ways to classify elements because of their rapidly increasing number. They started recognizing patterns in properties and began to develop classification schemes. Some such early attempts of classification are described below.

(i) Dobereiner's Triads

In 1817, a German Chemist, Dobereiner observed that certain elements which had similar chemical properties, could be grouped together. When these elements were arranged in increasing order of their atomic masses, they generally occurred in groups of three. These groups were called **triads**. He noticed that the atomic mass of the middle element of the triad was the arithmetic mean of the other two elements of the triad. This was known as the **Dobereiner's law of triads**. It is stated as follows.

When elements are placed in order of ascending order of atomic masses, groups of three elements having similar properties are obtained. The atomic mass of the middle element of the triad is equal to the mean of the atomic masses of the other two elements of the triad.

Examples of Dobereiner's Triads

In the alkali metal group, consider the elements Li, Na and K. All these elements are metals. They are highly reactive and they show valency 1. The Dobereiner's triad for alkali metal group can be shown as follows.

Element	Symbol	A (Atomic Mass)
Lithium	Li	7
Sodium	Na	23
Potassium	K	39

From the Dobereiner's law of triads, the atomic mass of the middle element, in this case Na, should be the arithmetic mean of Li and K.
Thus arithmetic mean of Li and K = (7 + 39) / 2 = 23.
It can be seen that arithmetic mean of atomic masses of Li and K = atomic mass of Na

Now consider elements in the halogen group viz. Cl, Br and I. All these elements are non-metallic, they are very reactive and form acids with water. They have valency 1. Due to their similar chemical properties, these three elements formed another Dobereiner's triad. So let us see whether Cl, Br and I obey Dobereiner's law of triads.

Element	Symbol	A (Atomic Mass)
Chlorine	Cl	35.5
Bromine	Br	80
Iodine	I	127

For Dobereiner's law to be valid, A (Br) = A(Cl) + A(I) / 2 = (35.5 + 127) /2 = 81.2
The actual atomic mass of Br is 80.
Thus the atomic mass of the middle element of the triad is nearly equal to the arithmetic mean of the atomic masses of the other two elements of the triad. Hence the Dobereiner's law of triads holds true for halogen triad also.

Drawbacks of Dobereiner's law of triads

(i) All the then known elements could not be arranged in the form of triads.
(ii) For very low mass or for very high mass elements, the law was not holding good. Take the example of F, Cl and Br. Atomic mass of Cl is not an arithmetic mean of atomic masses of F and Br.
(iii) As the technique improved for measuring atomic masses accurately, the law was unable to remain strictly valid.

Advantage of Dobereiner's law of triads

The only advantage of Dobereiner's research was that it made chemists look at elements in terms of groups of elements with similar chemical and physical properties. This eventually led to rigorous classification of elements and the modern periodic table of elements , as we know, it was discovered.

Activity 1 – Suppose two elements have atomic mass 8 and 40 respectively. Assuming that Dobereiner's law of triad is true, find the name of the middle element and its atomic mass.
Similarly, select one element from the second or third period of the periodic table. Pick up one element just above and one element just below it. Using their atomic weights, verify whether Dobereiner's law of triads is true.

(ii) Newlands' Octaves

In 1864, an English chemist Newlands observed that every eighth element showed similar physical and chemical properties when the elements were placed in the increasing order of their atomic masses. This was called as **Newlands law of octaves.** It was due to its similarity with musical notes that it was called the law of octaves where, after seven different notes the eighth note is repetition(harmonic) of the first one . Newlands law of octaves is stated as follows.

> When elements are arranged in the increasing order of their atomic masses, the eighth element resembles the first in physical and chemical properties.

Newlands arranged the elements then known in the following manner.

Li	Be	B	C	N	O	F	
Na	Mg	Al	Si	P	S	Cl	
Do	***Re***	***Me***	***Fa***	***So***	***La***	***Ti***	***Western notes***

Sa	Re	Ga	Ma	Pa	Dha	Ni	Indian notes

In his arrangement, a row of elements had seven elements and the eighth element fell under the first element. In those days, the number of elements was very limited and noble gases were not known.

Let us see the elements in the first column. Li is the first element. The eighth element after Li is Na. Similarly, the eighth element after Na is K. So according to Newlands law of octaves, we should expect the elements Li, Na and K to have similar chemical properties. This they do have. For example, all these elements are metallic, highly reactive and show a valency 1. They are known as **alkali metals**.

Next, if we take Be as the first element, the eighth element is Mg. If we continue in a similar fashion, the eighth element is Ca. According to Newlands law, Be, Mg and Ca should display similar chemical properties. They do. These elements are metallic in nature, their oxides are alkaline in nature and they show a valency 2. They are known as **alkaline earth metals.**

Similarly, Cl happens to be the eighth element after F and Br happens to be the eighth element after Cl.The elements F, Cl and Br form a group of **halogens**. They obey Newlands law of octaves.

Drawbacks of Newlands law of octaves

(i) It was not valid for elements beyond calcium.
(ii) When new elements like noble gases were discovered, his table had no place for them.

Advantage of Newlands law of octaves

The most important contribution in the process of classification of elements was the periodicity Newlands saw in every eighth element. The modern periodic table, that we shall study later, drew heavily from the concept of periods of eight. Also it must be noted that Dobereiner’s triads occurred in the octaves of Newlands.

(iii) Lothar Meyer’s Atomic volume curve

A German Chemist Lothar Meyer, in 1869, was studying the physical properties of elements along with their valence states. He made a table that contained a preliminary tabulation of 28 elements. The table showed how the integral valence changed as the atomic mass of elements increased.

Lothar Meyer considered the volume taken up by fixed weights of the various elements. Under such conditions, each weight contained the same number of atoms of its particular element. (Avogadro’s number)This meant that the ratio of the volumes of the various elements was equal to the ratio of the volumes of single atoms of the various elements. Thus Lothar Meyer could determine the atomic volumes of elements.

If the atomic volumes of the elements were plotted against the atomic weight, a series of peaks were produced. The peaks had alkali metals – Na, K, Rb and Cs. Each fall and rise to a peak, corresponded to a period like the waves. In each period, a number of physical properties other than atomic volume also fell and rose, such as valence and melting point. The figure shown below shows Lothar Meyer's atomic volume curve where atomic volume is plotted against atomic mass of an element.

Hydrogen, the first in the list of elements is a special case and can be considered as making up the first period all by itself.

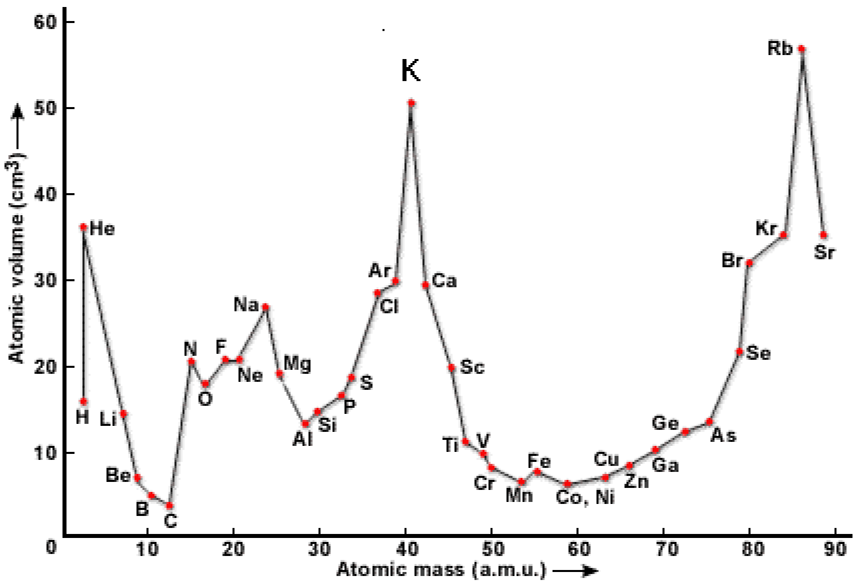

Figure 1 – Lothar Meyer's atomic volume curve

The second and the third period in Lothar Meyer's table included seven elements each and duplicated Newlands law of octaves as follows.

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl

However, the next wave had more than seven elements. The third wave had about 17 to 18 elements. This clearly showed where Newlands law had failed. One could not force the law of octaves to hold strictly throughtout the table of elements, with seven elements in each row. After the first two periods, the length of the period had to be longer. The atomic volume curve has following features.

(i) Alkali metals like Na, K Rb that have similar properties, occur as peaks of the curve.
(ii) Halogen elements like F, Cl, Br that have similar properties, occur at the rising or the ascending part of the curve.

