CHAPTER-4

PERIODICITY OF ELEMENTS

You will learn in this chapter about:

* The search for a classification.
* Dobereiner's classification.
* Newland's classification.
* Lother Meyer's classification.
* Mendeleev's classification.
* Salient Features of Mendeleev's Table
* Modern periodic table.
* Modern periodic Law.
* Periods and groups.
* Metals, non-metals and metalloids in the periodic table.
* Some periodic properties of atoms, like atomic radii, ionization energy, electron affinity and electronegativity.

4.1 THE SEARCH FOR A CLASSIFICATION

In the early era of science, only few elements were known, but new elements were being discovered. Hence, in order to facilitate their study, need arose for a frame work in which the elements may be arranged or classified. Previously the classification was based on atomic mass (atomic weight) of elements because it was thought that the properties of an element depended upon its atomic mass. But recently a complete classification has been made, which is based on atomic number of the elements, instead of their atomic masses.

We trace the history of development of classification of elements, as follows:

4.1.1 Dobereiner's Classification:

In 1829 Johann Dobereiner, noticed that of the three elements with very similar chemical behaviour i.e. Calcium (Ca), strontium (Sr) and Barium (Ba), the atomic mass of the middle element i.e. Sr is almost the arithmetic mean of the other two. This led him to call this group of three elements, a 'triad' and proposed the **Law or Rule of Triad,** which states that:

"Central atom of each set of triad had an atomic mass almost equal to the arithmetical mean of the atomic masses of other two elements".

Table 4.1 Triads

Elements		Atomic Mass	Mean Atomic Mass
Triad	Lithium Sodium Potassium	7 23 39	At. mass of Na $= \frac{7+39}{2} = 23$
Triad	Chlorine Bromine Iodine	35.5 81 126.5	At. mass of Br $= \frac{35.5+126.5}{2} = 81$

This law or rule cannot be extended to the classification of all the elements, because it is true only in the cases of very few elements only.

4.1.2 Newland's Classification:

In 1863 John Newland, a London industrial chemist proposed **NEWLAND'S LAW OF OCTAVE**, which states that:

"If elements are arranged in the order of increasing atomic masses, the eighth element starting from a given one, has similar properties as first one i.e. its properties are a kind of repetition of the first, like the eighth note in an octave of music."

Table 4.2 Octave Law

Element	Li	Be	B	C	N	O	F
Atomic Mass	7	9	11	12	14	16	19
Element	Na	Mg	Al	Si	P	S	Cl
Atomic Mass	23	24	27	28	31	32	35.5
Element	K	Ca					
Atomic Mass	39	40					

For example, Na is eighth element from Li and has similar properties, Mg is eighth element from Be and has similar properties, etc.

This arrangement of elements for the first time brought to light the existence of **PERIODICITY** i.e. recurrence of chemical and physical properties at regular interval and provided a great idea towards the development of modern periodic table. This law failed because it held good for the first sixteen elements but did not work after seventeenth element. Moreover hydrogen was not included in this sequence.

4.1.3 Lother Meyer's Classification:

Julius Lother Meyer, a German scientist, in December 1869 published a periodic table, in which the then known 56 elements were arranged on the basis of their atomic masses in nine vertical columns or groups from I to IX. But he laid down emphasis on the physical properties of elements.

Lother Meyer calculated the atomic volumes of elements. The atomic volume of an element is the volume which would be occupied by 1 gram atomic weight (1 mole) of atoms of element if it were a solid.

$$\text{Atomic volume} = \frac{\text{Gram atomic weight}}{\text{Density}}$$

In order to emphasize the concept of periodicity, he plotted a graph between atomic volumes of elements against their increasing atomic masses.

The curve obtained is shown in Fig 4.1. It consists of sharp peaks and broad minima.

