Periodic Classification of Element

OBJECTIVES

After completing this lesson, you will be able to:

- ✓ **state different historical classifications of elements in brief;**
- ✓ **state main features of Mendeleev's periodic table;**
- ✓ **explain the defects of Mendeleev's periodic table;**
- ✓ **state modern periodic law;**
- ✓ **describe the features of the long form of periodic table;**
- ✓ **define various periodic properties;**
- ✓ **discuss the trends in various periodic properties in the periodic table.**

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.

In the last two lessons you have studied about the structure of atomsand their electronic configurations. You have also studied that elements with similar electronic configurations show similar chemical properties.

Electrons are filled in various shells and subshells in a fairly regular fashion. Therefore, properties of elements are repeated periodically. Such trends in their physical and chemical properties were noticed by chemists in the nineteenth century and attempts were made to classify elements on their basis long before structure of atom was known.

In this lesson we shall study about the earlier attempts for classification, the first successful classification which included all the known elements at that time namely Mendeleev's periodic table, and about the long form of modern periodic table which is an improvement over Mendeleev's work. Finally, we shall learn about some properties of elements and their variations in the periodic table.

4.1 EARLIER ATTEMPTS OF CLASSIFICATION OF ELEMENTS

The first classification of elements was as metals and non-metals. This served only limited purpose mainly because of two reasons:

1. All the elements were grouped in to these two classes only. Moreover the group containing metals was very big.
2. Some elements showed properties of both-metals and non-metals and they could not be placed in any of the two classes.

After this, scientists made attempts to recognize some pattern or regularity in variation of properties of elements and to classify them accordingly. Now we shall learn about some of them.

4.1.1 Dobereiner's triads

In 1829, Dobereiner, a German scientist made some groups of three elements each and called them triads.

All three elements of a triad were similar in their properties.

He observed that the atomic mass* of the middle element of a triad was nearly equal to the arithmetic mean of atomic masses of other two elements.

Also, same was the case with their other properties.

Let us take the example of three elements lithium, sodium and potassium. They form a Dobereiner's triad.

Mean of the atomic masses of the first (Li) and the third (K) elements: $\frac{7+39}{2}$ =23 u

The atomic mass of the middle element, sodium, Na is equal to 23 u. Two more Examples of Dobereneir's triads are given below

Element	Atomic mass
Calcium, Ca	40
Strontium, Sr	88
Barium, Ba	137

Element	Atomic mass
Chlorine, Cl	35.5
Bromine, Br ,	80
Iodine, I	127

Mean of the atomic masses of the first and third elements = $\frac{40 + 137}{2}$ = 88.5 u

Mean of the first atomic masses of the and third elements = $\frac{35.5 + 127}{2}$ = 81.5 u

Actual atomic mass of the second element = 88 u
Actual atomic mass of the second element = 80 u
Dobereneir's idea of classification of elements into triads did not receive wide acceptance as he could arrange only a few elements in this manner.

Drawback of Dobereiner's triads:

- ❖ Dobereiner could identify only three triads. He was not able to prepare triads of all the known elements
- ❖ All the known elements could not be arranged in the form of triads.
- ❖ For very low mass or for very high mass elements, the law was not holding good.
- ❖ Take the example of F, Cl, Br. Atomic mass of Cl is not an arithmetic mean of atomic masses of F and Br

Q.1 Which of the following is(are) Dobereiner's triads?
(a) P, As, Sb (b) Cu, Ag, Au (c) Fe, Co, Ni (d) S, Se, Te
A . a and b B . b and c
C . a and d D . all of the above
ANSWER
correct option C

Hint P(31), As(74.9), Sb(122) , S(32), Se(79), Te(127.6)

4.1.2 Newland's law of Octaves

In 1864 John Alexander Newland, an English chemist noticed that "when elements are arranged in the increasing order of their atomic masses* every eighth element had properties similar to the first element." Newland called it the Law of Octaves.

It was due to its similarity with musical notes where, in every octave, after seven different notes the eighth note is repetition of the first one as shown below.

1	2	3	4	5	6	7	8
Sa	Re	Ga	Ma	pa	Dha	Ni	Sa

Look carefully at the Newland's arrangement of elements shown below:

Li (6.9)	**Be (9.0)**	**B (10.8)**	**C (12.0)**	**N (14.0)**	**O (16.0)**	**F (19.0)**
Na (23.0)	**Mg (24.3)**	**Al (27.0)**	**Si (28.1)**	**P (31.0)**	**S (32.1)**	**Cl (35.5)**
K (39.1)	**Ca (40.1)**					

With the help of the arrangement given above, can you tell starting from lithium which is the eighth element? Sodium. And starting from sodium? It is potassium.
Properties of all three are similar. Similarly, aluminnium is the eighth element from boron it shows properties similar to it.
However, Newland could arrange elements in this manner only up to calcium out of a total of over sixty elements known at his time.
Because of this shortcoming his work was not received well by the scientific community. The next break through in classification of elements came in the form of Mendeleev’s work.

Drawback of newlands octaves

- There was no proper place for hydrogen.
- The observation made by Newland does not hold well for those elements which lying beyond Ca. After Calcium,every eighth element did not possess properties similar to that of the first.
- After the discovery of Nobel gases,it became difficult to fit them in Newland's periodic table

Q.2 The given arranged sequence on the basis of their increasing atomic masses represents which law of classification of elements? F,Na,Mg,Al,Si,P,S,Cl & K