(iii) Noble gases (now included) like Ne, Ar, Kr, that have similar properties, occur just before the alkali elements.
(iv) H and He seem to be the exception to all rules.

Advantages of Lothar Meyer's classification

(i) Lothar Meyer's main contribution was his recognition of periodic behaviour i.e. a repeating pattern in physical properties like melting points, boiling points, valencies and so on.
(ii) It led to the confirmation of periodicity in chemical properties as well.

Activity 2 – Collect data of melting points of elements of as many elements as you can. Draw a graph of melting point as a function of atomic mass. See whether you get curve similar to Lothar Meyer's atomic volume curve.

Test your understanding

1) What were the criteria used by the scientists in the early attempts of classification of elements ?
2) Mention any three attempts of classification of elements prior to Mendeleev's classification. Why were these attempts not successful ?
3) How did the early attempts of classification help to arrive at today's periodic table ?

Concept

Mendeleev's Classification

Working on the researches of the earlier scientists, a Russian chemist, Dimitri Mendeleev presented in 1869 a much bolder and scientifically useful classification of elements. Mendeleev saw that there is a periodicity occurring in the physical and chemical properties, if the elements were arranged in order of their atomic weights. His classification in tabular form is called **Mendeleev's periodic table**. It gave order to the large amount of data available for all the elements. In Mendeleev's periodic table, about 60 to 70 elements were accommodated. Mendeleev made a statement which later came to be known as **Mendeleev's periodic law**. It is stated as follows.

The properties of elements are the periodic functions of their atomic weights.

Modified form of Mendeleev's periodic table is shown below.

PERIODIC TABLE (*Modified form of Mendleeff's Table*)

PERIODS	Group :	I	II	III	IV	V	VI	VII	VIII	Zero
	Oxide: Hydride:	R_2O RH	RO RH_2	R_2O_3 RH_3	R_2O_4 RH_4	R_2O_5 RH_3	RO_3 RH_2	R_2O_7 RH	RO_4 Transition Traids	Noble gases
		A B	A B	A B	A B	A B	A B	A B		
1		H 1 (At. No.) 1.008 (At. Wt.)								He 2 4.0026
2		Li 3 6.939	Be 4 9.012	B 5 10.811	C 6 12.011	N7 14.007	O 8 15.999	F 9 18.998		Ne 10 20.183
3		Na 11 22.99	Mg 12 24.312	Al 13 26.981	Si 14 28.086	P 15 30.974	S 16 32.06	Cl 17 35.453		Ar 18 39.948
4	First series	K 19 39.102	Ca 20 40.08	Sc 21 44.96	Ti 22 47.90	V 23 50.94	Cr 24 51.99	Mn 25 54.939	Fe 26 Co 27 Ni 28 55.85 58.93 58.71	
	second series	Cu 29 63.54	Zn 30 65.37	Ga 31 69.72	Ge 32 72.59	As 33 74.92	Se 34 78.96	Br 35 79.909		Kr 36 83.80
5	First series	Rb 37 85.47	Sr 38 87.62	Y 39 88.905	Zr 40 91.22	Nb 41 92.906	Mo 42 95. 94	Tc 43 (99)	Ru 44 Rh 45 Pd 46 101.07 102.91 106.4	
	second series	Ag 47 107.87	Cd 48 112.40	In 49 114.82	Sn 50 118.69	Sb 51 121.75	Te 52 127.60	I 53 126.9014		Xe 54 131.30
6	First series	Cs 55 132.90	Ba 56 137.34	*Rare Earths 57-71	Hf 72 178.49	Ta 73 180.948	W 74 183.85	Re 75 186.2	Os 76 Ir 77 Pt 78 190.2 192.2 195.09	
	second series	Au 79 196.97	Hg 80 200.59	Tl 81 204.37	Pb 82 207.19	Bi 83 208.98	Po 84 (210)	At 85 (210)		Rn 86 (222)
7		Fr 87 (223)	Ra 88 (226)	† Actinide Elements 89-103	Ku 104	Ha 105				

*** Lanthaandie Elements** *(Rare Earth Series)* (La 57 Ce 58 Pr 59 Nd 60 Pm 61 Sm 62 Eu 63 Gd 64 Tb 65 Dy 66 Ho 67 Er 68 Tm 69 Yb 70 Lu 71
(138.91 140.12 140.91 144.24 (147) 150.35 151.96 157.25 158.92 162.50 164.93 167.26 168.93 173.04 174.97

† Actinide Series (Ac 89 Th 90 Pa 91 U 92 Np 93 Pu 94 Am 95 Cm 96 Bk 97 Cf 98 Es 99 Fm 100 Md 101 No 102 Lr 103
((227) 232.04 (231) 238.3 (237) (244) (243) (245) (247) (249) (254) (253) (256) (253) (257)

Figure 2 - Modified form of Mendeleev's periodic table

A periodic function is the one which repeats itself after a certain interval. Thus, according to the periodic law, the chemical and physical properties of elements repeat themselves after certain intervals when they are arranged in the increasing order of their atomic mass. Mendeleev believed that atomic mass of elements was the most fundamental property and arranged them in its increasing order in horizontal rows till he encountered an element which had properties similar to the first element. He placed this element below the first element and thus started the second row elements. Proceeding in this manner he could arrange all the known elements according to their properties and thus created the first periodic table.

Main features of Mendeleev's periodic table

In Mendeleev's periodic table, elements are arranged in tabular form in rows and columns. Their description is given below.

(i) The horizontal rows present in the periodic table are called **periods**. There are seven periods in the periodic table. These are numbered from 1 to 7 (Arabic numerals)

(ii) Properties of elements in a particular period show regular gradation (i.e. increase or decrease) from left to right.

(iii) The vertical columns present in it are called **groups**. There are nine such columns and they are numbered from I to VIII and Zero (Roman numerals)
(The zero group is now included)

(iv) Groups I to VII are subdivided into A and B subgroups. Group Zero and VIII do not have any subgroups.

(v) All the elements in a particular group are chemically similar in nature. They show regular gradation in their physical properties and chemical reactivities.

Merits of Mendeleev's periodic classification

(i) Mendeleev's was the first classification which included all the elements known at that time and even the elements discovered later.
(ii) Mendeleev had left some gaps in the periodic table. These gaps were for the elements yet to be discovered. E.g. he proposed the existence of an unknown element that he called eka-aluminium. The element gallium was discovered four years later and its properties matched very closely with the predicted properties of eka-aluminium.
(iii) Scientists still use Mendeleev's method to predict the properties of undiscovered elements by looking vertically across groups and horizontally across periods of elements.

Defects in Mendeleev's periodic classification

(i) Hydrogen resembles alkali metals (it forms H^+ ion like Na^+ ion) as well as halogens (it forms H^- ion like Cl^- ion). Therefore, it could be neither placed with alkali metals (Group I) nor with halogens (Group VII).

(ii) Different isotopes of same element have different atomic masses. Therefore, each one of them should be given a different position in the periodic table. On the other hand, because they are chemically similar, they had to be given same position. So there was no proper position to isotopes in the periodic table.

(iii) At certain places, an element of higher atomic mass has been placed before an element of lower atomic mass. For example, argon (39.9) is placed before potassium (39.1),tellurium (127.6) was placed before iodine (126.9). This disturbed the principle of increasing atomic mass.

Activity 3 - Prepare a few teams of 4 to 5 students each. Your teacher will give each team a set of cards. Each card in the set will contain information about an element. Your challenge will be to arrange the cards into a two dimensional table in some way that makes sense to you and the other members of your team. When you have finished arranging your elements, explain to the other students why did you arrange the element-cards, the way you did.

Activity 4 - In Mendeleev's periodic table, 'similar elements were placed in dissimilar groups and dissimilar elements were placed in similar groups'. Find out such elements and justify this statement.

Test your understanding

1) What is Mendeleev's periodic law ? Why was it modified?
2) Who was responsible to modify Mendeleev's periodic law ?
3) How many columns and rows were present in Mendeleev's periodic table ?
4) State any two merits and any two drawbacks of Mendeleev's periodic table.

Concept

Modern Classification

Moseley, an English physicist in 1913 discovered that atomic number is a fundamental property of an element and not atomic mass. Atomic number Z of an element is the number of protons in the nucleus of the atom. The number of electrons in the neutral atom is also equal to its atomic number. This discovery changed the whole perspective about elements and their properties to such an extent that a need was felt to change the periodic law also. Mendeleev's periodic law was then modified and now it is called the **modern periodic law**. It is stated as follows.

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

After the change in the periodic law, many changes were suggested in the periodic table. Many versions of the periodic table are in use but the one which is more commonly used is called **Long Form of the Periodic Table.** One version, how the long form should look like, is given below.

