UG PART 1,(SUBSIDIARY)

Periodic properties of Elements

Elements in the periodic table are arranged in order of increasing atomic number. The elements display periodic properties in terms of atomic radius, ionisation energy, electron affinity, electronegativity metallic character and other properties. These properties are functions of their atomic number. However atomic number in itself was not adequate to explain the periodicity in properties of the elements. A more satisfactory concept known as effective nuclear charge or screening effect of the inner electrons on the nuclear charge was developed. It arose out of the uncertainty principle and the wave nature of the electrons.

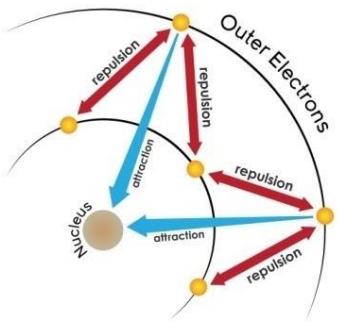
Effective nuclear charge

Effective nuclear charge is the attractive positive charge of the nuclear protons acting on the valance electrons of multi-electron atoms. The electrons are arranged in shells. There occurs electron- electron repulsion as well as nuclear protons and electrons attraction (fig. 1).The inner electrons shield the nuclear charge more effectively than the outer electrons. A precise calculation of Zeff is prohibitive because of uncertainty principle $\Delta X \times \Delta p = h/2\pi$ arising out of the wave character of electron. In other words the probability of finding the electron at a particular space is uncertain. In a simple and straight forward way we define an approximate effective nuclear charge as Zeff = Z- S where Z is the atomic number and S is the number of electrons in the inner closed shell. Under this formulation it is assumed that effective nuclear charge on all the electrons in the same shell is same.

For example, the electronic configuration of sodium atomic number Z = 11 is $1s^2\ 2s^2\ 2p^6\ 3s^1$

Zeff = Z-S = 11-10 = 1^+

The closed shell electronic configuration can be found by looking at the immediately preceding inert gas configuration. For Sodium electronic configuration is [Ne] $3s^1$.



In reality effective nuclear charge is always less than the calculated value as shown above. It is because inner electrons are attracted more strongly by the nuclear charge. They shield or block the nuclear charge more effectively. Effective nuclear charge is a dominant factor behind all periodic properties of the elements.

(Fig. 1)

For aluminum (Z=13) the electronic configuration is [Ne] $3s^2\ 3p^1$

Aluminum has the electronic structure $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

So, for aluminum Zeff = 13-10 = 3^+

For chlorine (Z=17) the electronic configuration is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

So for chlorine, Zeff =17-10 = 7^+

As the effective nuclear charge increases, its influences on the periodic properties like

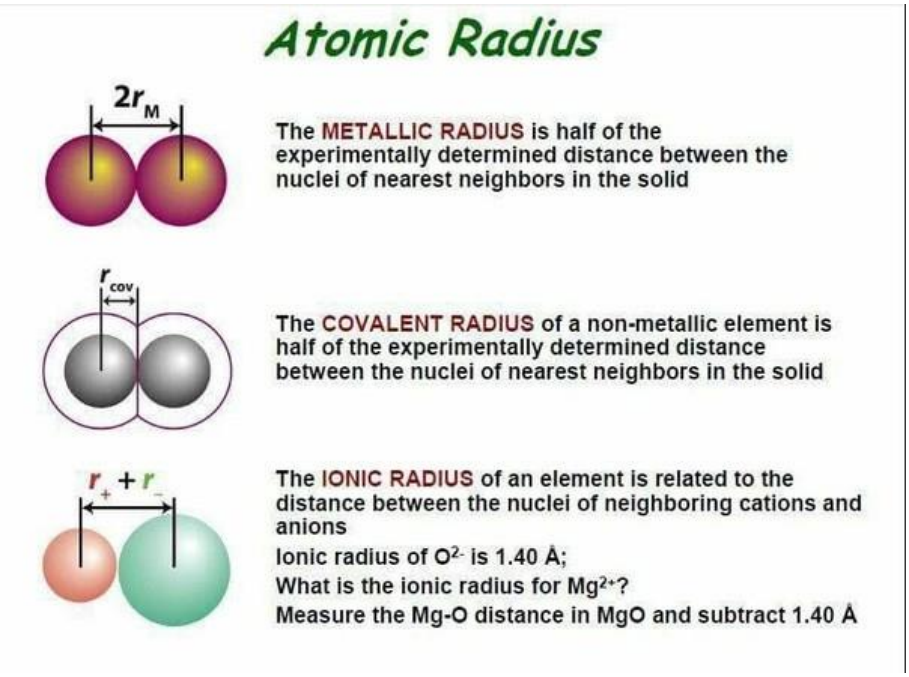
atomic size, ionic size, metallic radius, ionisation energy, electron affinity, electronegativity and several other properties becomes visibly significant. They are briefly discussed below.