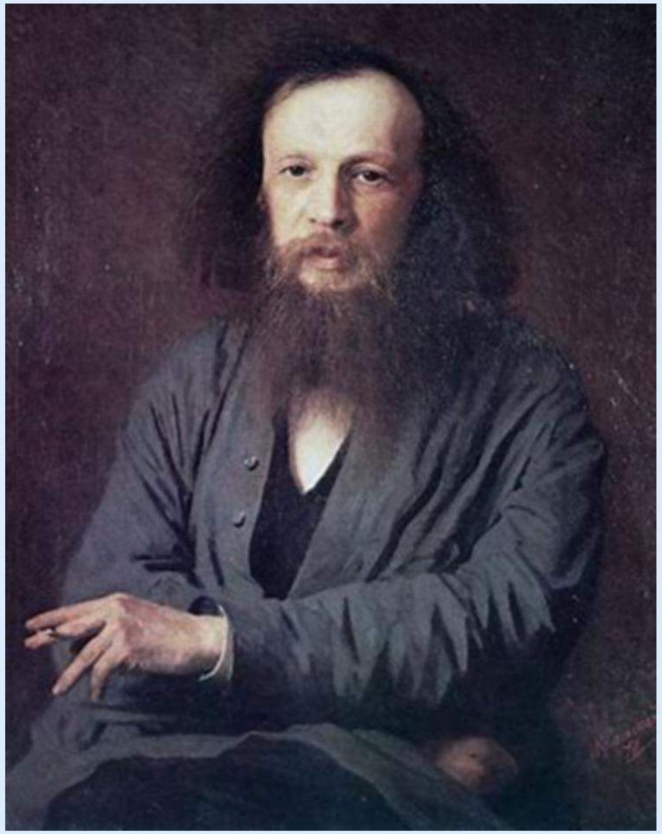
A .Modern table
B .Newton's law
C .Newland's law of octaves
D .Mendeleev's periodic table

4.1.3 MENDELEEV’S PERIODIC LAW AND PERIODIC TABLE

4.3.1a Mendeleev’s periodic law

Dmitry Mendeleev a Russian chemist while trying to classify elements discovered that on arranging in the increasing order of atomic mass*, elements with similar chemical properties occurred periodically. In1869, he stated this observation in the following form which is known as Mendeleev’s Periodic Law.**

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. Now we shall learn about the arrangement of elements on the basis of the periodic law



The chemical and physical properties of elements are a periodic function of their atomic masses*.

A tabular arrangement of the elements based on the periodic law is called periodic table. 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 of elements. Proceeding in this manner he could arrange all the known elements according to their properties and thus created the first periodic table

4.1.3b Main features of Mendeleev's periodic table

Look at the Mendeleev's periodic table shown in fig.4.2 carefully. What do you observe? Here, elements are arranged in tabular form in rows and columns. Now let us learn more about these rows and columns and the elements present in them.

1. The horizontal rows present in the periodic table are called periods. You can see that there are seven periods in the periodic table. These are numbered from 1 to 7 (Arabic numerals).
2. Properties of elements in a particular period show regular gradation (i.e. increase or decrease) from left to right.
3. The vertical columns present in it are called groups. You must have noticed that these are nine in number and are numbered from I to VIII and Zero (Roman numerals).
4. Groups I to VII are subdivided into A and B subgroups. Groups Zero and VIII don't have any subgroups.
5. All the elements in a particular group are chemically similar in nature. They show regular gradation in their physical properties and chemical reactivities. After learning about the main features we shall now learn about the main merits of Mendeleev's periodic table.

PERIODIC SYSTEM OF THE ELEMENTS IN GROUPS AND SERIES

SERIES	GROUPS OF ELEMENTS								
	0	I	II	III	IV	V	VI	VII	VIII
1	-	Hydrogen H 1.008	-	-					
2	Helium He 4.0	Lithium Li 7.03	Beryllium Be 9.1	Boron B 11.0	Carbon C 12.0	Nitrogen N 14.04	Oxygen O 16.00	Fluorine F 19.0	
3	Neon Ne 19.9	Sodium Na 23.5	Magnesium Mg 24.3	Aluminium Al 27.0	Silicon Si 28.4	Phosphorus P 31.0	Sulphur S 32.06	Chlorine Cl 35.45	
4	Argon Ar 38	Potassium K 39.1	Calcium Ca 40.1	Scandium Sc 44.1	Titanium Ti 48.1	Vanadium V 51.4	Chromium Cr 52.1	Manganese Mn 55.0	Iron Fe 55.9; Cobalt Co 59; Nickel Ni 59 (Cu)
5		Copper Cu 63.6	Zinc Zn 65.4	Gallium Ga 70.0	Germanium Ge 72.3	Arsenic As 75	Selenium Se 79	Bromine Br 79.95	
6	Krypton Kr 81.8	Rubidium Rb 85.4	Strontium Sr 87.6	Yttrium Y 89.0	Zirconium Zr 90.6	Niobium Nb 94.0	Molybdenum Mo 96.0	-	Ruthenium Ru 101.7; Rhodium Rh 103.0; Palladium Pd 106.5 (Ag)
7		Silver Ag 107.9	Cadmium Cd 112.4	Indium In 114.0	Tin Sn 119.0	Antimony Sb 120.0	Tellurium Te 127.6	Iodine I 126.9	
8	Xenon Xe 128	Caesium Cs 132.9	Barium Ba 137.4	Lanthanum La 139	Cerium Ce 140	-	-	-	
9		-	-	-	-	-	-	-	
10	-	-	-	Ytterbium Yb 173	-	Tantalum Ta 183	Tungsten W 184	-	Osmium Os 191; Iridium Ir 193; Platinum Pt 194.9 (Au)
11		Gold Au 197.2	Mercury Hg 200.0	Thallium Tl 204.1	Lead Pb 206.9	Bismuth Bi 208	-	-	
12	-	-	Radium Ra 224	-	Thorium Th 232	-	Uranium U 239		
					HIGHER SALINE OXIDES				
	R	R_2O	RO	R_2O_3	RO_2	R_2O_5	RO_3	R_2O_7	RO_4
					HIGHER GASEOUS HYDROGEN COMPOUNDS				
					RH_4	RH_3	RH_2	RH	

4.1.3c Merits of Mendeleev’s periodic classification

1. Classification of all elements :-

Mendeleev’s was the first classification which successfully included all the elements.

2. Prediction of new elements:-

Mendeleev’s periodic table had some blank spaces in it. These vacant spaces were for elements that were yet to be discovered.

For example, 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 ekaaluminium.

In this section we have learnt about the success of Mendeleev’s periodic classification and also about its merits.

Does it mean that this periodic table was perfect? No. Although it was a very successful attempt but it also had some defects in it. Now we shall discuss the defects in this classification.