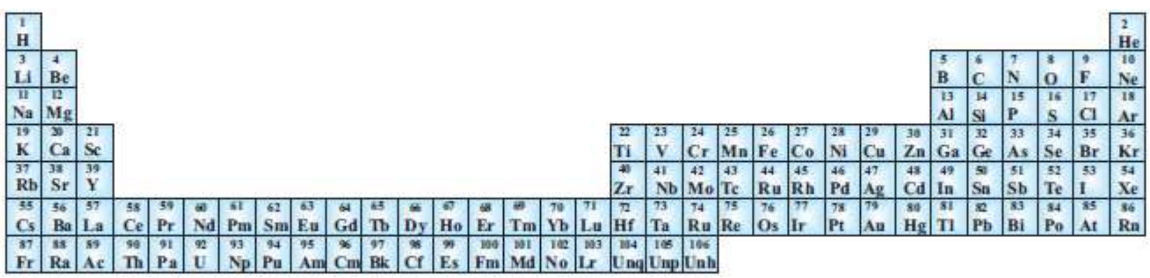

Fig.3 – The extended long form of the periodic table

Following figure shows the regularly used long form of the periodic table.

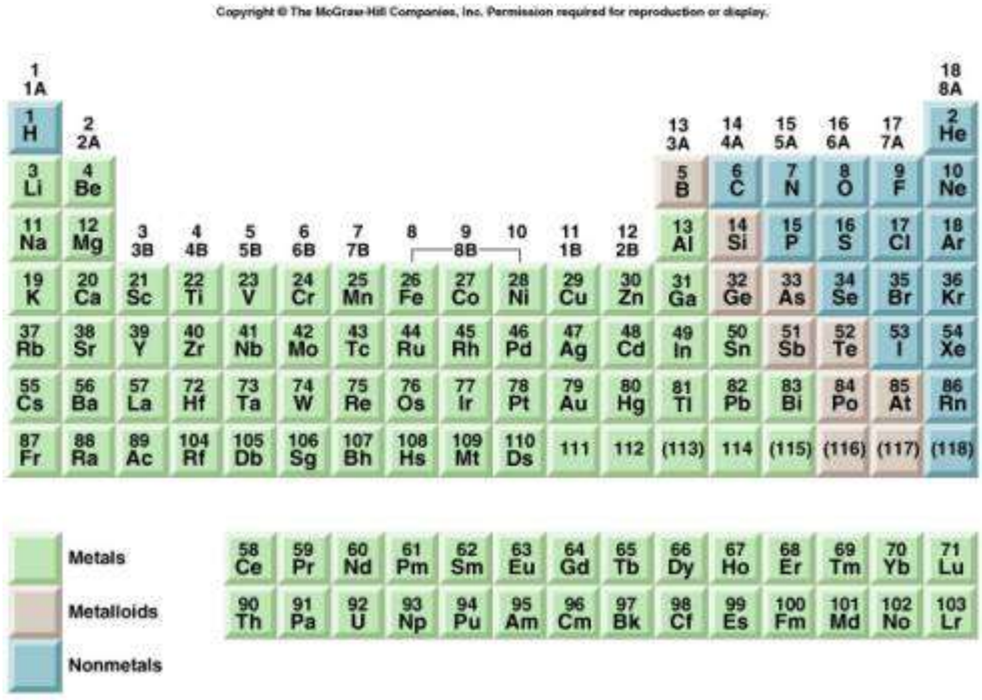

Figure 4 – Long form of the periodic table

Features of the long form of the periodic table

Groups

(i) There are 18 vertical columns in the long form of the periodic table. Each column is called a **group**. The groups have been numbered from 1 to 18 (in Arabic numerals) from left to right. Group 1 on the extreme left contains alkali metals and group 18 on the extreme right contains noble gases.
The International Union of Pure and Applied Chemistry (**IUPAC**) has recommended that groups be numbered from 1 to 18 instead of IA,IIA, IB, IIB etc. It may be noted that **Roman numerals** were used in the old system and **Arabic numerals** are used in IUPAC system. Thus Mg belongs to group IIA according to old system while it belongs to group 2 according to IUPAC system. Similarly, oxygen belongs to VIA group according to old system while it belongs to group 16 according to IUPAC system.

(ii) Elements in a given group have similar electronic configuration and have same number of valence electrons. For example, in alkali metals (Group 1) and in halogens (Group 17) as we move down the group, more and more shells are added.

Group 1		Group 17	
Element	**Electronic Configuration**	**Element**	**Electronic Configuration**
Li	2,1	F	2,7
Na	2,8,1	Cl	2,8,7
K	2,8,8,1	Br	2,8,8,7
Rb	2,8,18,8,1	I	2,8,18,18,7

All elements of group 1 have only one valence electron. Li has electrons in two shells, Na in three , K in four and Rb has electrons in five shells. Similarly, all the elements of group 17 have seven valence electrons however the number of shells is increasing from 2 in F to 5 in I.

(iii) Elements in groups 1 and 2 on the left side and groups 13 to 17 on the right side of the periodic table are called **normal elements** or **representative elements**. Their outermost shells are incomplete. They are also called **main group elements**.

(iv) Elements in groups 3 to 12 in the middle of the periodic table are called **transition elements** (Although groups 11 and 12 elements, in a strict sense, are not transition elements). Their two outermost shells are incomplete. It should be noted that electrons are added to the last incomplete shell only in case of normal elements. In case of transition elements, electrons are added to incomplete inner shells.

(v) Group 18 on the extreme right side of the periodic table contains noble gases. Their outermost shell contains 8 electrons.

(vi) Elements below the main body of the periodic table are called **inner transition elements**. They contain **lanthanides** and **actinides**.

The 14 elements with atomic numbers 58 to 71 (Ce to Lu) are called **lanthanides**. They are placed along with the element lanthanum (La), atomic number 57 in the same position (group 3 in period 6) because of very close resemblance between them. However, for the sake of convenience, they are shown separately below the main periodic table.

The 14 elements with atomic numbers 90 to103 (Th to Lr) are called **actinides**. They are placed along with the element actinium (Ac), atomic number 89 in the same position (group 3 in period 7) because of very close resemblance between them. However, for the sake of convenience, they are shown separately below the main periodic table

Periods

(i) There are seven rows in the periodic table. Each row is called a **period**. The periods have been numbered from 1 to 7 (Arabic numerals).

(ii) In each period, a new shell starts filling up. The period number is also the number of shell which starts filling up in it. For example, in elements of 3rd period, the third shell (M shell) starts filling up as we move from left to right. The first element of this period, sodium Na (2,8,1) has only one electron in its valence shell (third shell) while the last electron of this period, argon Ar (2,8,8) has eight electrons in its valence shell. The gradual filling of the third shell has been shown below.

Element	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
Electronic Configuration	2,8,1	2,8,2	2,8,3	2,8,4	2,8,5	2,8,6	2,8,7	2,8,8

(iii) First period is the **shortest period** of all and it contains only two elements H and He.
(iv) Second and third periods are called **short periods** and they contain 8 elements each.
(v) Fourth and fifth periods are called **long periods** and they contain 18 elements each.
(vi) Sixth and seventh periods are **very long periods** containing 32 elements each.

Advantages of the long form of the periodic table

(i) The modern periodic table is based on atomic number which is a more fundamental property of an atom than atomic mass. The long form of modern periodic table is therefore free from main defects of Mendeleev's periodic table.
(ii) The table shows why elements in the same group display similar properties. The table also shows how and why the properties of elements differ in the same period.
(iii)The periodicity is related to electronic configuration. All chemical and physical properties are a manifestation of the electronic configuration of the elements. Hence the periodic table is very systematic and follows the fundamental electronic structure of the elements to classify them.

Defects in the long form of the periodic table

(i) The position of hydrogen is still not settled since it shows properties of both alkali metals as well as halogens.

(ii) Lanthanides and actinides are not accommodated in the main body of the periodic table.

Activity 5 – Draw the skeleton of the long form of the periodic table and draw a line of demarcation between metals and non- metals.

Activity 6 - Suppose you have found an element with atomic number 120. Suggest a position to this element in the modern periodic table and predict any two chemical properties of this element.

Test your understanding

1) What is the IUPAC nomenclature of groups in the long form of the periodic table? How is it different from that in Mendeleev's periodic table ?
2) Is classification of elements in the long form of the periodic table related with electronic configuration of elements? Justify your answer.
3) How many elements are present in the 6^{th} period of the long form of the periodic table.?
4) Name the group and period of the element having atomic number 21.
5) Draw the outline of long form of the periodic table and show the position of normal elements and transition elements in it.
6) Name any two defects in Mendeleev's periodic table which stand removed in the long form of the periodic table.
7) You are given an element which shows valency 2, it reacts with cold water producing hydrogen and it forms an ionic chloride. In which group of the periodic table will you place this element?

