15

PERIODIC TABLE AND VARIATION IN ATOMIC PROPERTIES

15.1 INTRODUCTION

We all have seen different heaps of onions & potatoes in a vegetable shop. Imagine, they are lying mixed and you want to buy 1kg of onions. What will happen? You will have to wait for long to sort that and then weigh them

When you possess a variety of material substances, you have to keep them classified for an easy access and quick use. You cannot afford to mix clothes with eatables, cosmetics or books. Classification assures you that your eatbles are in the kitchen, books on the study table or rack and your cosmetics are on the dressing table. Shopkeepers, business houses, storekeepers, administrators, managers, information technology experts and scientists etc. have to keep their materials duly classified.

Chemists faced a similar problem when they were to handle a large number of elements. The study of their physical and chemical properties and keeping a systematic record of this had been a great challenge to chemists; Classification of elements makes their study easy.

Grouping of elements with similarity of properties and their systematic study was finally made easy with the discovery of periodic law and its applications which are discussed in this lesson.

15.2 OBJECTIVES

After reading this lesson, you will be able to

- recognise the need for classification of elements.

- recall the earlier attempts on classification of elements
- define modern periodic law
- Co-relate the sequence of arrangements of elements in the periodic table with the electronic configuration of the elements
- recall the designations of the groups (1-18) in the periodic table
- locate the classification of elements into s-, p-, d-, and f- blocks of the periodic table.
- explain the basis of periodic variation of

  (a) atomic size (b) ionic size

  (c) ionization energy (d) electron affinity within a group or a period.

15.3 EARLY ATTEMPTS

Attempts were made to classify elements ever since the discovery of metals or may be even earlier. Some grouped metals separately from non-metals and some applied formulation of arithmetics to the known atomic weights of certain elements. J.W. Dobereiner in 1817 discovered that when closely related elements are grouped in a set of three, the atomic weight of the middle element had an atomic weight almost the arithmetical mean of the other two elements in that group. e.g.,

Element	Lithium	Sodium	Potassium
Atomic weight	6.94	22.99	39.10
mean atomic weight	——	23.02	——

He called such a group of three elements a Triad. He could group only a few elements due to lack of knowledge of correct atomic weights of the elements at that time.

In 1863, J.A.R. Newlands, developed a system of classification of elements and entitled it **Law of Octaves.** He arranged the elements is such a way that every eighth element had similar properties, like the notes of music. The law could not apply to a large number of known elements. However, the law indicated very clearly the recurrence of similar properties among the arranged elements. Thus the periodicity was visualised for the first time in a meaningful way.

> **Periodicity**: Re-occurrence of properties after regular intervals.

More significant results were obtained when Lother Meyer's work reflecting the periodicity was found to based on physical properties of the elements. He clearly showed that certain properties showed a periodic function.

15.4 MENDELEEV'S PERIODIC TABLE

In 1869, Mendellev's first research work on classification of the elements got published under the title "The Relation of the Properties to the Atomic Weights of the Elements", The table showing the arrangement of elements stressed the fact that periodicity is a certainity and that the vacant spaces in the table will be filled when the correct atomic weights of the elements are known or when the element of that particular atomic weight is discovered.

S. Nos.	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
1	H = 1							
2	Li = 7	Be = 9.4	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	Sg = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59 Ni = 59, Cu = 63
5	Cu = 63	Zn = 65	Ga = 68	Ge = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	Y = 88	Zr = 90	Nb = 94	Mo = 96	- = 100	Ru = 194, Rh = 104 Pd = 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, In = 197 Pt = 198, Au = 199
11	Au = 199	Hg = 200	Tl = 204	Pb = 207	Bi = 208	...	...	
12	...	...	...	Th = 231	...	U = 240	...	

***Table 15.1 Mendeleev's Table of 1871**

The extent of knowledge regarding the chemical properties of the elements and his insight into the system of periodicty possessed by the elements under certain arrangement have no parallel in the history of chemistry. This work layed strong foundation for the existence of fundamental principles of the periodic law. One of his most important conclusions include that the elements if arranged according to their atomic weights, exhibit an evident systematic reoccurence of properties (periodicity of properties) and even the properties of some elements were listed much before their discovery. Mendeleev's periodic table (table 15.1) was quite useful till the discovery of atomic number. There existed certain inherent defects which opposed the system.

15.5 MODERN APPROACH

Mosely in 1913 showed that the properties of the elements are periodic function of their atomic numbers. He arranged all the elements according to the increasing atomic number.

> **MODERN PERIODIC LAW:** The properties of the elements are periodic function of their atomic numbers.

Atomic number itself was discovered in 1913 by a team lead by Mosely. Table based on atomic number is term as Modern Periodic Table

15.6 LONG FORM OF PERIODIC TABLE

The arrangement of elements in the long form of periodic table is a perfect matching of electronic configuration of the elements on one hand and physical and chemical properties of the elments on the other. Some important considerations of the modern atomic structure applied to the classification are discussed below :

* Many of the symbol are not in use now.