4.3.1d Defects in Mendeleev's periodic classification

In spite of being a historic achievement Mendeleev's periodic table had some defects in it. The following were the main defects in it:

1. Position of hydrogen

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

2. Position of isotopes

Different isotopes of same elements 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.

3. Anomalous pairs of elements

At certain places, an element of higher atomic mass has been placed before an element of lower atomic mass. For example, Argon (39.91) is placed before potassium (39.1)

CHECK YOUR PROGRESS

1. Elements A, B and C constitute a Dobereiner's triad. What is the relationship in their atomic masses?
2. How many elements were included in the arrangement given by Newland?
3. Which property of atoms was used by Mendeleev to classify the elements?
4. How many groups were originally proposed by Mendeleev in his periodic table?
5. Where in the periodic table are chemically similar elements placed, in a group or in a period?
6. Mendeleev's periodic table had some blank spaces in it. What do they signify?
7. What name was given to the element whose properties were similar to the element eka-aluminium predicted by Mendeleev?

4.4 MODERN CLASSIFICATION

Henry Moseley, an English physicist discovered in the year 1913 that atomic number, is the most fundamental property of an element and not its atomic mass. Atomic number, (Z), of an element is the number of protons in the nucleus of its 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. Now we shall learn about the changes made in the periodic law.

4.4.1 Modern periodic law

After discovery of atomic number the periodic law was modified and the new law was based upon atomic numbers in place of atomic masses of elements.

The Modern Periodic Law states "The chemical and physical properties of elements are a periodic function of their atomic numbers"

After the change in the periodic law many changes were suggested in the periodictable. Now we shall learn about the modern periodic table which finally emerged.

4.4.2 Modern periodic table

The periodic table based on the modern periodic law is called the Modern Periodic Table.

IUPAC Periodic Table of the Elements

Key: atomic number / Symbol / name / conventional atomic weight / standard atomic weight

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H hydrogen																	2 He helium
3 Li lithium	4 Be beryllium											5 B boron	6 C carbon	7 N nitrogen	8 O oxygen	9 F fluorine	10 Ne neon
11 Na sodium	12 Mg magnesium											13 Al aluminium	14 Si silicon	15 P phosphorus	16 S sulfur	17 Cl chlorine	18 Ar argon
19 K potassium	20 Ca calcium	21 Sc scandium	22 Ti titanium	23 V vanadium	24 Cr chromium	25 Mn manganese	26 Fe iron	27 Co cobalt	28 Ni nickel	29 Cu copper	30 Zn zinc	31 Ga gallium	32 Ge germanium	33 As arsenic	34 Se selenium	35 Br bromine	36 Kr krypton
37 Rb rubidium	38 Sr strontium	39 Y yttrium	40 Zr zirconium	41 Nb niobium	42 Mo molybdenum	43 Tc technetium	44 Ru ruthenium	45 Rh rhodium	46 Pd palladium	47 Ag silver	48 Cd cadmium	49 In indium	50 Sn tin	51 Sb antimony	52 Te tellurium	53 I iodine	54 Xe xenon
55 Cs caesium	56 Ba barium	57-71 lanthanoids	72 Hf hafnium	73 Ta tantalum	74 W tungsten	75 Re rhenium	76 Os osmium	77 Ir iridium	78 Pt platinum	79 Au gold	80 Hg mercury	81 Tl thallium	82 Pb lead	83 Bi bismuth	84 Po polonium	85 At astatine	86 Rn radon
87 Fr francium	88 Ra radium	89-103 actinoids	104 Rf rutherfordium	105 Db dubnium	106 Sg seaborgium	107 Bh bohrium	108 Hs hassium	109 Mt meitnerium	110 Ds darmstadtium	111 Rg roentgenium	112 Cn copernicium	113 Nh nihonium	114 Fl flerovium	115 Mc moscovium	116 Lv livermorium	117 Ts tennessine	118 Og oganesson

57 La lanthanum	58 Ce cerium	59 Pr praseodymium	60 Nd neodymium	61 Pm promethium	62 Sm samarium	63 Eu europium	64 Gd gadolinium	65 Tb terbium	66 Dy dysprosium	67 Ho holmium	68 Er erbium	69 Tm thulium	70 Yb ytterbium	71 Lu lutetium
89 Ac actinium	90 Th thorium	91 Pa protactinium	92 U uranium	93 Np neptunium	94 Pu plutonium	95 Am americium	96 Cm curium	97 Bk berkelium	98 Cf californium	99 Es einsteinium	100 Fm fermium	101 Md mendelevium	102 No nobelium	103 Lr lawrencium

INTERNATIONAL UNION OF PURE AND APPLIED CHEMISTRY

For notes and updates to this table, see www.iupac.org. This version is dated 1 December 2018.

United Nations Educational, Scientific and Cultural Organization

International Year of the Periodic Table of Chemical Elements

IUPAC

NOMENCLATURE OF ELEMENTS WITH ATOMIC NUMBERS > 100

The naming of the new elements had been traditionally the privilege of the discoverer (or discoverers) and the suggested name was ratified by the IUPAC. In recent years this has led to some controversy. The new elements with very high atomic numbers are so unstable that only minute quantities, sometimes only a few atoms of them are obtained. Their synthesis and characterisation, therefore, require highlysophisticated costly equipment and laboratory.
Such work is carried out with competitive spirit only in some laboratories in the world.
Scientists, before collecting the reliable data on the new element, at times get tempted to claim for its discovery. For example, both American and Soviet scientists claimed credit for discovering element 104. The Americans named it Rutherfordium whereas Soviets named it Kurchatovium. To avoid such problems, the IUPAC has made recommendation that until a new element's discovery is proved, and its name is officially recognized,,,,,,, a systematic nomenclature be derived directly from the atomic number of the element using the numerical roots for 0 and numbers 1-9. The roots are put together in order of digits which make up the atomic number and "ium" is added at the end.