Concept

Periodicity in properties

We have learnt about the main features of the Modern Periodic Table. We have also learnt that in a period the number of valence electrons and the nuclear charge increases from left to right. It increases the force of attraction between them. In a group the number of filled shells increases and valence electrons are present in last shells. This decreases the force of attraction between them and the nucleus of the atom. These changes affect various properties of elements and they show gradual variation in a group and in a period and they repeat themselves after a certain interval of atomic number.

> The recurrence of properties after a regular interval is called periodicity.

A property which shows similarity as well as a gradation (gradual variation) across a period and within a group with regular interval of atomic number is called a periodic property. The distribution of electrons in the various shells determines the physical and chemical character of the element. It is observed that the properties of elements depend more on the arrangements of the electrons in the outermost shell (valence shell) and not on the inner shells. The similarity in properties of elements is due to the similarity in electronic configuration while gradation in properties is due to the gradually changing atomic volume of the elements. Let us see some periodic properties and their trends in the periodic table.

(i) Valency – The combining capacity of an element is called its valency. Usually, it is found with reference to oxygen, hydrogen or chlorine.

a) Variation of valency in a period - The number of valence electrons increases in a period. In normal elements, the valency increases from 1 to 8 across a period from left to right. It reaches 8 in group 18 elements (noble gases) which show practically no chemical activity under ordinary conditions and their valency is taken as zero. Valency of normal elements with respect to oxygen increases from 1 to 7 as shown below for elements of the third period. This valency is equal to the number of valence electrons or group number for group 1 and 2 or groups 13 to 17.

Group	1	2	13	14	15	16	17
Element	Na	Mg	Al	Si	P	S	Cl
No. of Valency e^-	1	2	3	4	5	6	7
Valency w.r.t. O	1	2	3	4	5	6	7
Formula of oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7

In the following table, for the elements of second period, it is seen that valency with respect to hydrogen and chlorine increases from 1 to 4 and then decreases to 1 (or 8 – group number) again.

Group	1	2	13	14	15	16	17
Element	Li	Be	B	C	N	O	F
No. of valency e$^-$	1	2	3	4	5	6	7
Valency w.r.t. H or Cl	1	2	3	4	3	2	1
Formula of hydride	LiH	BeH_2	BH_3	CH_4	NH_3	H_2O	HF
Formula of chloride	$LiCl$	$BeCl_2$	BCl_3	CCl_4	NCl_3	Cl_2O	ClF

b) Variation of valency in a group - Elements of any given group have the same number of valence electrons. Therefore, they have the same valency. For example, all alkali metals show valency 1. All halogens show valency 1. All alkaline earth metals show valency 2.

The general trend of valency in the periodic table can be given as follows.

> The valency of elements goes on increasing across a period from left to right and remains the same in a given group of elements in the periodic table.

(ii) Atomic radii - Atomic size determines many physical properties of elements such as density, melting and boiling point and so on. It is difficult to define the size of an atom exactly. Usually, atomic size is expressed in terms of atomic radii. To develop a physical picture, atomic radius is taken to be the distance between the nucleus and the outermost shell of its electrons. It is experimentally determined by measuring the internuclear distance between two covalently bound atoms of the same element. It is then defined as one-half the distance between the nuclei of two atoms when they are linked to each other by a single covalent bond. Different experimental methods have been used to determine the values of atomic radii. Hence there is a little variation in the reported values. Following figure shows the method of measurement of the atomic radius.

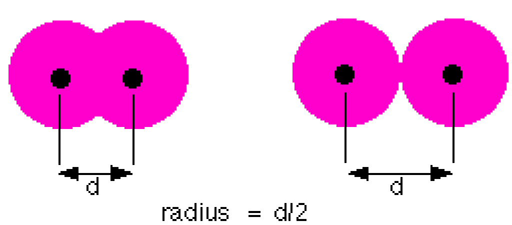

Figure 5 – Method of measurement of atomic radius

Following table shows the atomic radii of elements in pico meters in the periodic table.

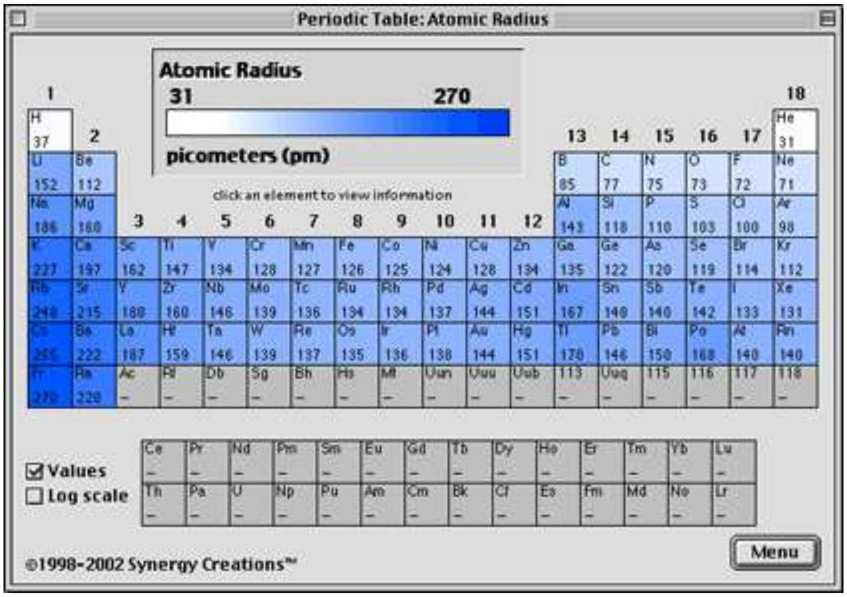

Figure 6 - Periodic table of atomic radii of elements

Following graph shows the variation of atomic radius as a function of atomic number.

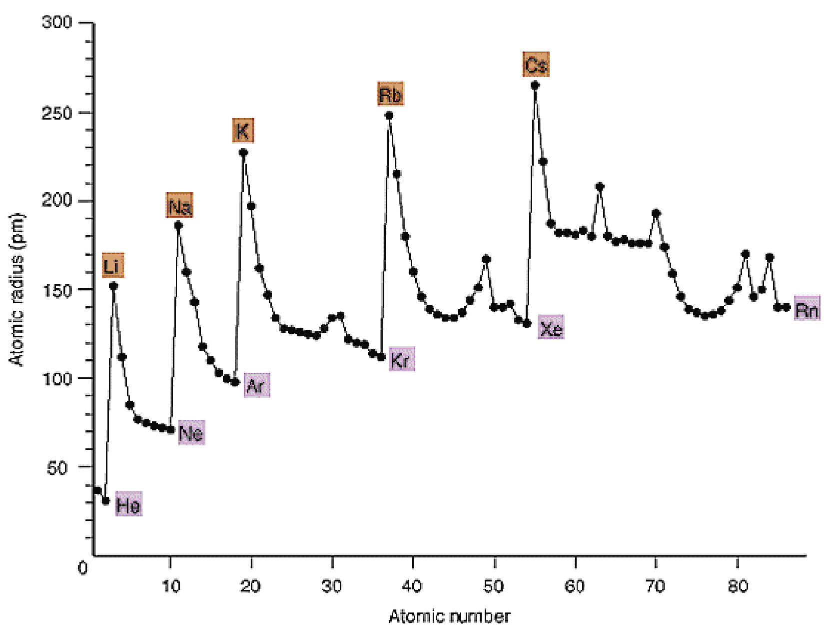

Figure 7 - Graph of atomic radius as a function of atomic number

Variation of atomic radii in a period - Atomic radii in picometers are given for the elements of the second and third periods below.

2nd period -	Li	Be	B	C	N	O	F
	155	112	98	91	92	73	72
3rd period -	Na	Mg	Al	Si	P	S	Cl
	190	160	143	132	128	127	99

Following figure shows how atomic radius changes across the periods 2 and 3.

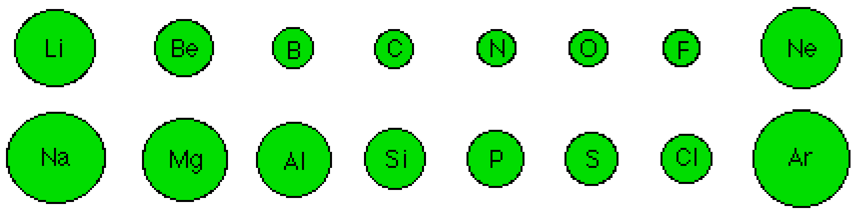

Figure 8 - Atomic radius decreases across the periods 2 and 3.