(i) An atom looses electrons from or gains electrons in the outermost orbit of an atom during a chemical reaction.

(ii) The sharing of an electron or electrons by an atom with other atom or atoms is largely through the outer most orbit. Thus the electrons in the outermost orbit of an atom largely determine the chemical properties of the elements.

We may therefore conclude that the elements possessing identical outer electronic configuration should possess similar plysical and chemical properties and therefore they should be placed together for an easy and systematic study.

Keeping in mind the reasoning given above, all the known elements when arranged in a table according to their increasing atomic number, the properties of the elements show periodic function (reappear at definite intervals). The table so developed is given as 15.2

15.7 STRUCTURAL FEATURES OF THE LONG FORM OF PERIODIC TABLE

(i) In this table there are 18 vertical columns called groups. They are numbered from 1 to 18. Every group has a unique configuration.

(ii) There are seven rows of boxes filled with elements. These rows are called **PERIODS** Thus the periodic table has seven periods, numbered from 1 to 7.

(iii) There are a total of 110 elements known. Of all the elements known 90 are naturally occurring and others are made through nuclear transformations or are synthesised artificially, Either way they are **Manmade Elements**, but you will find the term specifically applied to transuranic elements only (elements listed after uranium)

(iv) First period consists of only two elements (very short period). Second and third periods consists of only eight elements each (short periods). Fourth and Fifth period consists of 18 elements each (long periods). Sixth period consists of 32 elements. (long period). Seventh period yet incomplete and more and more elements are likely to be added as the scientific research advances.

(v) There are also nick names given to the groups or a cluster of groups on the basis of the similarity of their properties, e.g., transition metals.

Group 1 elements except hydrogen, are call ALKALI METALS

Group 2 elements are called ALKALINE EARTH METALS.

Group 3 to 12 elements are called Transition metals.

Group 16 elements are called oxygen and Sulphur (CHALCOGENS)

Group 17 elements are called Halogens

Group 18 elements are called NOBLE GASES, (Inert Gases)

Apart from what has been said above elements with atomic numbers 58 to 71 are called LANTHANIDES or INNER TRANSITION ELEMENTS — First series. Elements from atomic numbers 90 to 103 are called ACTINIDES or INNER TRANSITION ELEMENTS — Second series. All elements except transition or inner transition elements are also collectively called MAIN GROUP ELEMENTS.

GROUPS

REPRESENTATIVE ELEMENTS

Legend	
20	← Atomic number (Z)
Calcium	← Name of element
Ca	← Symbol of element
40.08	← Atomic weight or mass number (A)

TRANSITION ELEMENTS

PERIODS	Group IA 1	2	IIIA 3	IVA 4	VA 5	VIA 6	VIIA 7	VIIIA 8	VIIIA 9	VIIIA 10	IB 11	IIB 12	13	14	15	16	17	18 Noble Gases	
P 1	1 Hydrogen H 1.0079	IIA 2											IIIB 13	IVB 14	VB 15	VIB 16	VIIB 17	2 Helium He 4.00260	2
E 2	3 Lithium Li 6.941	4 Beryllium Be 9.01218											5 Boron B 10.81	6 Carbon C 12.011	7 Nitrogen N 14.0067	8 Oxygen O 15.9994	9 Fluorine F 18.9984	10 Neon Ne 20.179	2 8
R 3	11 Sodium Na 22.9898	12 Magnesium Mg 24.305											13 Aluminium Al 26.9815	14 Silicon Si 28.0855	15 Phosphorus P 30.9738	16 Sulfur S 32.06	17 Chlorine Cl 35.453	18 Argon Ar 39.948	2 8 8
I 4	19 Potassium K 39.0983	20 Calcium Ca 40.08	21 Scandium Sc 44.9559	22 Titanium Ti 47.88	23 Vanadium V 50.9415	24 Chromium Cr 51996	25 Manganese Mn 54.9380	26 Iron Fe 55.847	27 Cobalt Co 58.9332	28 Nickel Ni 58.69	29 Copper Cu 63.546	30 Zinc Zn 65.39	31 Gallium Ga 69.72	32 Germanium Ge 72.59	33 Arsenic As 74.9216	34 Slenium Se 78.96	35 Bromine Br 79.904	36 Krypton Kr 83.80	2 8 18 8
O 5	37 Rubidium Rb 85.4678	38 Strontium Sr 87.62	39 Yttrium Y 88.9059	40 Zirconium Zr 91.224	41 Niobium Nb 92.9064	42 Molybdenum Mo 95.94	43 Technetium Tc 98.91	44 Ruthenium Ru 101.07	45 Rhodium Rh 102.906	46 Palladium Pd 106.42	47 Silver Ag 107.868	48 Cadmium Cd 112.41	49 Indium In 114.82	50 Tin Sn 118.71	51 Antimony Sb 121.75	52 Tellurium Te 127.60	53 Iodine I 126.905	54 Zenon Xe 131.29	2 8 18 18 8
D 6	55 Cesium Cs 132.905	56 Barium Ba 137.33	57 Lanthanum La 138.906	72 Hafnium Hf 178.49	73 Tantalum Ta 180.948	74 Wolfram (Tungsten) W 183.85	75 Rhenium Re 186.207	76 Osmium Os 190.2	77 Iridium Ir 192.22	78 Platinum Pt 195.08	79 Gold Au 196.967	80 Mercury Hg 200.59	81 Thallium Tl 204.383	82 Lead Pb 207.2	83 Bismuth Bi 208.980	84 Polonium Po (209)	85 Astatine At (210)	86 Radon Rn (222)	2 8 18 32 18 8
S 7	87 Francium Fr (223)	88 Radium Ra 226.025	89 Actinium Ac 227.028	104 Unnil-quadium Unq (261)	105 Unnil-pentium Unp (262)	106 Unnil-hexium Unh (263)	107 Unnil-septium Uns (262)	108 Unnil-octium Uno (265)	109 Unnil-ennium Une (266)	110	111	112	113	114	115	116	117	118	2 8 18 32 32 18 8