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

Atomic number	Name acc. to IUPAC Nomenclature	Symbol	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	Ununnilium	Uun	Darmstadtium	Ds
111	Unununium	Uuu	Roentgenium	Rg
112	Ununbium	Uub	Copernicium	Cn
113	Ununtrium	Uut	Nihonium	Nh
114	Ununquadium	Uuq	Flerovium	Fl
115	Ununpentium	Uup	Moscovium	Mc
116	Ununhexium	Uuh	Livermorium	Lv
117	Ununseptium	Uus	Tennessine	Ts
118	Ununoctium	Uuo	Oganesson	Og

If you look at the modern periodic table you will observe that it is not much different from Mendeleev's periodic table. Now let us learn the main features of this periodic table.

4.4.2a Groups

There are 18 vertical columns in 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 extreme left position contains alkali metals (Li, Na, K, Rb, Cs and Fr) and group 18 on extreme right side position contains noble gases (He, Ne, Ar, Kr, Xe and Rn).

All elements present in a group have similar electronic configurations and have same number of valence electrons. You can see in case of group 1 (alkali metals) and group 17 elements (halogens) that as one moves down a group, more and more shells are added

Group 1		Group 17	
Element	Electronic configuration	Elemet	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	2,8,8,8,1	I	2,8,18,8,7

All elements of group 1 have only one valence electron. Li has electrons in two shells, Na in three, K in four while 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 two in F to five in I.

- **Elements present in groups 1 and 2 on 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 typical or main group elements.**
- **Elements present in groups 3 to 12 in the middle of the periodic table are called transition elements. (Although groups 11 and 12 elements are, strictly speaking, not transition elements). Their two outermost shells are incomplete.**
- **However, it should be noted here that more and more electrons are added to valence shell only in case of normal elements. In transitions elements, the electrons are added to incomplete inner shells.**
- **Elements 113, 115 and 117 are not known but included at their expected positions.**
- **Group 18 on extreme right side of the periodic table contains noble gases. Their outermost shells contain 8 electrons.**
- **Inner transition elements:14 elements with atomic numbers 58 to 71 (Ce to Lu) are called lanthanides and 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 convenience sake they are shown separately below the main periodic table 14 elements with atomic numbers 90 to103 (Th to Lr) are called actinides* and 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. They are shown also separately below the main periodic table along with lanthanides.**

4.4.2b Periods

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).
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 element of this period, argon Ar (2,8,8) has eight electrons in its valence shell.
The gradual filing of the third shell can be seen 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

- **The first period is the shortest period of all and contains only 2 elements, H and He.**
- **The second and third periods are called short periods and contain 8 elements each.**
- **Fourth and fifth periods are long periods and contain 18 elements each.**
- **Sixth and seventh periods are very long periods containing 32 elements* * each.**

4.4.2c Merits of modern periodic table over Mendeleev’s periodic table

The modern periodic table is based on atomic number which is more fundamental property of an atom than atomic mass. The long form of modern periodic table is therefore free of main defects of Mendeleev’s periodic table.

1. Position of isotopes

All isotopes of the same elements have different atomic masses but same atomic number. Therefore, they occupy the same position in the modern periodic table which they should have because all of them are chemically similar.

2. Anomalous pairs of elements

When elements are arranged in the periodic table according to their atomic numbers the anomaly regarding certain pairs of elements in Mendeleev’s periodic table disappears. For example, atomic numbers of argon and potassium are 18 and 19 respectively. Therefore, argon with smaller atomic number comes before potassium although its atomic mass is greater and properties of both the elements match with other elements of their respective groups.

CHECK YOUR PROGRESS 4.2

1. According to the modern periodic law the properties of elements are periodic function of which property of theirs?

2. List any two defects of Mendeleev’s periodic table which have been corrected in the modern periodic table?

3. How many group and periods are present in the long form of periodic table?

4. What is the name of the family of elements present in group 2 of the modern periodic table?

5. The elements that are present in the right hand portion of the periodic table are metals or non-metals?

4.5. Division of Elements into S, p, d & f blocks and their Electronic Configurations:

Elements in the periodic table have been divided into four blocks. This division is based upon the electronic configuration of the element. These are

(i) S-block elements:

Elements in which the last electron enter in the S-orbital of their respective outermost shells are called S-block elements. Since S-subshell contains only two electrons, so their are only 2 groups of S-block, first group is called alkali Metal (ns^1) and second group is called alkaline earth metal (ns^2). The general electronic configuration of S-block element are ns^{1-2} where n = 2-7

General characteristics of S-block Elements:

1. They are soft metals with low melting and boiling points.

2. They have low Ionization energy and highly Electropositive.

3. They are very reactive metals.

4. They are strong reducing agents.

5. All are good conductors of heat and electricity.

(ii) p-block Elements:

Elements in which the last electron enters in anyone of the three p-orbitals of their respective outermost shells are called p-block elements. p-orbital can accommodate six electrons , So there are total six groups in the p-block.The group 13 is called Boron family (ns^2 np^1), the group 14 is called carbon family (ns^2 np^2), the group 15 is called nitrogen family (ns^2 np^3), the group 16 is called oxygen family or chalcogens (ns^2np^4), the group 17 is called halogens (ns^2np^5) and last group is called Inert gases (ns^2np^6). The general electronic configuration of p-block elements is ns^2np^{1-6} where n=2-6

General characteristics :

1. Their Ionization energies are comparatively higher than S-block elements.
2. They mostly form covalent compounds.
3. Some of them show more than one oxidation state.
4. They include both metal and non-metals. Non metals are much higher than metals. Metallic characteristics increase down the group

(iii) d-block elements:

Elements in which the last electron enters in any one of the five d-orbitals of their respective penultimate shells are called d-block elements. d-orbital can accommodate 10 electrons, so there are total 10 groups in the d-block. These are from 3-12 in the periodic table. They are further divided into four series, called as first transition series, second , third and four transition series . General Electronic configuration of d-block is $(n-1)d^{1-10}$ ns^{0-2}

General characteristics :

1. They are hard, malleable and ductile with high melting and boiling points.
2. They are good conductors of heat and electricity.
3. They show variable oxidation states.
4. They form coloured complexes.
5. They form both ionic and covalent compounds.
6. Most of transition metals form alloys.