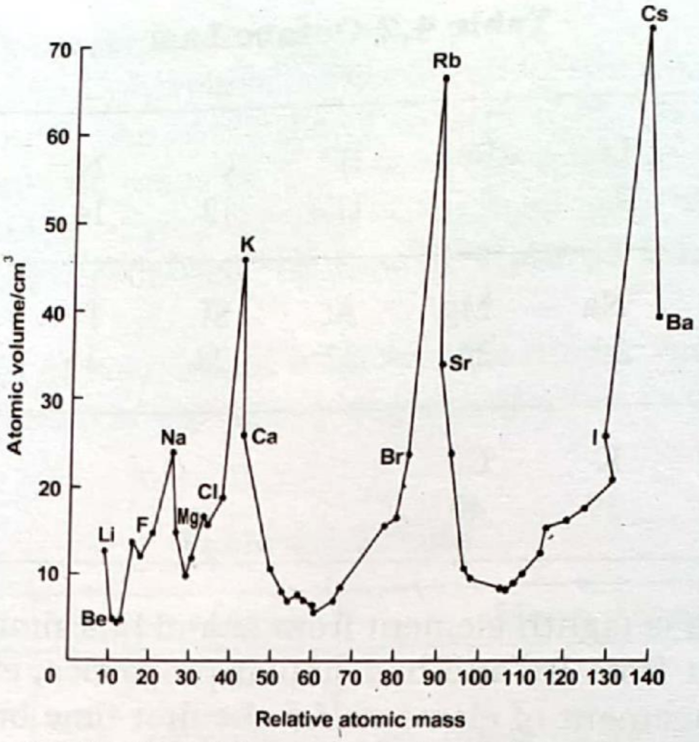


Fig 4.1 Lothar Meyer Atomic Volume Curve

He observed that the elements with similar properties occupy similar positions on the curve. For example, the highly reactive alkali metals (Li, Na, K, Rb, Cs) occupy the peaks there by showing that these elements have largest atomic volumes. The regular spacing of the highest points confirms the idea of periodicity, suggested by Newland.

4.1.4 Mendeleev's Classification:

In March 1869 Dimitri Mendeleev a Russian chemist arranged the elements in order of increasing atomic mass, placing the elements with similar chemical properties vertically beneath each other. By doing so, he observed that the properties of the elements with slight modification repeated themselves at intervals. So he put forward his Periodic Law which states that.

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

The periodic table published by Mendeleev consisted of eight vertical columns called groups (i.e. Group I to VIII) and horizontal rows called series. Now-a-days these series are called periods. The table is shown below:

Table 4.3 Mendeleev's Periodic Table of 1872

ROW	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
1	H = 1							
2	Li = 7	Be = 94	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27.3	Si = 28	P = 31	S = 32	Cl = 35.5	
4	K = 39	Ca = 40	___ = 44	Ti = 48	B = 51	Cr = 52	Mn = 55	Fe = 56 , Co = 59 Ni = 59 , Cu = 63
5	(Cu = 63)	Zn = 65	___ = 68	___ = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nh = 94	Mo = 96	___ = 100	Ru = 104 , Rh = 104 Pb = 106 , Ag = 108
7	(Ag=108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	I = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	...	...	...	
9	...	...	...	...	...	...	...	
10	...	...	?Er = 178	?La = 180	Ta = 182	W = 184	...	Os = 195 , Ir = 197 Pt = 198 , Au = 199
11	(Au = 199)	Hg = 200	T1 = 204	Pb = 207	Bi = 208	...	...	
12	...	...	...	Th = 231	...	U = 240	...	

Spaces were left for the unknown elements with atomic masses 44, 68, 72 and 100.

Salient Features of Table:

1. It has eight vertical columns called groups and twelve horizontal rows called periods.
2. Elements in each vertical column have similar properties.
3. Vacant spaces were left for the elements not discovered until then.

   He proposed their names as eka-boron, eka-aluminium and eka-silicon.
4. The group number indicate the highest valence that can be attained by elements of that group.

Advantages of Mendeleev's Periodic Table:

1. It helped in systematic study of elements. For example the study of sodium helps means to a large extent in pridicting the properties of other alkali metals as potassium, rubidium, cesium. It forcefully proved the concept of periodicity.
2. Prediction of new elements was made possible.
   The physical and chemical properties of eka-boron, eka- aluminium and eka-silicon were predicted by Mendeleev. This helped in their discovery. These have been named as scandium (Sc), Gallium (Ga) and Germanium (Ge). Their properties are remarkably the same as were predicted by Mendeleev.
3. Mendeleev's periodic table helped in correcting many doubtful atomic masses.

INTERESTING TO NOTE

Discovery of Gallium and Confirmation of Mendeleev's Prediction

Gallium was unknown when Mendeleev developed his periodic table Mendeleev left many vacant spaces in his periodic table, stating that such vacant spaces would be filled by elements not yet discovered. He did not only left vacant spaces but he also predicted the properties of the elements which when found, would occupy these spaces, like those, eka (below) aluminium and silicon. In 1875, when the French Chemist Paul Emile Lecoq, discovered the new element Gallium (from Latin Gallia for France) and gave it density as 4.7g/cm^3. Mendeleev pointed out that the density should be 5.94g/cm^3, that was proved true as predicted. The scientific community was astounded that Mendeleev (Theorist) knew the properties of new element better than the chemist who has discovered it. The new element Gallium which was called by Mendeleev Eka-Aluminium, so named because it occupied the first vacant space in his periodic table under aluminium.

Mendeleev's periodic table proved to be the key for unlocking the mysteries of atomic structure and chemical bonding. In the nineteenth century, element 101, was named Mendeleevium (Md) in his honour.