In a period, atomic radius generally decreases from left to right. It can be explained as follows. As we go from left to right, electrons are added, one at a time, to the same outermost shell. A proton is also added one at a time. The outermost electrons experience increasingly strong nuclear attraction, so the electrons come closer to the nucleus and more tightly bound to it. This results in decreasing the atomic radius.

We have to ignore the noble gas at the end of each period because noble gases do not form bond under normal conditions. Their van-der-Walls' radius has been shown in the above figure which is larger than its atomic or covalent radius.

Variation of atomic radii in a group – Atomic radii in picometers are given below for the alkali metals and halogens.

Element	Atomic radius	Element	Atomic radius
Li	155	F	72
Na	190	Cl	99
K	235	Br	114
Rb	248	I	133

In a group, atomic radius increases from top to bottom. It can be explained as follows. As we go down the group, filled shells go on increasing in number. The outermost shells go away from the nucleus. For example, lithium Li, has one filled shell while sodium Na, has two filled shells. The outermost shell is more away in sodium than in lithium from the nucleus. This decreases the attraction between the nucleus and the valence electrons. The outermost electrons go away from the nucleus and are loosely bound. This results in increasing the atomic radius.

The general trend of atomic radius in the periodic table can be given as follows.

> Atomic radius of an element goes on decreasing across a period from left to right and goes on increasing from top to bottom in a group in the periodic table.

(iii) Ionic radii – An ion is an atom or group of atoms that carries a positive or negative charge as a result of having lost or gained one or more electrons. Ionic radius is the radius of the ion. When an atom is converted to an ion, the size of the neutral atom changes. An anion is bigger than a neutral atom. This is because addition of one or more electrons increases repulsion among electrons and they move away from each other. On the other hand, a cation is smaller than the neutral atom. When one or more electrons are removed, the repulsive force between the remaining electrons decreases and they come a little closer. This is shown in the following figure.

Figure 9 – Cation is smaller than atom. Anion is bigger than atom.

The ionic radius is estimated from the distance between cations and anions that are adjacent in ionic crystals. Ionic radius also refers to the size of an atom after losing or gaining valence electrons. There is always some variation in the reported values of ionic radii because the methods used for the determination of ionic radii are different.

Following figure shows the ionic radii of some elements. The figures adjacent to ions are radii in picometers.

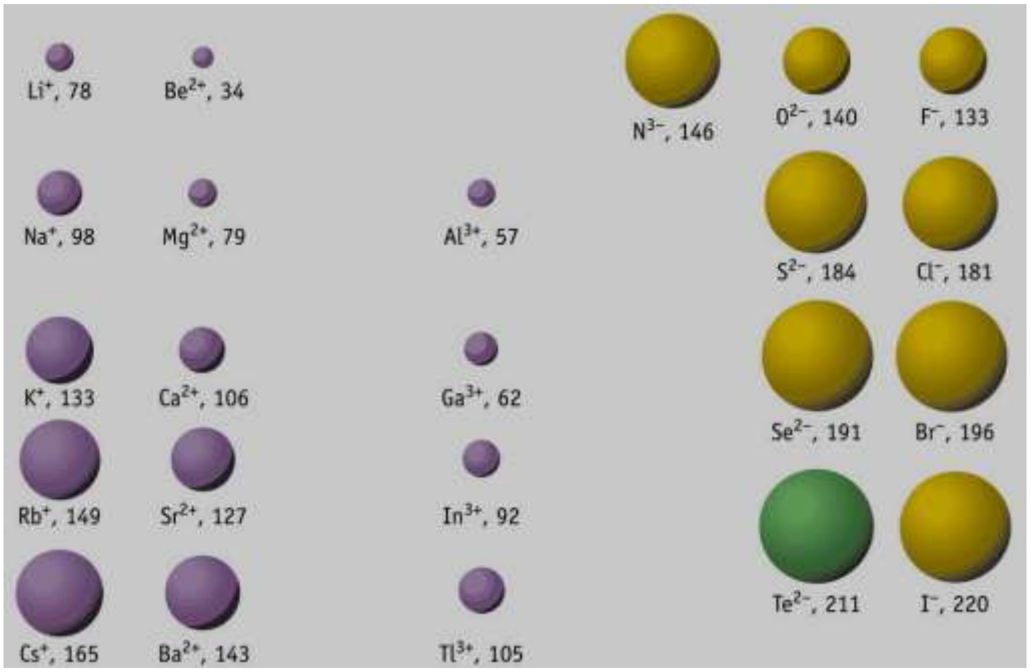

Figure 10 – Ionic radii of some elements

Variation of ionic radii across a period - The above figure shows the ionic radii of some elements. For metals and non-metals separate considerations are necessary. Among metals the cationic radius and among non – metals the anionic radius goes on decreasing across the period from left to right. For a given set of cations or anions, the number of electrons remain the same (isoelectronic ions) while atomic number goes on increasing. The electrons are attracted more to the nucleus hence the ionic radius goes on decreasing. If only cationic radius is considered, it goes on decreasing across the period from left to right continuously.

Variation of ionic radii within a group - The ionic radii of elements in a group are shown in the above figure. As we go down the group, ionic radii increase with increasing atomic number of the element. As we go down the group, the number of shells increases and valence electrons present in higher shells go away from the nucleus. For example, in lithium, the valence electrons are present in 2nd shell while in sodium, valence electrons are present in 3rd shell. Thus, the distance between nucleus and the outermost shell i.e. ionic radius goes on increasing from top to bottom in a group.

The general trend of ionic radius in the periodic table can be given as follows.

Ionic radius of an element goes on decreasing across a period from left to right and goes on increasing from top to bottom in a group in the periodic table.

(iv) Ionisation energy - The amount of energy required to remove an electron from a free gaseous isolated atom from its ground state to infinity to form a gaseous ion is called ionization energy. It is expressed in $kJmol^{-}$ or $kcalmol^{-}$. When expressed in eV per atom, it is called ionization potential. Ionisation energy is a measure of force of attraction between the nucleus and the outermost electron. Stronger the attraction, greater is the value of ionization energy.

If only one electron is removed, the ionization energy is known as **first ionization energy**. If second electron is removed, the ionization energy is called the **second ionization energy** and so on. The ionization energy goes on increasing for successive removal of electrons.

Following table shows the first ionization energy ($kJmol^{-1}$) of elements in the periodic table.

1314 **H** 2.1			Ionization Energies and Electronegativites														2368 **He** --
519 **Li** 1.0	900 **Be** 1.5												799 **B** 2.0	1088 **C** 2.5	1401 **N** 3.0	1036 **O** 3.5	1682 **F** 4.0	2076 **Ne** --
498 **Na** 0.9	736 **Mg** 1.2											578 **Al** 1.5	787 **Si** 1.8	1063 **P** 2.1	1000 **S** 2.5	1256 **Cl** 3.0	1519 **Ar** --
418 **K** 0.8	590 **Ca** 1.0	632 **Sc** 1.3	661 **Ti** 1.5	649 **V** 1.6	653 **Cr** 1.6	716 **Mn** 1.5	762 **Fe** 1.8	757 **Co** 1.8	736 **Ni** 1.8	745 **Cu** 1.9	904 **Zn** 1.6	578 **Ga** 1.6	782 **Ge** 1.8	1013 **As** 2.0	942 **Se** 2.4	1142 **Br** 2.8	1352 **Kr** --
402 **Rb** 0.8	548 **Sr** 1.0	636 **Y** 1.2	670 **Zr** 1.4	653 **Nb** 1.5	695 **Mo** 1.8	720 **Tc** 1.9	724 **Ru** 2.2	745 **Rh** 2.2	804 **Pd** 2.2	728 **Ag** 1.9	866 **Cd** 1.7	557 **In** 1.7	707 **Sn** 1.8	833 **Sb** 1.9	870 **Te** 2.1	1008 **I** 2.5	1172 **Xe** --
377 **Cs** 0.7	502 **Ba** 0.9	540 **La** 1.1	531 **Hf** 1.3	586 **Ta** 1.5	770 **W** 1.7	757 **Re** 1.9	841 **Os** 2.2	887 **Ir** 2.2	862 **Pt** 2.2	887 **Au** 2.4	1008 **Hg** 1.9	590 **Tl** 1.8	716 **Pb** 1.8	770 **Bi** 1.9	820 **Po** 2.0	-- **At** 2.2	1038 **Rn** --
-- **Fr** 0.7	510 **Ra** 0.9	678 **Ac** 1.1	-- **Ku** --														
1008 – Ionization energy C - symbol 2.5 - electro-negativity			665 **Ce** 1.1	557 **Pr** 1.1	607 **Nd** 1.1	-- **Pm** 1.1	540 **Sm** 1.1	548 **Eu** 1.1	594 **Gd** 1.1	649 **Tb** 1.1	657 **Dy** 1.1	-- **Ho** 1.1	-- **Er** 1.1	-- **Tm** 1.1	598 **Yb** 1.1	481 **Lu** 1.2	
			-- **Th** 1.3	-- **Pa** 1.5	385 **U** 1.7	-- **Np** 1.3	-- **Pu** 1.3	-- **Am** 1.3	-- **Cm** 1.3	-- **Bk** 1.3	-- **Cf** 1.3	-- **Es** 1.3	-- **Fm** 1.3	-- **Md** 1.3	-- **No** 1.3	-- **Lr** --	

Figure 11 – Periodic table of first ionization energies of elements

Following figure shows the variation of first ionization energy as a function of atomic number.