(vi) f-block Elements:

Elements in which the last electron enters in any one of the seven f-orbitals of their respective anti penultimate shells are called f-block elements. General outermost shell electronic configuration of f-block element is $(n-2)\ f^{0-14}(n-1)d^{0-2}\ ns^2$ where n=6-7

They are also called as inner transition elements.

General characteristics:

1. They are heavy metals.
2. They are generally high melting and boiling points.
3. They show variable oxidation states.
4. Their compounds are generally coloured.
5. They have high tendency to form complexes.
6. Most of elements of the actinoid series are radioactive.

4.6 PERIODIC PROPERTIES

In the previous section 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 higher 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. Such properties are called periodic properties. In this section we shall learn about some periodic properties and their variation in the periodic table.

4.6.1 VALENCY

- 4.5.1(a) Valency in a period : You have already learnt in the previous section that the number of valence electrons increases in a period.
- In normal elements it increases from 1 to 8 in 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.
- Carefully look at the table given below. What do you observe? Valency of normal elements with respect oxygen increases from 1 to 7 as shown below for elements of third period.
- This valency is equal to the number of valence electrons or group number for groups 1 and 2, or (group number-10) for groups 13 to 17.

Group	1	2	13	14	15	16	17
Element	Na	Mg	Al	Si	P	S	Cl
No. of valence electrons	1	2	3	4	5	6	7
Valency with respect to oxygen	1	2	3	4	5	6	7
Formula of oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O

In the following table for elements of second period you will observe that valency of elements of with respect to hydrogen and chlorine increases from 1 to 4 and then decreases to 1 again.

Group	1	2	13	14	15	16	17
Element	Li	Be	B	C	N	O	F
No. of valence electrons	1	2	3	4	5	6	7
Valency with respect to hydrogen and chlorine	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

4.6.1(b) Valency in a group :

All the elements of a group have the same number of valence electrons. Therefore, they all have the same valency. Thus valency of all group 1 elements, alkali metals, is 1. Similarly valency of all group 17 elements, halogens, is 1 with respect to hydrogen and 7 with respect to oxygen.

4.6.2 Atomic Radius

Atomic radius is the distance from the centre of the nucleus to the outermost shell containing electrons. In other words, it is the distance from the center of the nucleus to the point up to which the density of the electron cloud is maximum.

Types of Atomic Radii

Atomic radii are divided into three types:

Covalent radius

Van der Waals radius

Metallic radius

Therefore, we will study these three types of radius because they are vital for a better understanding of the subject.

1) Covalent Radius

Covalent radius is one half the distance between the nuclei of two covalently bonded atoms of the same element in a molecule. Therefore, r covalent = ½ (internuclear distance between two bonded atoms). The internuclear distance between two bonded atoms is called the bond length.

Therefore, r covalent = ½(bond length)

2) Van der Waals Radius

It is one half the distance between the nuclei of two identical non-bonded isolated atoms or two adjacent identical atoms belonging to two neighboring molecules of an element in the solid-state. The magnitude of the Van der Waals radius is dependent on the packing of the atoms when the element is in the solid-state.

For example, the internuclear distance between two adjacent chlorine atoms of the two neighboring molecules in the solid-state is 360 pm. Therefore, the Van der Waals radius of the chlorine atom is 180 pm.

3) Metallic Radius

A metal lattice or crystal consists of positive kernels or metal ions arranged in a definite pattern in a sea of mobile valence electrons. Each kernel is simultaneously attracted by a number of mobile electrons and each mobile electron is attracted by a number of metal ions.

Force of attraction between the mobile electrons and the positive kernels is called the metallic bond.
It is one half the internuclear distance between the two adjacent metal ions in the metallic lattice.
In a metallic lattice, the valence electrons are mobile, therefore, they are only weakly attracted by the metal ions or kernels.

In a covalent bond, a pair of electrons is strongly attracted by the nuclei of two atoms. Thus, a metallic radius is always longer than its covalent radius.
For example, the metallic radius of sodium is 186 pm whereas its covalent radius as determined by its vapor which exists as Na_2 is 154 pm. The metallic radius of Potassium is 231 pm while its covalent radius is 203 pm

4.6.2a Variation of atomic radii in a period

Atomic radii (in picometer) of 2nd and 3rd period elements are given in the table given below. What do you observe? In a period, atomic radius generally decreases from left to right.

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

Can you explain this trend? You have learnt in the beginning of this section that in a period there is a gradual increase in the nuclear charge. Since valence electrons are added in the same shell, they are more and more strongly attracted towards nucleus. This gradually decreases atomic radii.

NOTE :-Effective nuclear

Effective nuclear charge refers to the charge felt by the outermost (valence) electrons of a multi-electron atom after the number of shielding electrons that surround the nucleus is taken into account. The trend on the periodic table is to increase across a period and increase down a group.

Effective Nuclear Charge Formula

The formula for calculating the effective nuclear charge for a single electron is:

$$Z_{eff} *=* Z - \sigma$$

Zeff is the effective nuclear charge, or Z effective

Z is the number of protons in the nucleus, the atomic number

σ (Sigma) is the average amount of electron density between the nucleus and the electron (screening constatnt)

1.Find Z: Atomic Number

Firstly, determine the value of Z which is the number of protons in the nucleus of the atom. Besides, in the example of Lithium, the value for Z is 3.

2. Find σ: Slater's Rules

Use Slater's rule to find the value of σ. Moreover, it provides numerical values for the effective nuclear charge concept. Furthermore, we can achieve this by writing out the electron configuration of the element in the following groupings and order:

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

Besides, the electron configuration corresponds to the shell level of the electrons in the atom (how far is the electron from the nucleus) and the letter corresponds to the given shape of an electron's orbit.

For example, the electron configuration of Lithium (3 electrons) will look like this: $(1s)^2$, $(2s)^1$. It means that there are two electrons in the first shell and one in second shell spherical orbital shapes

3. Find σ: Assign electron value

In this assign the value of electron according to their shell level and orbital shape. Moreover, an electron in an "s" or "p" orbit in the same shell as the electron for which you're solving contribute 0.35, also one energy level orbital lower than this contributes 0.85, and an orbital shell which is two-level lower contribute 1.