Defects In Mendeleev's Periodic Table:

1. There are three pairs of elements i.e. elements of higher atomic masses placed before elements of lower atomic masses i.e.
   (a) Argon (40) placed before potassium (39)
   (b) Cobalt (59.9) placed before nickle (58.6)
   (c) Tellurium (127.6) placed before iodine (126.9)
2. No place for isotopes of elements.
3. Dissimilar elements placed in same group i.e. Alkali metals (Li, Na, K, Rb, Cs, Fr) were placed with coinage metals (Ag, Cu, Au)
4. Similar elements placed in different groups for example Barium (Ba) and lead (Pb) resemble in many properties but they are placed in separate groups.
5. It failed to give the idea of atomic structure.

4.1.5 Modern Periodic Table:

The arrangement of elements on the basis of their atomic masses left many anomalies in the position of different elements in the periodic table. Moreover the existence of isotopes showed that the atomic mass of an element is not the fundamental property of an element.

The modern periodic table is the result of discovery of atomic number by Moseley in 1914.

Based on the concept of atomic number Bohr, Werner and Bury proposed the **Modern Periodic Law** which states that,

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

In modern periodic table, also known as Bohr's Long Form Of Periodic Table, elements are arranged in order of their increasing atomic number. The elements having similar properties are repeated at regular intervals. As atomic number is related to the number of protons in an atom, so the real basis of periodicity of properties is due to recurrence of similar valency shell configuration of the next element in the same group. The modern periodic table is shown in table 4.4

Table 4.4 Modern Periodic Table of the Elements

1	Atomic number
H	Symbol
1.008	Atomic weight

Main groups | Transition metals | Main groups

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H 1.00794																	2 He 4.00260
3 Li 6.941	4 Be 9.01218											5 B 10.81	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.998403	10 Ne 20.1797
11 Na 22.98977	12 Mg 24.305											13 Al 26.98154	14 Si 28.0855	15 P 30.97376	16 S 32.066	17 Cl 35.453	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.69	29 Cu 63.546	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.9055	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.41	49 In 114.82	50 Sn 118.710	51 Sb 121.757	52 Te 127.60	53 I 126.9045	54 Xe 131.29
55 Cs 132.9054	56 Ba 137.33	57 *La 138.9055	72 Hf 178.49	73 Ta 180.9479	74 W 183.85	75 Re 186.207	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.9665	80 Hg 200.59	81 Ti 204.383	82 Pb 207.2	83 Bi 208.9804	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.0254	89 †Ac 227.0278	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 Uun (269)	111 Uuu (272)	112 Uub (277)	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

*Lanthanide series	58 Ce 140.12	59 Pr 140.9077	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.9254	66 Dy 162.50	67 Ho 164.9304	68 Er 167.26	69 Tm 168.9342	70 Yb 173.04	71 Lu 174.967
†Actinide series	90 Th 232.0381	91 Pa 231.0359	92 U 238.0289	93 Np 237.048	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

The modern periodic table contains seven horizontal rows called periods and sixteen vertical columns called groups.

4.1.6 Periods:

The elements within a period have dissimilar properties from left to right across any period, the physical and chemical properties of elements change from metallic to non metallic along a period.

All periods except the first, start with an alkali metal with one electron in their valence shell and end up with zero group element with valence shell having 8 electrons, except He which has only 2 electrons.

The First Period:

It contains only two elements i.e. H and He. This period signifies the completion of K-shell or first orbit. It is the shortest period with two elements.

The Second and Third Period (Short Periods):

Each of these periods contains 8 elements. They signify the filling up of L-shell and M-shell respectively.

The second period starts with Li and ends up with Ne; whereas the third period starts with Na and ends at Ar.

The Fourth and Fifth Period (Long Periods):

Each of these periods contains 18 elements. In these periods the electrons fill M and N shells. Fourth period starts from K and ends at Kr. Fifth period starts from Rb and ends at Xe.

The Sixth Period (Longest Period):

It contains 32 elements. It starts from Cs and ends with Rn.
Besides, fourteen elements called **Lanthanides,** are placed at the bottom of periodic table.

The Seventh Period (Incomplete Period):

It starts with Francium (Fr). This period is incomplete as to date about 109 elements have been discovered.

This period also includes a group of fourteen elements starting from **Actinium** (Ac). These elements are called **Actinides**. They are also placed at the bottom of the table.

4.1.7 Group:

The vertical columns are called groups. Basically there are eight groups (I to VIII) but each group is further sub-divided into "A" and "B", sub-groups.

The elements of sub-group 'A' are called **'Main'** or **Representative Elements,** as the properties of these elements are represented by valency electrons.

The elements of sub-group 'B' are called **Transition Elements,** because the properties of these elements show a gradual change or transition between the two sets of representative elements, on either side of them.

Elements of a group have similar valency shell configuration hence have similar properties.

The group number indicates the total number of electrons in valency shell of that element.