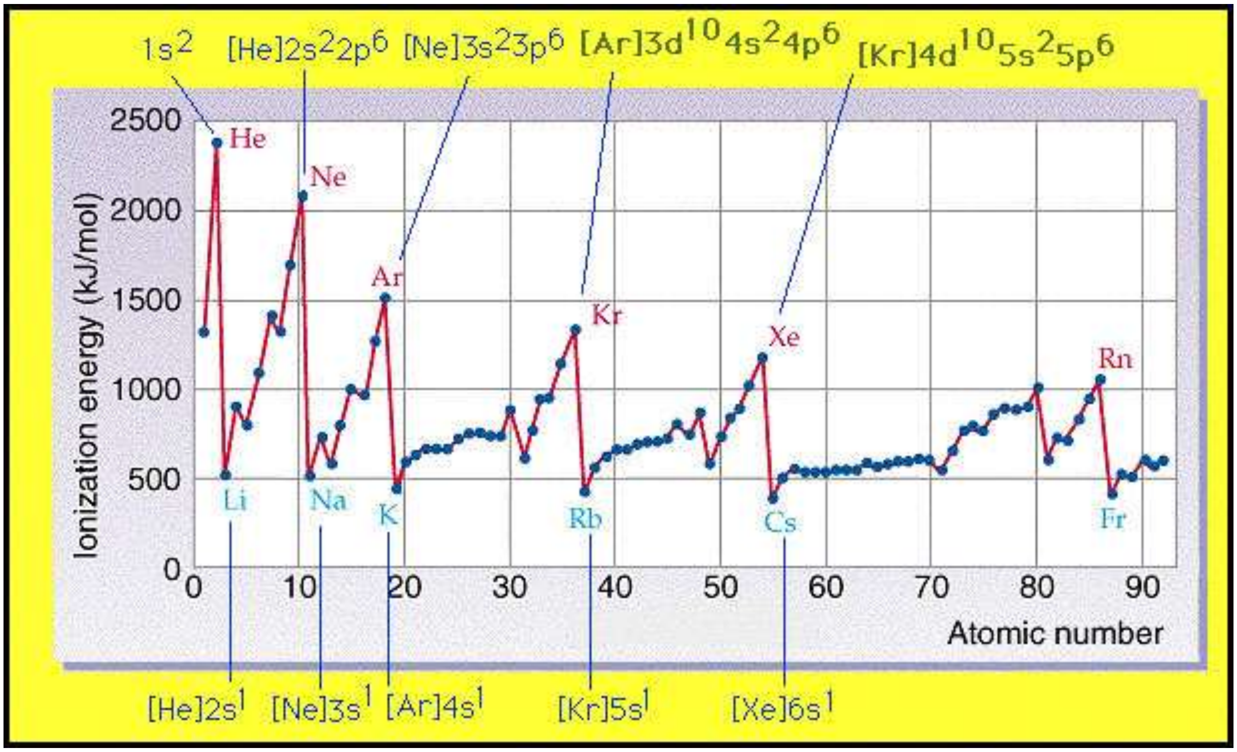

Figure 12 – Graph of first ionization energy as a function of atomic number of element

Variation of ionization energy in a period - As we go from left to right across a period, atomic size goes on decreasing and hence the force of attraction between the nucleus and valence electrons goes on increasing. As a consequence, the ionization energy goes on increasing from left to right in a period. This is clear from the table of first ionization energies (in $kJmol^{-}$) of elements of the second row elements.

Element	Li	Be	B	C	N	O	F	Ne
Ionisation energy	520	899	801	1086	1400	1314	1680	2080

The ionization energies of B and O are exceptionally low due to the stability achieved by the atom on losing electron. The graph shown below shows the variation of first ionization energy as a function of atomic number for the second row elements.

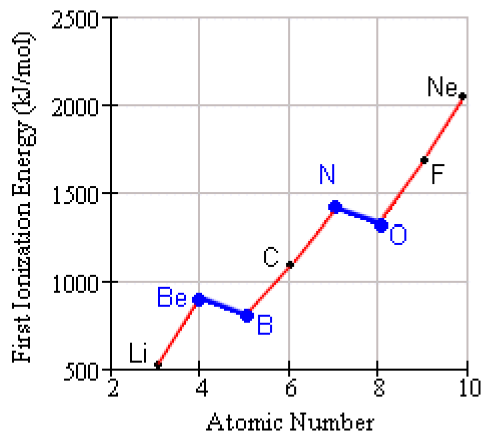

Figure 13 – Variation of first ionization energy in second row elements.

Variation of ionization energy within a group - As we go from top to bottom in a group, atomic size goes on increasing and hence the force of attraction between the nucleus and valence electrons goes on decreasing. As a consequence, the ionization energy goes on decreasing from top to bottom in a group. This is clear from the table of first ionization energies (in $kJmol^{-}$) of elements of alkali metals and halogens.

Element	**Ionisation Eenergy**	**Element**	**Ionisation Energy**
Li	520	F	1680
Na	496	Cl	1251
K	419	Br	1143
Rb	403	I	1009

The general trend of ionization energy in the periodic table can be given as follows.

> Ionisation energy of an element goes on increasing across a period from left to right and goes on decreasing from top to bottom in a group in the periodic table.

Activity 7 – Draw a graph of first ionization energy against atomic number of maximum number of elements you can. Pick up the elements which occupy similar positions on the graph and place them one below the other. Prepare their table and see whether it is similar to long form of the periodic table.

(v)Electron affinity - It is defined as the energy released by adding an electron to a gaseous atom (negative quantity) or the energy required to remove an electron from a gaseous anion (positive quantity). The negative values indicate that the anions are more stable than the neutral atoms. Electron affinity is expressed in $kJmol^{-}$ or $kcalmol^{-}$. Sometimes it is also expressed in electron volts per atom.

If only one electron is added, the electron affinity is known as **first electron affinity**. If second electron is added, the electron affinity is called the **second electron affinity** and so on. The electron affinity goes on decreasing for successive addition of electrons. Electron affinity is a measure of the attraction between the incoming electron and the nucleus. The stronger the attraction, the more energy is evolved.

The electron affinities of atoms are difficult to measure, hence precise values are not available . These values are obtained from measurements of heats of formation and lattice energies of ionic compounds of the elements. The electron affinity of an element is a measure of that element's tendency to act as an oxidizing agent (an electron acceptor) .

Following table shows the values of electron affinities of elements.

Group → / ↓ Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 73																	He *
2	Li 60	Be *											B 27	C 122	N *	O 141	F 328	Ne *
3	Na 53	Mg *											Al 42	Si 134	P 72	S 200	Cl *349*	Ar *
4	K 48	Ca 2	Sc 18	Ti 8	V 51	Cr 65	Mn *	Fe 15	Co 64	Ni 112	Cu 119	Zn *	Ga 41	Ge 119	As 79	Se 195	Br 324	Kr *
5	Rb 47	Sr 5	Y 30	Zr 41	Nb 86	Mo 72	Tc *	Ru 101	Rh 110	Pd 54	Ag 126	Cd *	In 39	Sn 107	Sb 101	Te 190	I	*295* Xe *
6	Cs 46	Ba 14	†	Hf	Ta 31	W 79	Re *	Os 104	Ir 150	Pt 205	Au 223	Hg *	Tl 36	Pb 35	Bi 91	Po	At	Rn *
7	Fr	Ra	‡	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Uuq	Uup	Uuh	Uus	Uuo
† Lanthanides				La 45	Ce 92	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm 99	Yb	Lu 33
‡ Actinides				Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Figure 14 – Periodic table of electron affinities of elements

Following figure shows the graph of electron affinity as a function of atomic number of an element.