Series															
Lanthanide Series 6*	58 Cerium Ce 140.12	59 Praseo-dymium Pr 140.908	60 Neodymium Nd 144.24	61 Promethium Pm (145)	62 Samarium Sm 150.36	63 Europium Eu 151.96	64 Gadolinium Gd 157.25	65 Terbium Tb 158.925	66 Dysprosium Dy 162.50	67 Holmium Ho 164.930	68 Erbium Er 167.26	69 Thulium Tm 168.934	70 Ytterbium Yb 173.04	71 Lutetium Lu 174.967	2 8 18 32 9 2
Actinide Series 7**	90 Thorium Th 232.038	91 Protactinium Pa 231.03	92 Uranium U 238.029	93 Neptunium Np 237.048	94 Plutonium Pu 239	95 Americium Am (243)	96 Curium Cm (247)	97 Berkelium Bk (247)	98 Californium Cf (252)	99 Einsteinium Es (252)	100 Fermium Fm (257)	101 Mendelevium Md (258)	102 Nobelium No (259)	103 Lawrencium Lr (260)	2 8 18 32 32 9 2

Table 15.2: Long from of Periodic Table.

15.8 POSITION OF METALS, NON-METALS AND METALLOIDS

In order to locate the position of metals, non-metals and metalloids in the periodic table, you may draw a diagonal line joining the element Boron (At no. 5) with that of tellurium (At no. 52) and passing through silicon and arsenic. Now we are in a position to make statements in the following manner.

(i) The elements that are above the diagonal line and to the right are non-metals (except selenium which shows slightly metallic character also.) The non-metallic character is more marked the farther an element is from the diagonal line and up.

(ii) The elements that are below the diagonal line and to the left are metals. The metallic character is more marked the farther an element is from the diagonal line and down. All lanthanides and actinides are metals (Hydrogen is a non-metal and is an exception)

(iii) The elements that fall under the diagonal line are metalloids and possess the characteristics of metals as well as non-metals. In addition. Germanium, Antimony and Selenium also the characteristics of metalloids.

INTEXT QUESTIONS: 15.1

1 Classify the elements of group 14, 15 and 16 into metals, non-metals and metalloids.

....

2. Compare the metallic character of Aluminium and Potassium.

....

3. Name the group number for the following type & elements
   (i) Alkaline earth metals
   (ii) Alkali metals
   (iii) Transition metals
   (iv) Halogens
   (v) Noble gases.

....

4. Name five man made elements.

....

15.9 CATEGORISATION OF ELEMENTS INTO 's', 'p', 'd', and 'f' BLOCKS.

Grouping of elements in the periodic table can be done in another systematic way, which is more related to their electronic configuration. Under this categorisation, the location of the

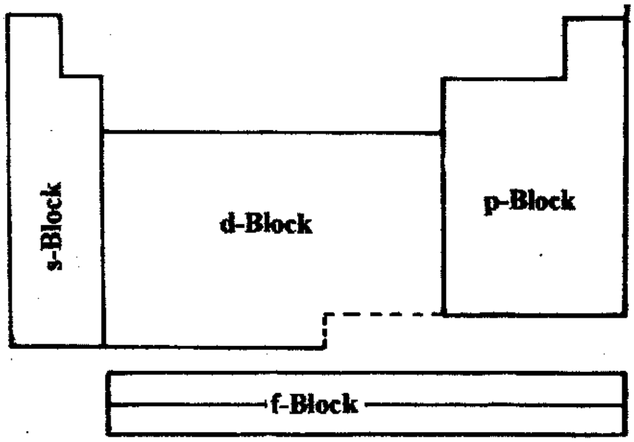


Table 15.3a: Blockwise categorization of Elements.