4. Find σ: Add values together

For calculating the value of S adds the number you assign to each electron in Slater's Rules.

For Example $\sigma = .85 + .85 = 1.7$ (the sum of values of two-electron)

5. Subtract σ from Z

Finally subtract the value of σ from Z to find the value of effective nuclear charge, Zeff.

For example, Us the Lithium atom, then Z =3 (atomic number) and σ = 1.7. Now put the variables in the formula to know the value of Zeff (effective nuclear charge).

$$Zeff = Z - \sigma$$

$$Zeff = 3 – 1.7 = 1.3$$

Zeff values for some elements

	H							He
Z	1							2
1s	1.00							1.69
	Li	**Be**	**B**	**C**	**N**	**O**	**F**	**Ne**
Z	3	4	5	6	7	8	9	10
1s	2.69	3.68	4.68	5.67	6.66	7.66	8.65	9.64
2s	1.28	1.91	2.58	3.22	3.85	4.49	5.13	5.76
2p			2.42	3.14	3.83	4.45	5.10	5.76
	Na	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
Z	11	12	13	14	15	16	17	18
1s	10.63	11.61	12.59	13.57	14.56	15.54	16.52	17.51
2s	6.57	7.39	8.21	9.02	9.82	10.63	11.43	12.23
2p	6.80	7.83	8.96	9.94	10.96	11.98	12.99	14.01
3s	2.51	3.31	4.12	4.90	5.64	6.37	7.07	7.76
3p			4.07	4.29	4.89	5.48	6.12	6.76

4.6.2b Variation of atomic radii in a group

What happens to atomic radii in a group? Atomic radii increase in a group from top to bottom. This can be seen from the data of atomic radii in picometers given for groups 1 and 17 elements below.

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

As we go down a group the number of shells increases and valence electrons are present in higher shell and the distance of valence electrons from nucleus increases. For example, in lithium the valence electron is present in 2nd shell while in sodium it is present in 3rd shell. Also, the number of filled shells between valence electrons and nucleus increases. Thus in group 1 Li (2,1) has one filled shell between its nucleus and valence electron while Na (2,8,1) has two filled shells between them. Both the factors decrease the force of attraction between nucleus and valence electron. Therefore, atomic size increases on moving down a group.

4.6.3 Ionic radii

Ionic radius is the radius of an ion. On converting into an ion the size of a neutral atom changes. Anion is bigger than the neutral atom. This is because addition of one or more electrons increases repulsions 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.

4.6.3a Variation of ionic radii in groups andperiods

Ionic radii show variations similar to those of atomic radii. Thus, ionic radii increase in a group. You can see such increases in groups 1 and 16 elements from the data given below.

Group 1		Group 16	
Element	ionic radius	Element	Ionic radius
Li^+	60	O^{2-}	140
Na^+	95	S^{2-}	184
K^+	133	Se^{2-}	198
Rb^+	148	Te^{2-}	221

Ionic radii decrease in a period . It can be seen from the data of ionic radii in picometer for 2nd period elements given below.

Element	Li^+	Be^{2+}	B	C	N^{3-}	O^{2-}	F^-
Ionic radii	60	31	-	-	171	140	136

In the data given above, the positions of boron and carbon have been left vacant as they do not form ions. Also, the trend in radii of cations is seen in Li^+ and $Be^{2+}$and in radii of anions is seen in N^{3-} , O^{2-} and F^-

4.6.4 Ionization energy

- ✓ Negatively charged electrons in an atom are attracted by the positively charged nucleus. For removing an electron this attractive force must be overcome by spending some energy.
- ✓ The minimum amount of energy required to remove an electron from a gaseous atom in its ground state to form a gaseous ion is called ionization energy.
- ✓ It is measured in unit of kJ mol^{-1}. It is a measure of the force of attraction between the nucleus and the outermost electron.
- ✓ Stronger the force of attraction, greater is the value of ionization energy. It corresponds to the following process:
- ✓ If only one electron is removed, the ionization energy is known as the first ionization energy. If second electron is removed the ionization energy is called the second ionization energy. Now we shall study the variation of ionization energy in the periodic table.

4.6.4a Variation of ionization energy in a group

We have already seen earlier, that the force of attraction between valence electrons and nucleus decreases in a group from top to bottom. What should happen to their ionization energy values? Ionization energy decreases in a group from top to bottom.
This can be seen from ionization energy values (in kJ mol-1) of groups 1 and 17 elements given below.

Group 1		Group 17	
Element	Ionization Energy	Element	Ionization Energy
Li	520	F	1680
Na	496	Cl	1251
K	419	Br	1143
Rb	403	I-	1009

4.6.4b Variation of ionization energy in a period

We know that the force of attraction between valence electron and nucleus increases in a period from left to right. As a consequence of this, the ionization energy increases in a period from left to right. This trend is can be seen in ionization energies (in kJ mol-1) of elements belonging to 2nd and 3rd periods.

2nd Period Elements	Li	Be	B	C	N	O	F	Ne
Ionization Energy	520	899	801	1086	1400	1314	1680	2080

3rd Period Elements	Na	Mg	Al	Si	P	S	Cl	Ar
Ionization Energy	496	738	578	786	1021	1000	1251	1521

4.6.5 Electronegativity

You have learnt in the previous section that electron affinity of an element is a measure of an isolated atom to attract electrons towards it self. We normally do not deal with isolated atoms. Mostly we come across atoms which are bonded to other atoms. There is another property which deals with the power of bonded atoms to attract electrons. This property is known as electronegativity.

Electronegativity is relative tendency of a bonded atom to attract the bond-electrons towards itself.

Electronegativity is a dimensionless quantity and does not have any units.

It just compares the tendency of various elements to attract the bond-electrons towards themselves. The most widely used scale of electronegativity was devised by Linus Pauling.

Electronegativity is a useful property. You will learn in the next chapter how it helps to understand the nature of chemical bond formed between two atoms.

Now let us learn about its variation in groups 1 and 17.