Atomic radius

It is not possible to determine the atomic radius accurately because there is never a zero probability of finding an electron. All that is determined is the distance between the two nuclei.

A covalent radius is one-half of the distance between two nuclei of two identical atoms.

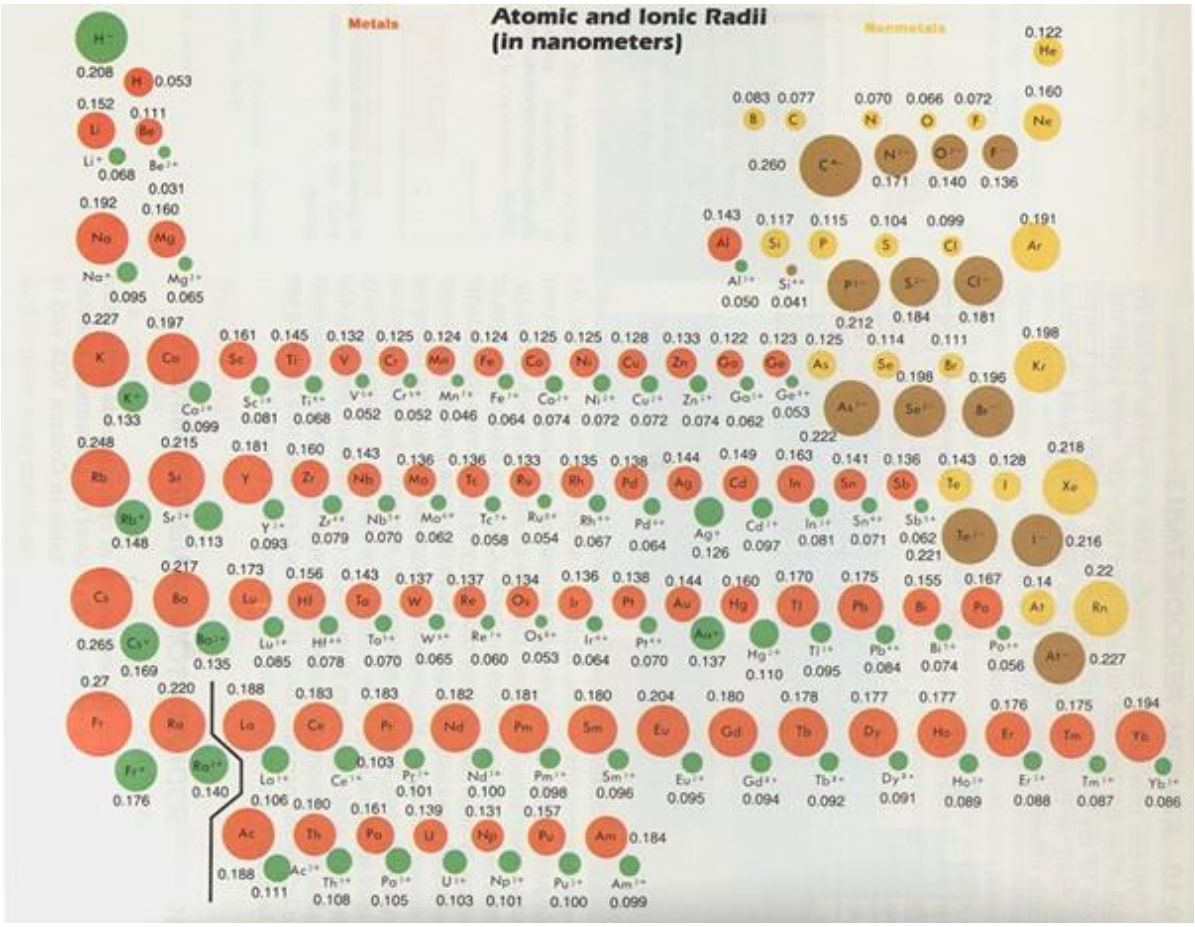
An ionic radius is one half of the distance between the nuclei of two in an ionic bond. The distance is apportioned to the smaller cation and larger anion.



(Fig. 2)

The metallic radius is one-half the distance between the nuclei of two adjacent atoms of the crystalline structure of the metal.

As regards noble gases are concerned the experimental data of atomic radii values have become a matter of contest among scientists. So they are left out of the trends in atomic radii. The S.I. unit of atomic radii are given in nanometre (10^{-9} m) or picometre (10^{-12} m). The periodic trend of atomic and ionic radii data are figuratively presented in (Table. 1)



(Table. 1)

Ionisation Energy

The 1st ionisation energy is given in (table 2). It can be seen that in general (1) the first ionisation energy increases as we go across a row in periodic table and (2) decreases as we go down a column in the periodic table.