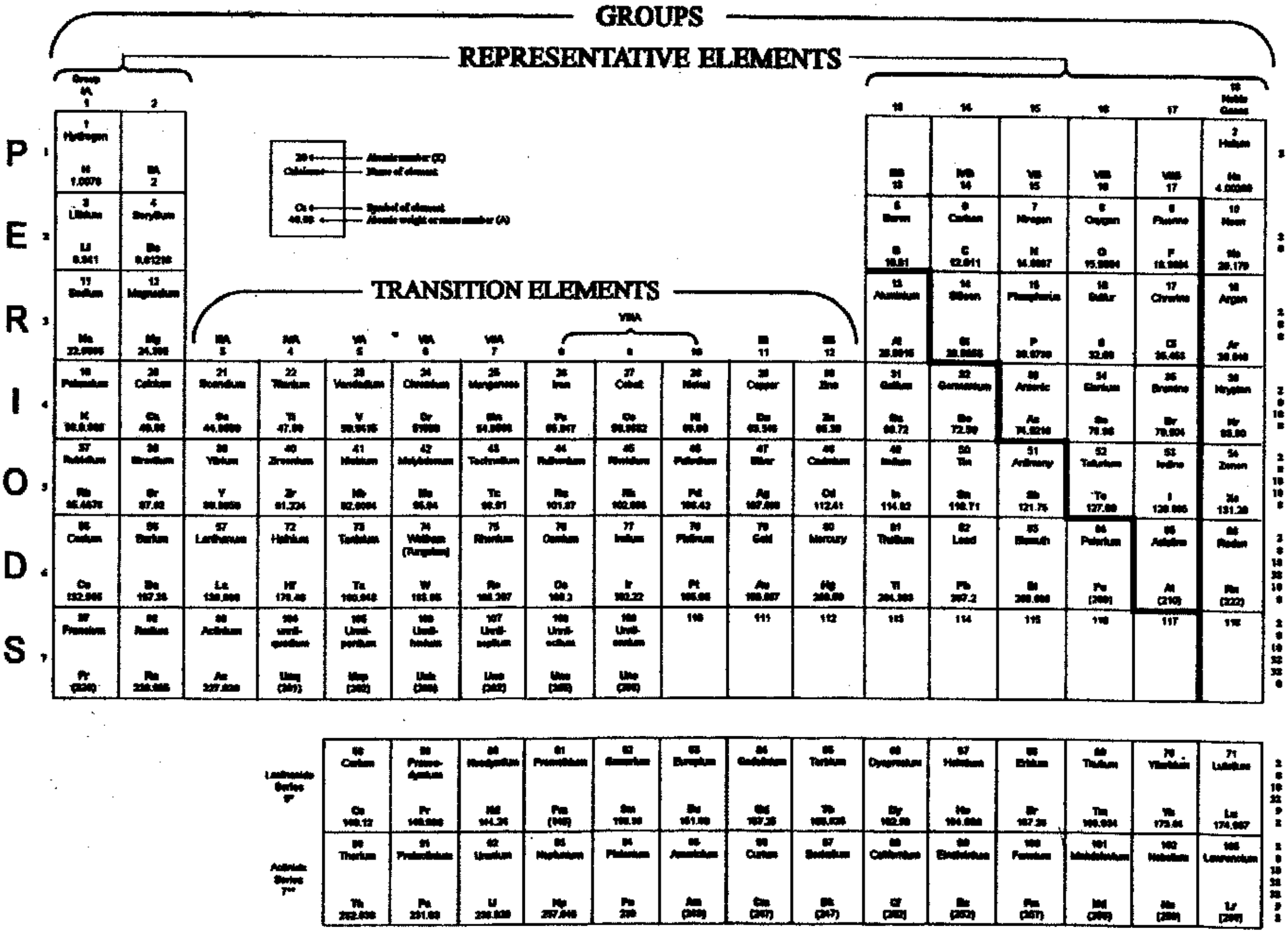


Table 15.3b: Blockwise filled Periodic table.

differentiating electron (the last electron) is most important. If for example, the electron has gone to 's' sub-shell, may be s^1 or s^2, the elements will fall in 's' block and if the last electron goes to 'p' sub-shell, then the element will belong to 'p' block. Similarly if the differenciating electron enters in the 'd' sub-shell, of an atom, then the elements comprising all such atoms will belong to 'd' block.

There are minor exceptions at Mn and Zn configurations. You will study more about the reasons for such exceptions in other chapters.

The grouping of elements explained above can be related to the type of elements discussed earlier:

(i) s-block elements : All alkali metals, alkaline earth metals & hydrogen

(ii) p-block elements : All elements of group no. 13 to group no. 18.

(iii) d-block elements : All elements from group no .3 to group no. 12 except Lanthanides and Actinides.

(iv) f-block elements : Lanthanides (at. no. 58 to 71) and Actinides (at. no. 90 to 103)

This is shown in Tables 15.3 a and 15.3b on page 7 of this Unit.