Group I A (The Alkali Metals) or (Lithium Family)

This group includes Lithium (Li), Sodium (Na), Potassium (K), Rubidium (Rb), Caesium (Cs) and Francium (Fr). Their valence shell contains one electron only, and on reaction they lose this electron and form univalent positive ions (M^{1+}). They are highly reactive metals with low melting points. Fr is radioactive.

Their atomic radii, atomic volumes, ionic radii increase from Li to Cs due to the addition of extra shell to each element and due to same reason, the melting and boiling points decrease downward. They are called **Alkali Metals** because they form water soluble base such as $NaOH$ and KOH.

Group IIA (The Alkaline Earth Metals); (Beryllium Family)

It includes Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium(Sr), Barium (Ba) and Radium (Ra).

Their valence shell contains two electrons. On reaction they lose these two electrons and form divalent positive ions (M^{2+}). Ra is radioactive.

These elements are a bit harder, having higher melting and boiling points than the alkali metals, but they have smaller atomic, ionic radii and atomic volumes,

Down the group they do not show a regular trend in melting, boiling points and densities.

Group IIIA (The Boron Family):

It includes Boron (B); Aluminium (Al); Gallium (Ga); Indium (In) and Thallium (Tl). Their valency shell contains three electrons. They exhibit a valency of 3^+ and form M^{3+} ions. Except boron they are highly electropositive elements i.e. having metallic character which increases down the group; due to increase in atomic volume. Boron is **Metalloid**. A metalloid is an element which has some properties of metals and some properties of non-metals.

Group IVA (Carbon Family):

It includes Carbon (C); Silicon (Si); Germanium (Ge), Tin (Sn) and Lead (Pb).

Their valence shell contains four electrons. C, Si and Ge form covalent compounds whereas Sn and Pb exhibit a variable valence of 2 and 4.

Of these elements C is non-metal; Si and Ge are metalloids, Sn and Pb are metals.

Down the group atomic radii and volumes increase due to addition of a new shell and for the same reason metallic character increases down the group. C and Sn exist in different allotropic forms.

Group V (Nitrogen Family):

It includes Nitrogen (N), Phosphorus (P) Arsenic (As), Antimony (Sb) and Bismuth (Bi).

Of these elements N and P are non-metals, As and Sb are metalloids and Bi is a metal. Their valence shells contain five electrons. There is a large variation of properties as we go down the group.

Nitrogen exists as diatomic molecules (N_2) and forms a number of oxides as NO, N_2O, NO_2, N_2O_4 and N_2O_5. Due to small atomic size and large ionization potential, nitrogen has a tendency to accept three electrons to form nitride ion (N^{3-}). Phosphorus exists as P_4 molecule.

Except Nitrogen all exist in more than one allotropic forms.

Group VI (Oxygen Family):

It includes oxygen (O); Sulphur (S); Selenium (Se), Tellurium (Te) and Polonium (Po).

Of these oxygen and sulphur are non-metals, selenium, tellurium are metalloids and polonium is metal.

All of the elements exhibit the property of allotropy. For example allotropic forms of oxygen are oxygen (O_2) and ozone (O_3).

Oxygen and sulphur form divalent negative ions O^{2-} and S^{2-}. Their valence shell contains six electrons.

Group VIIA (The Halogens):

It includes Fluorine (F); Chlorine (Cl); Bromine (Br); Iodine (I) and Astatine (At).

Except Astatine (which is a metalloid) all others are non-metals and exist as diatomic molecules. At room temperature F_2 and Cl_2 are gases; bromine is a liquid and iodine is a solid. Their valence shell contains seven electrons. They have high ionization energies and large negative electron affinities hence they easily accept an electron to form halide ions $(x)^{1-}$ i.e. (F^{1-}, Cl^{1-}, Br^{1-}, I^{1-})

Group VIIIA (Inert or Noble Gases) :

It includes Helium (He); Neon (Ne); Argon (Ar), Krypton (Kr); Xenon (Xe) and Radon (Rn).

Their valence shell contains eight electrons, except helium which has two electrons. With the exception of krypton and xenon (which have large atomic volumes so slightly reactive under drastic conditions) the rest of these elements are totally inert chemically. The reason is that these have completely filled outer shells, a condition that represents greater stability.