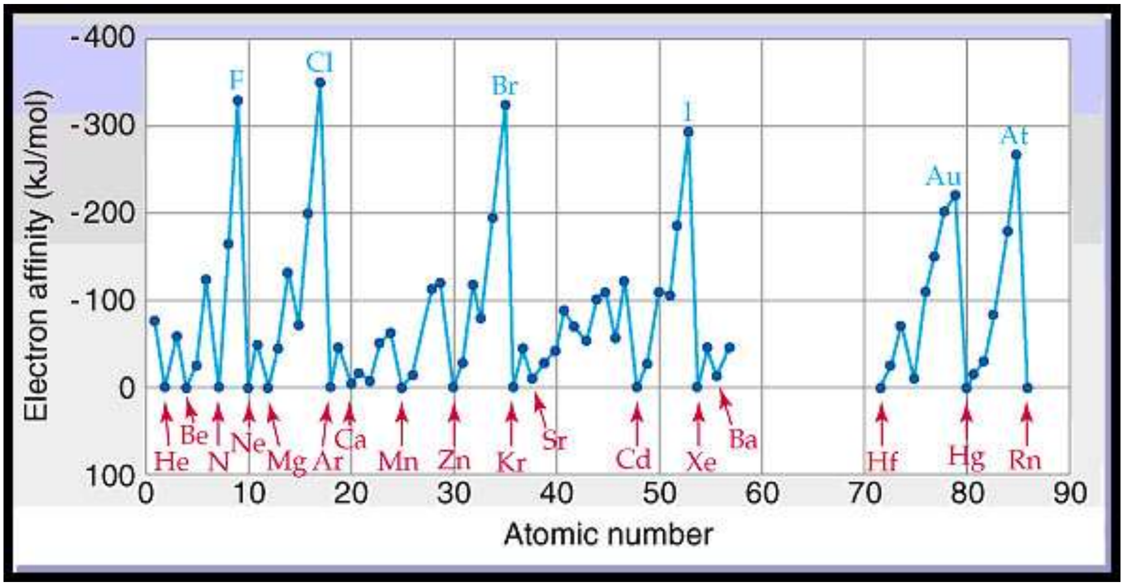

Figure 15 – Graph of electron affinity as a function of atomic number of element

Variation of electron affinity across a period - The first electron affinity values of second and third row elements are given below.

Element	Li	Be	B	C	N	O	F
Electron affinity	58	-	23	123	0	142	333
Element	Na	Mg	Al	Si	P	S	Cl
Electron affinity	53	-	44	120	74	200	348

Be, Mg and N do not form stable monovalent anion, so they have exceptionally low values of electron affinity. Be and Mg have electron affinity values below zero while N has it zero.

Following figure shows the graph of variation of electron affinity as a function of atomic number for the second row elements.

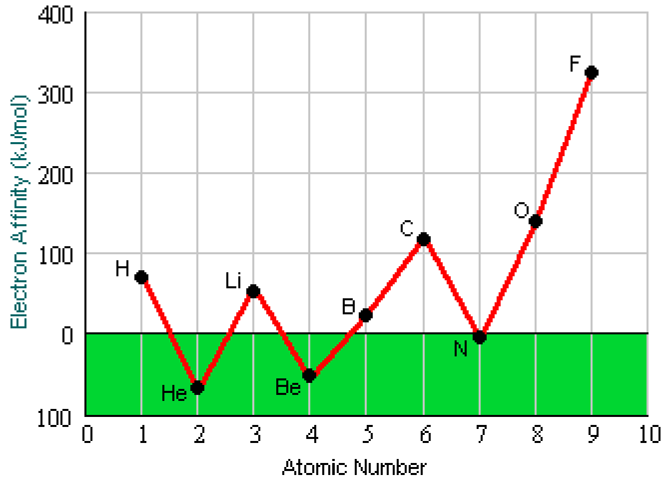

Figure 16 – Variation of electron affinity in the second row elements

Electron affinity increases from left to right across the period. This is explained as follows. Atomic size goes on decreasing across the period. So the valence electrons come nearer to the nucleus and the force of attraction between the nucleus and valence electrons goes on increasing. The valence shell comes nearer to completion of the octet and get stability. As a result the energy released i.e. electron affinity goes on increasing across the period from left to right.

Variation of electron affinity within a group - The first electron affinity values for alkali metals and halogens are given below.

Element	Li	Na	K	Rb
Electron affinity	58	53	48	45
Element	F	Cl	Br	I
Electron affinity	333	348	324	295

Following figure shows the graph of electron affinity as a function of atomic number for the oxygen and halogen family elements. Oxygen and fluorine have exceptionally low values of electron affinity because they have small atomic size and are reluctant to accept electrons.

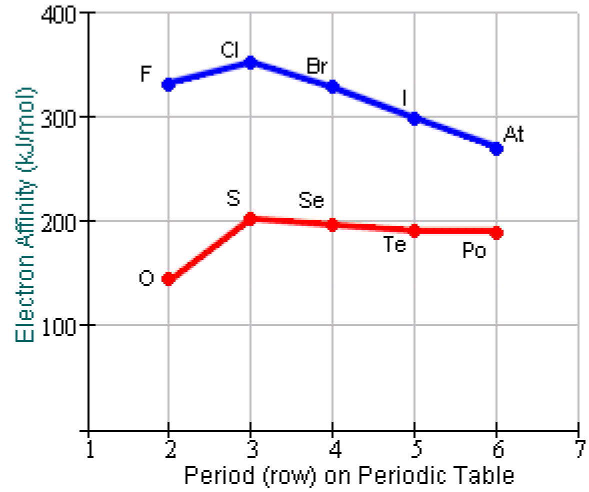

Figure 17 – Variation of electron affinity in oxygen and halogen family elements

Electron affinity decreases as we go down the group with increasing atomic number. This can be explained as follows. As the atomic number increases a filled shell is added to every next element. The valence electrons go away and the force of attraction between the nucleus and valence electrons decreases. The tendency to acquire electron decreases and hence electron affinity decreases within a group as we move from top to bottom in the group.

The general trend of electron affinity in the periodic table can be given as follows.

Electron affinity of an element goes on increasing across a period from left to right and goes on decreasing from top to bottom in a group in the periodic table

(vi) Electronegativity – The relative tendency (or ability) of an atom of an element to attract a bonded pair of electrons to itself in a molecule is called electronegativity. It is a dimensionless quantity and does not have any unit. It is a relative property.
Electron affinity is a property of gaseous isolated atoms. We normally do not deal with isolated atoms. Instead, we come across atoms which are bonded to each other. So electronegativity is a more useful property. It helps to understand the nature of chemical bond between two atoms.

Electronegativity cannot be directly measured and must be calculated from other atomic or molecular properties. Several methods of calculation have been proposed and, although there may be small differences in the numerical values of the electronegativity, all methods show the same periodic trends between elements. Pauling's scale of electronegativity is more in use. Following table shows electronegativities of elements according to Pauling scale.

→ **Atomic radius decreases** → **Ionization energy increases** →
Electronegativity increases ->

Group / Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 2.1																	He
2	Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
3	Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
4	K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr
5	Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe
6	Cs 0.7	Ba 0.9	Lu	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	Rn
7	Fr 0.7	Ra 0.9	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Uuu	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo

Periodic table of Electronegativity using the Pauling scale

Figure 18 – Periodic table of electronegativities of elements according to Pauling scale

Variation of electronegativity across a period – The electronegativities of third row elements are given below.

Element	Na	Mg	Al	Si	P	S	Cl
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0

Following figure shows a graph of electronegativity as a function of atomic number for the 3rd period elements.

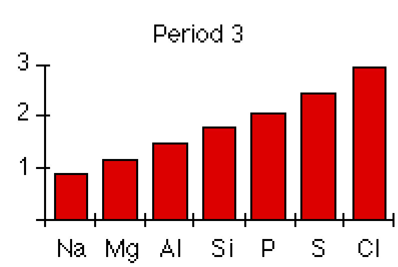

Figure 19 – Variation of electronegativity in the third row elements

It is seen that electronegativity goes on increasing from left to right in a period. The trend can be explained as follows. Atomic size goes on decreasing from left to right in a period.

The valence electrons come nearer to the nucleus. The attraction between the nucleus and valence electrons goes on increasing. So also the tendency of the nucleus and hence the atom to attract bonded pair of electrons goes on increasing. Hence electronegativity goes on increasing from left to right in a period.

Variation of electronegativity within a group – The electronegativities of alkali metals and halogens are given below.