15.10 PERIODICITY IN ATOMIC PROPERTIES

The term periodicity is used to indicate that some characteristic. properties reappear in the periodic table after moving definite intervals, again and again, however with a varying magnitude. Thus after starting from a certain point on the periodic table, we are almost certain that the movement in a particular direction will show steady increase or decrease of a said property. For example, on moving from left to right in a period we are likely to see decrease in the size of an atom. The maxima and minima in a property will be visible through the graphicrepresentations of the property (fig: 15.1)

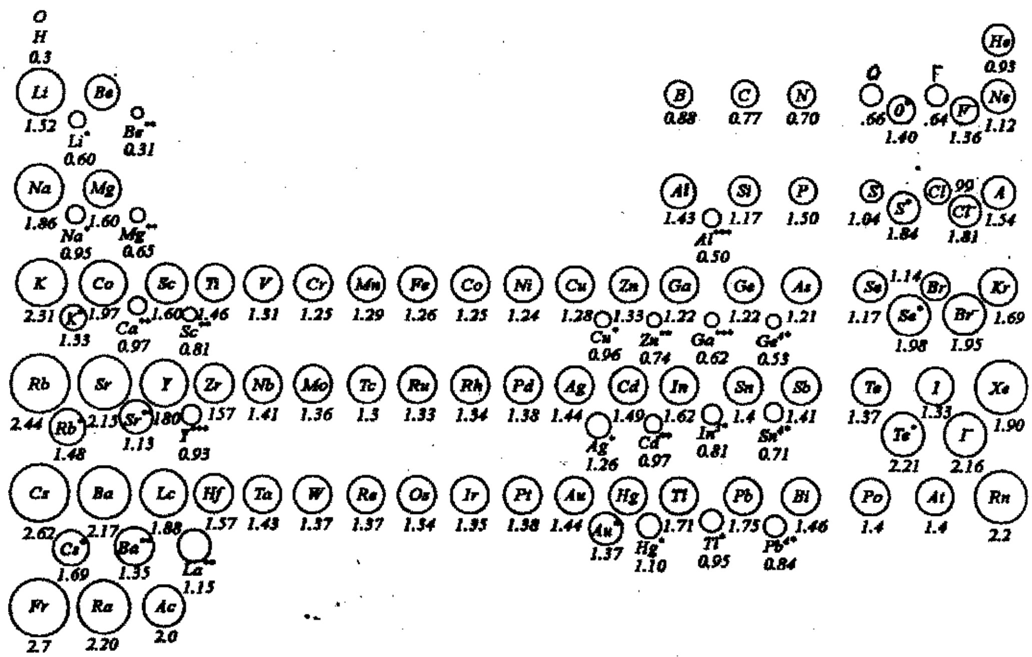


Fig. 15.1: The Compbell periodic table showing atomic and ionic sizes.

Some structural features, like size of an atom is closely linked to the ease with which the most loosely bound electron can be completely removed away from an atom. This ease will determine the chemical reactivity of the element also. Thus we see than the structural feature, the physical property and the chemical behaviour are closely related.

The periodic variation will acquaint us with the trends of variation and we shall thus form our opinion on their properties. We shall now discuss in some more detail the variation of atomic size, ionic size ionization energy, electron affinity and melting points of elements in the periodic table.

15.11 ATOMIC SIZE

A glance at Fig. 15.1 gives an idea about the variation in size of an atom as we move in the periodic table. The size of an atom is visualized from the radius of the atom, which is supposed to be spherical.

In homonuclear molecules the distance from the centre of one nucleus to the centre of another nucleus gives the bond length and half of this bond length is atomic radius. (Fig 15.2) The first member of each period is the largest in size. Thus we can say that the group 1 atoms are the largest in their respective horizontal rows. Similarly, atoms of group 2 elements are large but are definitely smaller than the corresponding atoms of group 1. This is due to the reason that the extra charge on the nucleus draws the electrons inward resulting in smaller size for the atoms under reference. This trend of decrease in size of atoms, as we move from left to right continues. An example is shown in fig. 15.3. However there may be some exceptions but then there will be other reasons to explain them.

Atomic radius = ½ AA = r

Fig. 15.2

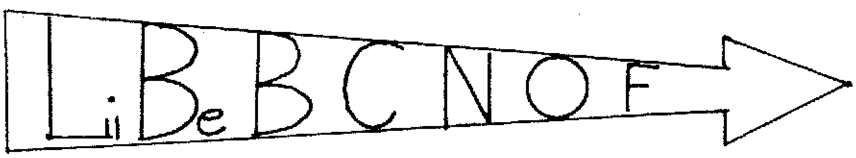


Fig. 15.3: From left to right, size of atoms decrease in the P.T.

In going down the group of elements (in any particular column) the atomic size increases at each step:

This increase may be explained in terms of a new electron shell being added, when we pass from one element to another in a group.