INTERESTING TO NOTE

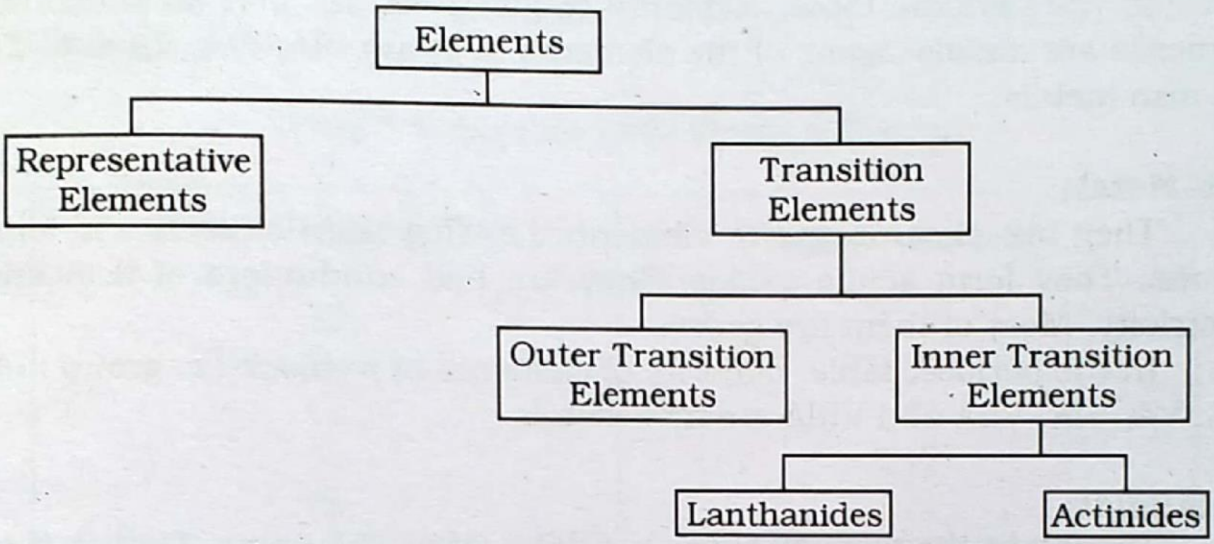
Discovery of Noble Gases

None of the noble gases was known when Mendeleev proposed his periodic table. In 1892. The English scientist Ramsay became interested in the discovery that nitrogen obtained from the air had a slightly higher density than that prepared by chemical reactions. After careful investigation, he concluded that higher density must be due to the presence of unknown gas. When he separated this gas from the air, he found that it was completely unreactive. He called it argon, the "idle or lazy" gas in Greek.In the same year Ramsay isolated helium (He) the lightest of all noble gases, from uranium ores. During 1898 Ramsay and Rayleigh isolated three additional noble gases from air, neon (Ne) krypton (Kr) and xenon (Xe).

Group IB To VIIIB (Transition Elements):

These are metals. In these elements, besides the valence shell penultimate shell is also incomplete. In chemical reactions they show more than one valencies. These elements in compounds having characteristic colours.

The total classification can be summarized in the following scheme:



4.1.8 Non-metals, Metalloids and Transition Metals in Periodic Table.

Majority of the known elements are metals; only 17 elements are non-metals and eight elements are metalloids.

Table 4.5

Metals (IA–IIA); METALS (IIIB–IIB); Non Metals and Metalloids (IIIA–VIIIA)

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Te	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 T1	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Unq	105 Unp	106 Unh	107 Uns	108 Uno	109 Une									

METALS

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Metals:

They are electropositive elements i.e. they lose electrons to form cations. They form basic oxides.

All of them have lustre and are, malleable (i.e. can be spread out into sheets) and ductile (i.e. can be drawn into wire), are good conductors of heat and electricity.

In the periodic table, elements of group IA, IIA and all transition elements are metals. Some of the elements of group IIIA, IVA, VA and VIA are also metals.

Non-Metals:

They are electronegative elements i.e. they gain electrons to form anions. They form acidic oxides. They are bad conductors of heat and electricity. Most of them are gases.

In the periodic table, majority of elements of p- block i.e. group IIIA, IVA, VA, VIA, VIIA and VIIIA are non-metals.

Metalloids:

These are the elements which exhibit dual character. That is they show the properties of both metals as well as non-metals. For example, their oxides are amphoteric i.e. have basic as well as acidic nature.

Examples are:
Boron (B) of group IIIA
Silicon (Si) and Germanium (Ge) of group IVA.
Arsenic (As) and Antimony (Sb) of group VA.
Tellurium (Te) and Polonium (Po) of group VIA
Astatine (At) of group VIIA.

4.2 SOME PERIODIC PROPERTIES OF ATOMS

4.2.1 Atomic Radii:

Modern research shows that an atom does not have strictly defined boundaries. So it is impossible to determine the exact radius of an atom, but for all practical purposes the atomic radius may be defined as half the distance between two adjacent nuclei of two similar atoms in touch with each other.

It is measured in Angstrom unit (A^0 or A.U)
$1A^0 = 10^{-8}$cm.