Group 1		Group 17	
Element	Electronegativity	Element	Electronegativity
Li	1.0	F	4.0
Na	0.9	Cl	3.0
K	0.8	Br	2.8
Rb	0.8	I-	2.5

What do you observe? Electronegativity decreases in a group from top to bottom.

Now let us see its variation in 2nd and 3rd period elements.

2nd Period Elements

Element	Li	Be	B	C	N	O	F
Electronegativity	1.0	1.5	2.0	2.5	3.0	3.5	4.0

3rd Period Elements

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

Now what do you observe? Electronegativity increases in a period from left to right.

4.6.6 Metallic and non-metallic character

You know what are characteristic properties of a metal?

They are its electropositive character (the tendency to lose electrons), metallic luster, ductility, malleability and electrical conductance.

Metallic character of an element largely depends upon its ionization energy.

Smaller the value of ionization energy, more electropositive and hence more metallic the element would be.

4.6.6a Variation of metallic character in a group

You know the variation of ionization energy in a group. Can you predict the variation of metallic character on its basis?

Metallic character of elements increases from top to bottom.

This can best be seen in elements of group 14.

Its first element, carbon is a typical nonmetal, next two elements Si and Ge are metalloids and the remaining elements Sn and Pb, are typical metals as shown below.

Group 14

Element	Nature
C	Non-metal
Si	Metalloid
Ge	Metalloid
Sn	Metal
Pb	Metal

4.6.6b Variation of metallic character in a period

How does metallic character change in a period? Metallic character of elements decreases in a period from left to right as shown below for 3rd period elements

Element	Na	Mg	Al	Si	P	S	Cl
Element Nature	Metal	Metal	Metal	Metalloid	Non-metal	Non-metal	Non-metal

CHECK YOUR PROGRESS

Fill in the blanks with appropriate words.

1. The force of attraction between nucleus and valence electrons ________________ in a period.
2. Atomic radii of elements ________________ in a period from left to right.
3. Radius of cation is ________________ than that of the neutral atom of the same element
4. Electronegativity ________________ in a period from left to right and ________________ in a group from top to bottom.
5. Metallic character of elements ________________ from top to bottom in a group.
6. Ionization energy of the 1st element in a period is ________________ in the entire period.

4.7 Electron gain enthalpy

Electron gain enthalpy of an element is the energy released when a neutral isolated gaseous atom accepts an extra electron to form the gaseous negative Ion i.e. anion. We can denote it by $\Delta_{eg}H$. Greater the amount of energy released in the above process, higher is the electron gain enthalpy of the element.

The electron gain enthalpy of an element is a measure of the firmness or strength with which an extra electron is bound to it. It is measured in electron volts per atom or kJ per mole. It can be an endothermic or exothermic reaction when you add an electron to the atom.

ELECTRON GAIN ENTHALPY	$kJ\ mol^{-1}$
FLUORINE	-333
CHLORINE	-348
BROMINE	-324
IODINE	-295
ASTATINE	-270.1

NOTE The negative of the enthalpy change for the process depicted in equation $X_{(g)} + e^{-} \rightarrow X^{-}_{(g)}$ is defined as the ELECTRON AFFINITY (Ae) of the atom under consideration. If energy is released when an electron is added to an atom, the electron affinity is taken as positive, contrary to thermodynamic convention. If energy has to be supplied to add an electron to an atom, then the electron affinity of the atom is assigned a negative sign. However, electron affinity is defined as absolute zero and, therefore at any other temperature (T) heat capacities of the reactants and the products have to be taken into account in DegH = –Ae – 5/2 RT.

4.7.1Some Facts about Electron Gain Enthalpy

Energy is released when an electron is added to the atom. Therefore, the electron gain enthalpy is negative.
The electron gain enthalpy for halogens is highly negative because they can acquire the nearest stable noble gas configuration by accepting an extra electron.
Noble gases have large positive electron gain enthalpy. This is because the extra electron is placed in the next higher principal quantum energy levels. Thus, a highly unstable electronic configuration is produced.

4.7.2Factors affecting Electron Gain Enthalpy

1) Atomic Size

As the size of the atom increases, the distance between the nucleus and the last shell which receives the incoming electrons increases. This decreases the force of attraction between the nucleus and the incoming electron. Hence, the electron gain enthalpy becomes less negative.

2) Nuclear charge

As the nuclear charge increases, the force of attraction between the nucleus and the incoming electron increases. Hence, the enthalpy becomes more negative.

3) Electronic configuration

Elements with exactly half filled or completely filled orbitals are very stable. You have to supply energy to add an electron. Hence, their electron gain enthalpy has large positive values. The electron gain enthalpy becomes less negative in going from top to bottom in a group. It becomes more negative in going from left to right in a period.

Anomalous Periodic Properties of Second Period Elements

It has been observed that Lithium, Beryllium, Boron, Carbon, Nitrogen, Oxygen, and Fluorine have slightly different periodic properties than the rest of the elements belonging to Groups 1, 2, 13-17 respectively. For example, Lithium and Beryllium form covalent compounds, whereas the rest of the members of Groups 1 and 2 form ionic compounds. Also, the oxide that is formed by Beryllium when it reacts with Oxygen is amphoteric in nature, unlike other Group 2 elements that form basic oxides. Yet another example is that of Carbon which can form stable multiple bonds, whereas Si=Si double bonds are not very common.