15.12 IONIC SIZE

An ion is formed when an atom undergoes loss or gain of electrons.

$$M (g) \longrightarrow M^{+} (g) + e^{-} \quad \text{(cation formation)}$$
$$M (g) + e^{-} \longrightarrow M^{-} (g) \quad \text{(anion formation)}$$

> A cation is always smaller than its atom and an anion is always bigger than its atom.e.g., Na+ is smaller than Na, Cl^- is bigger than Cl.

A cation is formed when an atom looses the most loosely bound electron from its outermost orbit. The atom acquires a positive charge and becomes an ion (a cation). A cation is smaller than its atom. On the removal of an, electron, the positive charge of the nucleus acts on lesser number of electrons than in the neutral atom and thus greater pull is exerted by the nucleus, resulting in a smaller size of the cation.

An anion is bigger than its atom because on receipt of an electron in the outermost orbit the number of negative charges increase and it outweighs the positive charges. Thus the hold of the nucleus on the shells decrease resulting in an increase in the size of the anion.

INTEXT QUESTION 15.2

Refer the periodic table (Table 15.2) and answer the following:-

(i) Name two 's' block elements, one of which should be a non-metal.

....

(ii) Name a 'p' block element which is a halogen.

....

(iii) Name a 'p' block element which is neither an inert gas nor a hologen.

....

15.13 IONIZATION ENERGY

Ionization energy is the energy required to remove the most losely bound electron from an isolated atom in the gaseous state for one mole of the substance. It is expressed in kJ mol^{-1} (kilojoules per mole).

As we move from left to right in the periodic table, there is a nearly regular increase in the magnitude of the ionization energy of elements.

Similarly, on moving down a group the magnitude of the ionization energy indicates a regular decline. The ionization energy of the first member of any (group on the top) is the highest within that group and the ionization energy of the last member in the same group is the least. This is shown an table 15.4

Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H • 1311																	He • 2372
2	Li • 520	Be • 899											B • 801	C • 1086	N • 1403	O • 1410	F • 1681	Ne • 2081
3	Na • 496	Mg • 737											Al • 577	Si • 786	P • 1012	S • 999	Cl • 1255	Ar • 1521
4	K • 419	Ca • 590	Sc • 631	Ti • 656	V • 650	Cr • 652	Mn • 717	Fe • 762	Co • 758	Ni • 736	Cu • 745	Zn • 906	Ga • 579	Ge • 760	As • 947	Se • 941	Br • 1142	Kr • 1351
5	Rb • 403	Sr • 549	Y • 616	Zr • 674	Nb • 664	Mo • 685	Tc • 703	Ru • 711	Rn • 720	Pd • 804	Ag • 731	Cd • 876	In • 558	Sn • 708	Sb • 834	Tc • 869	I • 1191	Xe • 1170
6	Cs • 376	Ba • 503	La • 541	Hf • 760	Ta • 760	W • 770	Re • 759	Os • 840	Ti • 900	Pt • 870	Au • 889	Hg • 1007	Tl • 589	Pb • 715	Bi • 703	Po • 813	At • 912	Rn • 1037
7	Fr	Ra	Ac															

Table 15.4: FIRST IONIZATION ENERGIES OF THE ELEMENTS (in kJ mol^{-1})

The variation in the magnitude of ionization energy of elements in the periodic table is mainly dependent on the following factors:

(a) The size of the atom
(b) The magnitude of the nuclear charge on the atom,
(c) The extent of screening
(d) The type of electron involved (s, p, d, or f.)

These factors taken together contribute largely to decide the extent of the force of attraction between the nucleus and the electrons around it. The resultant of these factors thus determine the magnitude of ionization energy of any element. You can see the variation in the magnitude of the ionization energy of elements with atomic number in the Fig 15.4