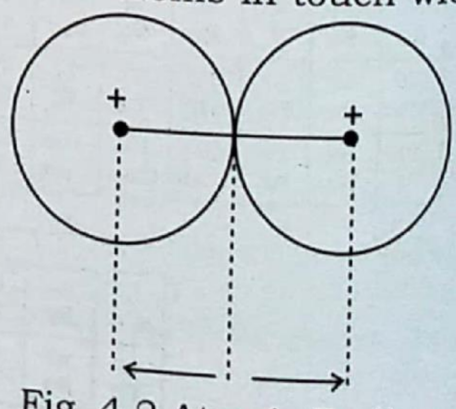


Fig. 4.2 Atomic Radius

The atomic radius depends upon the number of shells and nuclear charge in an atom.

In the periodic table the atomic radii increase down the group due to addition of new shell in each atom. But in a period the atomic radii decrease from left to right due to increase in number of protons i.e. increase in nuclear charge, which results in stronger pull on orbiting electrons by the nucleus. The variation of atomic radii in a group and a period is shown is following tables.

Table 4.6 Atomic radii Down a Group.

Elements of Group IIA	Atomic Radii in A^0
Be	1.12
Mg	1.36
Ca	1.97
Sr	2.15
Ba	2.22

Table 4.7 Atomic Radii in a Period

Elements	Na	Mg	Al	Si	P	S	Cl
Atomic Radii in A^0	1.51	1.36	1.25	0.77	0.70	0.66	0.64

4.2.2 Ionization Energy:

Ionization energy is one of the few fundamental properties which can be measured directly. It is defined as the minimum energy required to remove an electron from a gaseous atom in its ground state. It is measured in K.Jole/mole or electron volt (ev)per atom.

Ionization energy depends upon atomic size and nuclear charge. The higher the ionization energy the more difficult is to remove an electron. The ionization energy of hydrogen is 1312 K.J/mol. i.e.

$$H(g) + \text{Energy} \longrightarrow H^+ + e^- \quad I.E = +1312\ KJ/mol$$

Down a group in the periodic table, the ionization energy decreases because the addition of a new shell decreases the hold of nucleus on valence electron.

Ionization energy increases from left to right in a period because the addition of protons in the nucleus, increases the nuclear charge there by increasing the force of attraction on electrons.

The amount of energy required to remove first electron is called **first ionization energy.** For subsequent electrons it is called second, third, fourth ionization energy, etc.

The periodicity of ionization energy can be observed in the plot between first ionization potential and atomic number as shown in Figure 4.3.

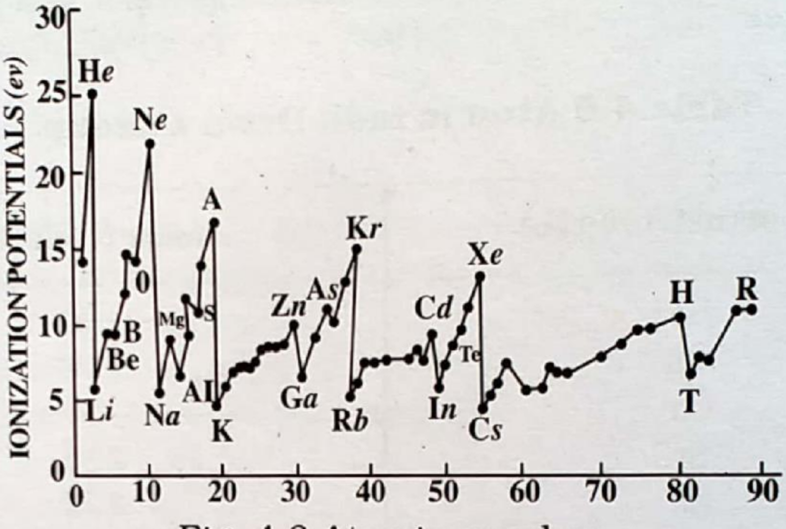


Fig. 4.3 Atomic number

4.2.3 Electron Affinity:

Electron affinity is defined as the energy change that occurs when an electron is gained by an atom in the gaseous state. It is measured in KJ/mol or in e.v per atom. Electron affinity for the addition of first electron is negative i.e. Energy is released but for further addition of electrons it is positive, because energy has to be added to overcome repulsion between negative ion and electron, as shown below:

$O(g) + e^- \longrightarrow O^-(g)$ E.A = -142 KJ/mol. (Exothermic)

$O^- + e^- \longrightarrow O^{2-}(g)$ E.A = +780 KJ/mol. (Endothermic)

Electron affinity depends upon the atomic size and nuclear charge.

Down a group in the periodic table, electron affinity decreases because the addition of a new shell to each atom decreases its force of attraction.

Table 4.8

Element	Electron Affinity in KJ/mol
F	-333
Cl	-348
Br	-324
I	-295

Flourine (F) has abnormally low electron affinity because due to its very small atomic size it does not accept electron easily.