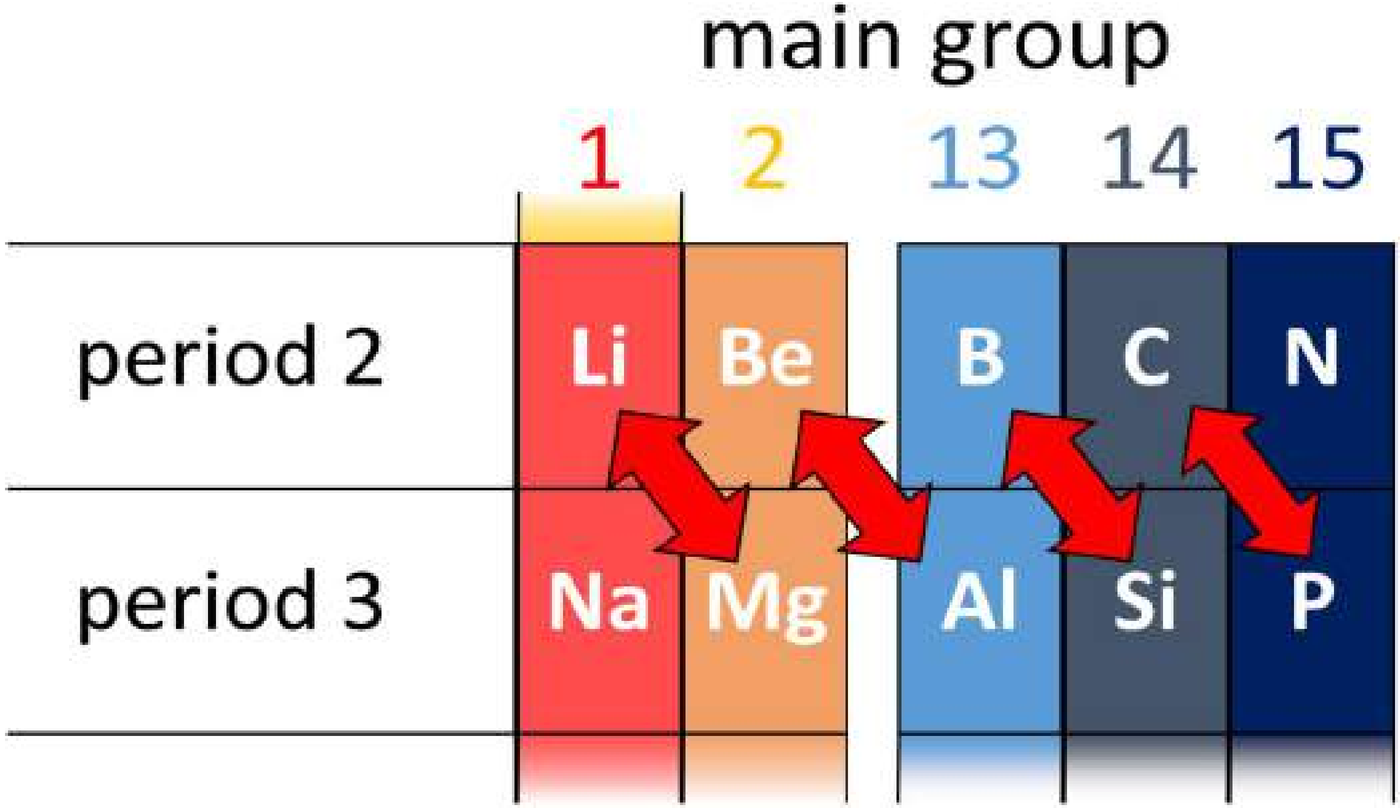
So, it has clearly been established that the second-period elements are different. In fact, they display periodic properties that are similar to the second element of the next group (i.e. Lithium is similar to Magnesium and Beryllium to Aluminium) or in other words, they have a diagonal relationship.

The reasons for differences in periodic properties and hence in chemical behavior are:

Small size of these atoms
High electronegativity
Large charge/radius ratio
These elements also have only 4 valence orbitals available (2s and 2p) for bonding as compared to the 9 available (3s, 3p, and 3d) to the other members of the respective groups, so their maximum covalency is 4. (This is why Boron can only form $[BF_4]_-$ whereas Aluminium can form $[AlF_6]^{3-}$).

Diagonal relationship

A diagonal relationship is said to exist between certain pairs of diagonally adjacent elements in the second and third periods (first 20 elements) of the periodic table. These pairs (lithium (Li) and magnesium (Mg), beryllium (Be) and aluminium (Al), boron (B) and silicon (Si), etc.) exhibit similar properties; for example, boron and silicon are both semiconductors, forming halides that are hydrolysed in water and have acidic oxides.



GLOSSARY

Actinides: A group of 14 elements with atomic numbers 90-103 (Th–Lr) which are placed along with the element actinium (Ac), atomic number 89 in the some position in group 3 in the periodic table.

Atomic number: It is the number of protons in the nucleus of the atom of an element.

Atomic radius: It is defined as one-half the distance between the nuclei of two atoms when they are linked to each other by a single covalent bond.

Dobereiner's triad: A group of three chemically similar elements in which the atomic mass and properties of the middle element are mean of the other two.

Electron affinity: It is the energy change when an electron is accepted by an atom in an isolated gaseous state. By convention, it is assigned a positive value when energy is released during the process.

Electronegativity: It is a measure of the tendency of a bonded atom to attract the bond-electrons towards itself.

Groups: The vertical columns present in periodic table.

Ionic radius: It is the radius of an ion i.e. the distance between the centre of ion and its outermost shell.

Ionization energy: It is the minimum amount of energy required to remove an electron from an isolated gaseous atom in its ground state to form a gaseous ion.

Lanthanides: A group of 14 elements with atomic numbers 58 to 71 (Ce to Lu) which are placed along with the element lanthanum (La), atomic number 57 in the some position in group 3 in the periodic table

Mendeleev's periodic law: The chemical and physical properties of elements are a periodic function of their atomic masses.

Modern periodic law : The chemical and physical properties of elements are a periodic function of their atomic numbers.

Newland's law of octaves: When elements are arranged in the increasing order of their atomic weights every eighth element has properties similar to the first.

Noble gases: The elements present in group 18 on extreme right side of the periodic table. Their outermost shells contain 8 electrons.

Normal elements: These are the elements present in groups 1 and 2 on left side and groups 13 to 17 on the right side of the periodic table whose only outermost shells are incomplete.

Periodic properties: These are the properties which repeat themselves after a certain interval of atomic number.

Periodic table: A tabular arrangement of the elements based on the periodic law.

Periods: The horizontal rows present in the periodic table. Transition elements: These are the elements present in groups 3 to 12 in the middle of the periodic table whose two outermost shells are incomplete.