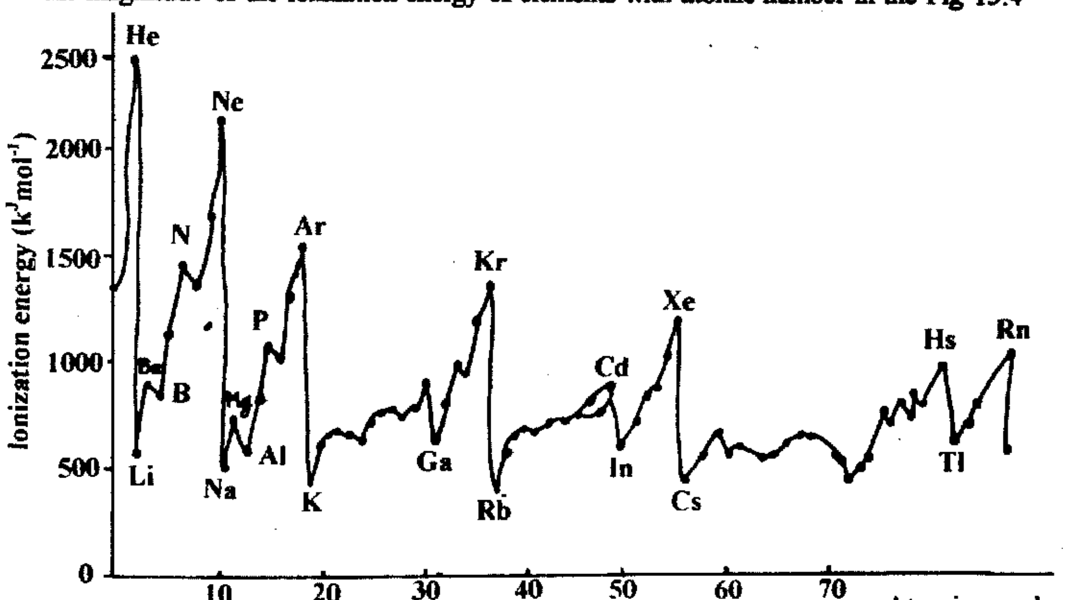


Fig. 15.4: Variation of ionization energy of elements.

> Ionization energy is the energy required to remove the most loosely bound electron from an atom in the gaseous state for one mole of the substance. It is an absolute value and can be determined experimentally.

15.14 ELECTRON AFFINITY

Every atom, in general, has the tendency to gain or loose electrons in order to acquire a noble gas configuration. The atoms which have five, six or seven electrons in their outermost orbit show tendency to accept electrons andachieve the nearest noble gas configuration. Halogens, for example, have seven electrons in their outermost orbit. Thus they show a tendency to accept one more electron and achieve the nearest noble gas configuration. The energy change (ΔE) for this process is called **electron affinity** of that atom:

$$X(g) + e^{-} \longrightarrow X^{-}(g).$$

where X represents an atom. Electron affinity can be position or negative. We can represent any element for this energy change.

$$\underset{1\ mol}{Cl(g)} + e^{-} \longrightarrow \underset{1\ mol}{Cl^{-}(g)}:\ \Delta E = -349\ kJ\ mol^{-1}$$

The electron affinity increases in negative value as we move from left to right in a period. This is because it is easier to add an electron to a smaller atom since the added electron would be closer to the positively charged nucleus. The negative value shows the release of energy and hence tendency to greater stabilization. Halogens release maximum energy when they accept an electron. On the other hand, metals do not accept electrons and show a high positive value for ΔE.

Electron affinities decrease in negative values, as we go down the group showing that the positive character of the atoms increase on the same trend. This is because the size of the atom increase as we go down the group and the electron added goes the higher shells. Electron affinity values for some elements are shown in table 15.5, alongwith their position in the periodic table

	← GROUPS →							
PERIOD	1	2	13	14	15	16	17	18
1	H – 73							He (0)
2	Li – 59.6	Be (0)	B – 26.7	C – 154	N – 7	O – 111	F – 328	Ne (0)
3	Na – 53						Cl – 349	Ar (0)
4	K – 48						Br – 325	Kr (0)
5	Rb – 47						I – 295	Xe (0)
6								Rn (0)

Table 15.5: Electron affinities in kJ mol^{-1}

Electron Affinity is the energy released or absorbed for one mole of neutral atoms in a gaseous state when electron in accepted by each atom.

15.15 ELECTRONEGATIVITY

It is an indicator of the extent of attraction by which electrons of the bond pair are attracted in a diatomic molecule. The value of electronegativity is assigned arbitrarily to one atom such as hydrogen. Then the value of electronegativity is assigned to all other atoms with respect to hydrogen. On such scale is the Pauling Scale of electronegativity (Table 15.6)

Electronegativity is defined as a measure of the ability of an atom to attract the electron pair in a co-valent bond to it self.

In a homonuclear diatomic molecule such as hydrogen (H_2) or Fluorine (F_2), the electron pair of the co-valent bond in each molecule experiences equal attraction by each atom. Thus none of the two atoms are able to shift the bond pair of electrons to it self. However in a hetronuclear diatomic molecule, the bond pair electrons get shifted towards the atom which is more electronegative than the other. For example, in HF or HCl the bond pair of electrons are not shared equally but the more electronegative atom is able to shift the bond pair towards itself, resulting in the polarization of the molecule.

A large difference between electronegativities of the two atoms indicate highly ionic character(ion pair).For example Cs^+F^-. On the other hand, zero difference in the electronegativities between the two atoms indicate that the percentage ionic character is zero therefore the molecule is purly covalent e.g.H_2, Cl_2,N_2 etc. These molecules are non-polar.

Li	Be	B	C	N	O	F
1.0	1.5	2.0	2.5	3.0	3.5	4.0
Na	Mg	Al	Si	P	S	Cl
0.9	1.2	1.5	1.8	2.1	2.5	3.0
K	Ca	Sc	Ge	As	Se	Br
0.8	1.0	1.3	1.8	2.0	2.4	2.8
Rb	Sr	Y	Sn	Sb	Te	I
0.8	1.0	1.3	1.7	1.8	2.1	2.5
Cs	Ba					
0.7	0.9					

Table : 15.6: Electronegatives of elements on Pauling scale.