In a period, the electron affinity increases from left to right because successive atoms have higher nuclear charge and attract the incoming electron more towards itself. See table 4.9

Table 4.9

Element	Li	Be	B	C	N	O	F	Ne
E.A in KJ/mol	-58	0	-23	-123	0.2	-142	-333	O

4.2.4 Electronegativity (E.N):

Electronegativity is defined as the relative tendency of an atom in a molecule to attract shared pair of electrons to itself. It is denoted by a number and has no unit.

Linus Pauling calculated the electronegativities of different elements taking fluorine as standard with its electronegativity = 4

Down a group electronegativity decreases as due to addition of new shell, the power of a nucleus to attract electron decreases. In a period from left to right it increases due to increase in nuclear charge.

Generally speaking, elements with high ionization energy and large electron affinity have high electronegativity. Variation of electronegativity with atomic number and the periodicity in it is shown in the following figure.

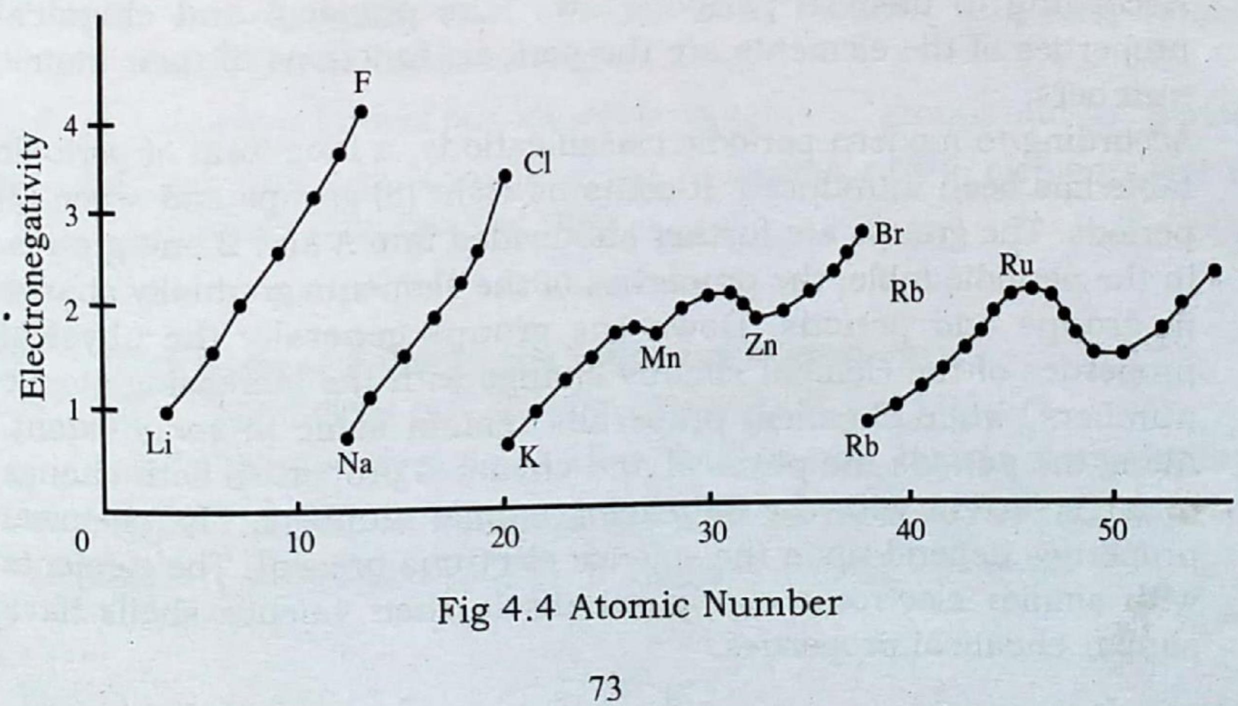


Fig 4.4 Atomic Number

The difference in electronegativities of two combining atom decides the nature of bond that is formed between them and affects the properties of molecules.

SUMMARY

1. With the discovery of new elements, scientists tried to arrange them in a scientific and organised manner, on the basis of their atomic masses. Dobereiner in 1829 was the first to classify the similar elements in the groups of three elements, known as **"triads"**. Newland in 1863 then classified the elements in the increasing atomic masses and put forward the **law of octaves.** In 1869 German chemist Lothar Meyer improved the Newland's classification and plotted a graph between the atomic volumes and atomic masses of the elements. He included only 56 elements and discovered that the elements with similar properties occupied the similar positions on the curve.
2. In 1869 Russian scientist Mendeleev arranged the elements in the increasing order of atomic masses and put forward a periodic law, known as Mendeleev's periodic law. In 1873, he published his periodic table which consisted of eight vertical columns, known as groups and twelve (12) horizontal rows known as periods. Many places were left vacant for undiscovered elements in his periodic table and predicted their properties. He also discovered correct atomic masses of many elements.
3. In (1913-1914), after the discovery of atomic numbers, the elements were then arranged in their increasing atomic numbers and modern periodic law was introduced.
   According to modern periodic law. "The physical and chemical properties of the elements are the periodic functions of their atomic numbers.
4. According to modern periodic classifications, a long-form of periodic table has been introduced. It contains eight (8) groups and seven (7) periods. The groups are further subdivided into A and B sub-groups.
5. In the periodic table, the properties of the elements gradually change in groups and periods. Down the groups generally the physical properties of the element slightly change with the increasing atomic numbers, while chemical properties remain same to some extent. Along the periods the physical and chemical properties both change to larger extent with the increasing atomic numbers. The chemical properties depend upon the valence electrons present. The elements with similar electronic configurations in their valence shells have similar chemical properties.