Element	Li	Na	K	Rb
Electronegativity	1.0	0.9	0.8	0.8
Element	F	Cl	Br	I
Electronegativity	4.0	3.0	2.8	2.5

Following figure shows a graph of electronegativity as a function of atomic number for the alkali metals and halogens.

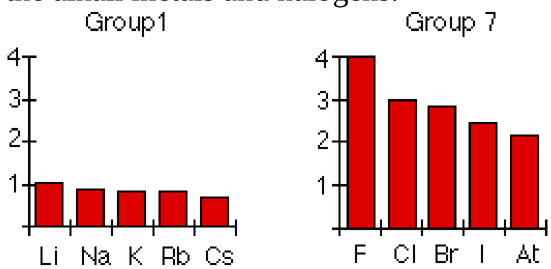

Figure 20 - Variation of of electronegativity in alkali metals and in halogens

It is seen that electronegativity goes on decreasing within a group from top to bottom. The trend can be explained as follows. Atomic size goes on increasing from top to bottom in the group. The valence electrons go away from the nucleus. The attraction between the nucleus and valence electrons goes on decreasing. So the tendency of the nucleus and hence the atom to attract bonded pair of electrons goes on decreasing. Hence electronegativity goes on decreasing from top to bottom in the group with increasing atomic number.

The general trend of electronegativity in the periodic table can be given as follows.

> Electronegativity of an element goes on increasing across a period from left to right and goes on decreasing from top to bottom in a group in the periodic table

(vii) Metallic character - Unlike ionization energy and electron affinity, metallic character can not be quantified nor can it be arbitrarily assigned any value like that of electronegativity. Metallic character is represented by physical properties like shining appearance, conductivity, ductility, malleability and so on. These days the terms ‘metallic character' and 'electropositive character' are used, more or less, with the same meaning. Electropositive character actually describes the chemical behaviour of an element. Electropositive character is the ability of an element to lose electrons. Certain properties arise due to this ability to lose electrons. They are – formation of basic oxides, displacement of hydrogen from dilute acids, formation of ionic chlorides and their reducing action. The follow up of these properties decide the metallic or electropositive character of an element.

The ability of an element to lose electrons can be related with its ionization energy which in turn depends upon the atomic size. As the atom becomes bigger in size, the valence electrons go away from the nucleus and the force of attraction between the nucleus and the valence electrons is reduced. This reduces the ionization energy and increases the ability of the element to lose the electrons. Conversely, when the atomic size becomes smaller, this increases the ionization energy and decreases the ability of the element to lose electrons .

Variation of metallic character across the period – The atomic radii (in pm), ionization energies (in $kJmol^-$) and metallic character of third row elements are given below.

Element	Na	Mg	Al	Si	P	S	Cl
Atomic radius	190	160	143	132	128	127	99
Ionisation energy	496	738	578	786	1021	1000	1251
Metallic character	Metal	Metal	Metal	Metalloid	nonmetal	nonmetal	nonmetal

The elements to the left of the periodic table have a tendency of losing electrons easily as compared to those on the right. This is clear from the values of ionization energies of the third row elements. As we move from left to right in a period, the electrons of the outer shell experience greater pull of the nucleus due to increased effective nuclear charge. So electrons of the elements on the right find it increasingly difficult to lose electrons. Hence metallic character of the element goes on decreasing from left to right across a period in the periodic table. Conversely, non-metallic character of the element goes on increasing from left to right across a period in the periodic table.

Variation of metallic character within a group – The atomic radii (in pm), ionization energies (in $kJmol^-$) and metallic character of elements of group 14 (carbon family) are given below.

Element	Atomic radius	Ionisation energy	Metallic character
C	77	1088	Non-metal
Si	118	787	Metalloid
Ge	122	782	Metalloid
Sn	140	707	Metal
Pb	150	716	Metal

As we move down the group, the number of shells increases. This causes the effective nuclear charge to decrease due to the outer shells being further away. In effect, the atomic size increases. The electrons of the outermost shell experience less nuclear attraction and so can lose electrons easily. This is confirmed from the values of ionization energies of the elements. Hence, metallic character of the element increases as we go down the group in the periodic table.
The general trend of metallic character in the periodic table can be given as follows.

> Metallic character of an element goes on decreasing across a period from left to right and goes on increasing from top to bottom in a group in the periodic table

Activity 8 – Do this activity in presence of a teacher. Take a small piece of sodium, magnesium and aluminium metal each. Put all the three pieces in a bowl containing water at room temperature. Sodium metal reacts with water but Mg and Al do not. Put the same pieces of Mg and Al in boiling water taken in a bowl. Magnesium reacts with hot water but Al does not. Prepare steam and pass it over heated Al metal. Hot Al metal reacts with steam. Write the reaction of each metal with water. Draw an inference about the pattern of reactivity of these metals.

Test your understanding

1) What is periodicity ? Why do elements show periodicity in properties?
2) Mention any three properties of elements which show periodicity.
3) Arrange each of the following sets of atoms in the ---
(i) increasing order of atomic radius-
a) Br, As, Ca, K b) Be, Ba, Ca, Sr c) Na, Si, Cl, S
(ii) decreasing order of size –
a) O^{2-}, Na^+, F^- b) Br^-, Sr^{2+}, Rb^+, Se^{2-} c) Se^{2-}, S^{2-}, O^{2-}
(iii) increasing order of first ionization energy –
a) Sr, Ca, Be b) Br, Rb, Se c) N, B, Ne
(iv) decreasing electron affinity –
a) P, Na, Ar b) B, C, N, O c) F, Cl, Br, I
(v) decreasing electronegativity –
a) Si, S, N, Cl b) C, Si, Fe, Sn c) Na, Al, Cl, P
(vi) increasing metallic character –
a) Bi, As, P, N b) Ca, K, Ge, Ga c) Rb, Br, Se, As

References / Figures / Diagrams –

1) Figure 1 – Lothar Meyer's atomic volume curve
http://www.educationalelectronicsusa.com/c/class_elem-III.htm

2) Figure 2 – Modified form of Mendeleev's periodic table
http://www.nios.ac.in/secscicour/CHAPTERo4.pdf (page 4)

3) Figure 3 – The extended long form of the periodic table
http://www.nios.ac.in/secscicour/CHAPTERo4.pdf (page 6)

4) Figure 4 - Long form of the periodic table
http://martine.people.cofc.edu/111LectWeek2_files/image006.jpg

5) Figure 5- Method of measurement of atomic radius
http://www.chemguide.co.uk/atoms/properties/atradius.html

6) Figure 6 - Periodic table of atomic radii of elements
http://www.atpm.com/8.08/images/period-table-radius-window.gif

7) Figure 7 – Graph of atomic radius as a function of atomic number of element
http://www.chem.ufl.edu/~itl/2045/lectures/lec_12.html

8) Figure 8 – Atomic radius decreases across the periods 2 and 3
http://www.chemguide.co.uk/atoms/properties/atradius.html

9) Figure 9 - Cation is smaller than atom. Anion is bigger than atom.
http://www.chemguide.co.uk/atoms/properties/atradius.html

10) Figure 10 – Ionic radii of some elements

http://www.chem.umass.edu/people/botch/Chem121F06/Chapters/Ch15/IonicRadii.jpg

11) Figure 11 – Periodic table of first ionization energies of elements

http://www.sciencebyjones.com/PDF%20files/Ionization%20Energy%20Graphing%201%20PDF.pdf

12) Figure 12 – Graph of first ionization energy as a function of atomic number of element
http://www.kentchemistry.com/images/links/bonding/ionization-energy.jpg

13) Figure 13 – Variation of first ionization energy in second row elements
http://www.chem.tamu.edu/class/majors/tutorialnotefiles/trends.htm

14) Figure 14 – Periodic table of electron affinities of elements
http://www.wikipedia.org/wiki/Electron_affinity

15) Figure 15 – Graph of electron affinity as a function of atomic number of element
http://www.iun.edu/~cpanhd/C101webnotes/modern-atomic-theory/images/electron-affinity.html

16) Figure 16 - Variation of electron affinity in second row elements
http://www.chem.tamu.edu/class/majors/tutorialnotefiles/trends.htm

17) Figure 17 - Variation of electron affinity in oxygen and halogen family elements
http://www.chem.tamu.edu/class/majors/tutorialnotefiles/trends.htm

18) Figure 18 – Periodic table of electronegativities of elements according to Pauling Scale
www.knowledgerush.com/kr/encyclopedia/Pauling_Electronegativity_Scale

19) Figure 19 – Variation of electronegativity in third row elements
www.chemguide.co.uk/atoms/bonding/electroneg.html

20) Figure 20 - Variation of electronegativity in alkali metals and in halogens
www.chemguide.co.uk/atoms/bonding/electroneg.html