The most electronegative elements have been placed on the farthest right hand, upper corner (noble gases are not included). The value of electronegativity decreases as we go down in any group or move in a period towards left. Thus fluorine is the most electronegative and caesium is the least electronegative element. (We have not considered Francium, being radioactive.)

INTEXT QUESTIONS 15.3

1 How much energy is required (in kJ) to ionise all the atoms in 23g of sodium atoms (I E of Na is = 496 kJ mol^{-1}).

....

2 Which species, in each pair is expected to have higher ionisation energy.

(i) $_{3}Li$, $_{11}Na$ (ii) $_{7}N$, $_{15}P$

(iii) $_{20}Ca$, $_{12}Mg$ (iv) $_{13}Al$, $_{14}Si$

(v) $_{17}Cl$, $_{18}Ar$ (vi) $_{18}Ar$ and $_{19}K$

(vii) $_{13}Al$, $_{14}C$

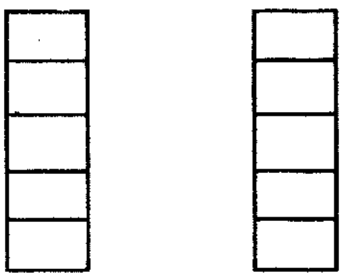
....

3. Calculate electron affinity for 355g of chlorine atoms. you may use the table given above.

....

4. Arrange the following elements into different categories on the basis of their outer orbit configuration and then put them in Vertically arranged boxes. All elements in a particular vertical column should have similar outer orbit configuration. (Atomic numbers are given with hypothetical symbol X)

$_{9}X$, $_{11}X$, $_{17}X$, $_{19}X$, $_{3}X$, $_{35}X$, $_{37}X$, $_{53}X$, $_{55}X$, $_{85}X$



....

15.16 WHAT YOU HAVE LEARNT

- The classification of elements make their study systematic.
- The arrangement of elements in the long form of the periodic table depends on their electronic configuration.
- All the known elements are arranged in 18 groups in the long form of periodic table.

- There are seven horizontal rows (periods) in the long from of the periodic table.
- Elements of groups 1 and 2 are known as alkali metals and alkaline earth metals respectively.
- Elements of groups 17 and 18 an known as halogens and noble gases respectively.
- s, p, d and f are the four blocks in the periodic table classified on the basis of their orbit most electron residing in s, p, d, or f sub-shell.
- The properties of the elements are the periodic function of their atomic number.
- The elements can be classified into metals, non-metals and metalloids on the basis of their proprieties and their position in the periodic table.
- The atomic size, ionic size, ionization energy, electron affinity and electronegativity show regular trends along a group and a period.

15.17 TERMINAL EXERCISE

1. Define modern periodic law.
2. Refer the periodic table given in Table 15.2 and answer the following questions.
   (i) The elements placed in group number 18 are called ....
   (ii) Alkali and alkaline earth metals are collectively called ....block metals.
   (iii) The general configuration for halogens is...................
   (iv) Name a p-block element which is a gas other than a noble gas or a halogen.
   (v) Name the groups that comprise the 's' block of elements.
   (vi) Element number 118 has not yet been established, to which block, will it belong tb?
   (vii) How many elements should be there in total if all the 7s, 7p, 6d and 5f, blocks have to be full.?
3. Describe the variation of Electron Affinity and lonization Energy in the periodic table.
4. Define the following:
   (a) Electron affinity (b) Ionization energy
   (c) Ionic radius (d) Electronegativity.

CHECK YOURE ANSWERS

INTEXT QUESTIONS 15.1

1.

Metals	Non metals	Metalloids
Sn, Pb	C	Si, Ge.
Sb, Bi	N, P	As,
Te, Po,	O, S	Se,

2. Potassium is more metallic than aluminium.

3. (i) 2 (ii) 1 (iii) 3 to 12 (iv) 17 (v) 18

4. Np, Lw, No, Kw, Ha.

INTEXT QUESTIONS 15.2

(I) Lithium and calcium

(ii) Chlorine

(iii) Oxygen.

INTEXT QUESTIONS 15.3

1. 496 kJ mol^{-1}

2. (i) $_{3}Li$ (ii) $_{12}Mg$ (iii) $_{7}N$

(iv) $_{14}Si$ (v) $_{18}Ar$ (vi) $_{14}C$

3.

$_{9}X$	$_{11}X$
$_{17}X$	$_{19}X$
$_{35}X$	$_{3}X$
$_{53}X$	$_{87}X$
$_{85}X$	$_{55}X$

TERMINAL EXERCISE

1 When elements are arranged according to their increasing atomic number. the properties of the elements show periodic function.

2. (i) Inert gases
   (ii) 's block elements
   (iii) $ns^2 np^5$
   (iv) Oxygen/nitrogen
   (v) Alkali metals and Alkaline earth metals.
   (vi) p-block
   (vii) 118

3. Electron affinity increases as we move to the right and up in the periodic table lonization Energy becomes lesser as we move to the left and down but increases as we move to the right and up in the periodic table.

4. See text.