6. In the periodic table, elements on the left are metals, while on the right side are non-metals and in the middle, the elements are generally metalloids showing both the characteristics of metals and non-metals.
7. The minimum amount of energy required to remove the most losely bound electron from the valence electrons in a gaseous atom to form positive ion. i.e. cation is called first ionization energy or ionization potential and the minimum amount of energy released from a gaseous atom when an electron is added in the valence shell to form negative ion i.e. anion is called Electron affinity. Down the groups the values of ionization potential and electron affinity both decrease with the increasing atomic numbers due to increase in atomic sizes. While along the periods the values of ionization potential and electron affinity both increase with the increasing atomic numbers.
8. The relative power of an atom to attract the shared pair of electrons towards itself is called electronegativity. The most electronegative atom is that of fluorine (F) with electronegativity = 4.0.
The repetition or reccurrence of properties after regular intervals is called periodicity.

EXERCISE

1. Fill in the blanks:

(i) The rule of triad was introduced by ....

(ii) The repetition of properties after regular intervals is called ......

(iii) The longest period is .... period and contains total ..... elements.

(iv) The elements that contain both metallic and non-metallic characteristics are called ....

(v) The long form of periodic table contains ..... groups and ...... periods.

(vi) According to Mendeleev the properties of the elements are the periodic functions of their ....

2. Tick True or False in the following statements:

(i) Mendeleev put forward his periodic law in 1856.

(ii) The first period contains two elements, hydrogen and hellium.

(iii) The longest period in the periodic table is 7th period.

(iv) Lanthanides and Actinides are d-block elements.

(v) Down the group the electronegativity increases with increasing atomic number.

(vi) The law of octaves was introduced by John Newland.

(vii) Li^7, Na^{23} and K^{39} form a triad.

3. Pick up the correct answer: (Multiple Choice Questions).

(i) Mendeleev's periodic table contained .... periods
(7, 8, 12, 10)

(ii) The incomplete period in the periodic table is ....
(7, 6, 3, 1)

(iii) The most reactive metal is ....
(Na, Cu, Fe, Ca)

(iv) The only liquid metal is ......
(Molybdenum, Gold, Mercury, Bromine)

(v) Lothar Meyer's curve included about ...... elements.
(Thirty, forty, fifty six, sixty two)

(vi) To which family does Ga belong ?
(Boron, Carbon, Nitrogen, Fluorine)

(vii) The elements of VII A group are known as,
(Halogens, Lanthanides, Actinides)

4. Write answer to the following questions:

(i) Define the followings:
- (a) Doberneir's rule of triad
- (b) Periodicity
- (c) Modern periodic law
- (d) Electronegativity

(ii) If an element contains two shells only and its outershell contains five electrons then to which group the element belongs in the periodic table. Name the element. Predict its period.

(iii) State Mendeleev's periodic law. Describe Mendeleev's periodic table. Write down the advantages and disadvantages of Mendeleev's periodic table.

(iv) Explain Newland's law of Octave. How this law provided the larger scope for the classification of the elements?

(v) Which pair of elements is chemically similar?
- (a) K, Cr
- (b) Cu, Ca
- (c) F, Cl
- (d) N,O

(vi) What do you understand by long form of periodic table? Explain some of its applications?

(vii) Discuss some of the physical properties of the elements which exhibit periodicity.

(viii) The elements having eight valence electrons are known as,
(a) Noble or inert gases (b) Halogens (c) Nitrogen family
(d) Transition elements.

(ix) How does the modern periodic law differ from Mendeleev's periodic law? Explain clearly groups and periods in the modern periodic table.

(x) What do you understand by the periodic classification of elements? What are the merits and demerits of the classification of the elements in the periodic tabular form?

(xi) What do you understand by representative and transition elements?

(xii) Discuss the following physical properties of the elements.
(a) Ionization energy (b) Electron affinity
(c) Electronegativity (d) Atomic Radii

(xiii) What are lanthanides and actinides? Are they d or f-block elements?