We expect the 1st ionisation energy to be larger as we go across the row because the force of attraction between the nucleus and an electron becomes larger as the number of protons in the nucleus of the atoms increases, the electron shells remaining the same. The 2nd observation is the consequence of the fact that the principal quantum number of the orbital holding the outermost electron becomes larger as we go down a column. Though the number of protons in the nucleus increases the electrons in inner shells and sub-shells screen the outer most electrons from some of the force of the attraction of the nucleus.

Subsequent ionisations can be considered in a similar way.

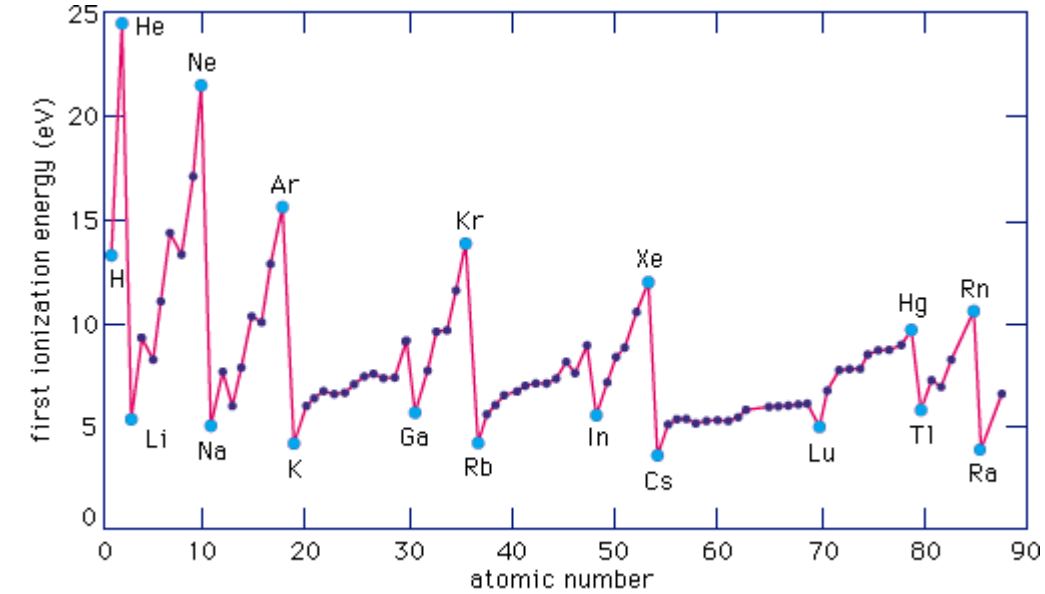
$M = M^+ + e^-$ *(First ionisation)*

$M^+ = M^{2+} + e^-$ *(Second ionisation)*

$M^{2+} = M^{3+} + e^-$ *(Third ionisation)*

Subsequent ionisations become more and more difficult and higher in energy as a

consequence of increase in effective nuclear charge. They are shown in (Fig. 3) and (Table. 2)



(Fig. 3)

Element	Atomic Number (Z)	Atomic radius (A)	Ionisation potentials (in eV)							
			I_1	I_2	I_3	I_4	I_5	I_6	I_7	I_8
Hydrogen	1		13.595							
Helium	2		24.580	54.40						
Lithium	3	1.225	5.390	75.61	122.42					
Berylium	4	0.869	9.320	18.20	153.85	217.65				
Boron	5	0.80	8.296	25.14	37.92	259.29	340.12			
Carbon	6	0.771	11.264	24.57	47.86	64.47	391.98	489.84		
Nitrogen	7	0.74	14.54	29.60	47.42	74.45	97.86	551.92	666.83	
Oxygen	8	0.74	13.614	35.14	54.93	77.39	113.81	138.08	739.11	871.12
Fluorine	9	0.72	17.42	34.98	62.64	87.23	114.21	157.11	185.13	953.60
Neon	10		21.559	41.07	64.0	97.16	126.4	157.91		
Sodium	11	1.572	5.138	47.29	71.65	98.86	138.6	172.36	208.44	264.15
Magnesium	12	1.364	7.644	15.03	80.12	109.26	141.23	186.86	225.31	265.97
Aluminium	13	1.248	5.984	18.82	28.44	119.96	153.77	190.42	241.93	285.13
Silicon	14	1.173	8.149	16.34	33.46	45.13	166.75	205.11	246.41	303.87

(Table. 2) Successive ionisation potential of some elements