PERIODIC PROPERTIES

ATOMIC AND IONIC RADII

The term atomic or ionic radius is generally used for the distance between the nucleus and the outer most shell of electrons of the atomic or ionic particle.

It is not possible to isolate an individual atoms or an ion. Therefore these quantities are derived indirectly. These values are obtained by measuring the distance between the nuclei of two bonded atoms in a gaseous molecule (internuclear distance) or between the nuclei of two ions in crystals or solids (interionic distance).

Periodic variations of atomic and ionic radii

(a) Variation in a period

Across a period, both atomic and ionic radii decrease from left to right in the periodic table.

For example in the elements of 2^{nd} period, the covalent radii (Atomic radii) decrease as we move from Li to F as shown below

Elements of 2^{nd} period	:	Li	Be	B	C	N	O	F
Atomic radii (Å)	:	1.23	0.90	0.82	0.77	0.75	0.73	0.7

$\longrightarrow$ values decreasing $\rightarrow$

Therefore in any period, the alkali metals (left) have the largest size while the halogens (extreme right) have the smallest size.

Reason

We know that as we proceed from left to right in a period, the electrons are added to the orbitals of the same main energy level. Addition of differentiating electrons to the same main energy level puts the electron no farther from the nucleus and hence cannot add to the size.

(b) variation in a group

On moving down a group, both atomic and ionic radii increases with increase in atomic number.

For example in the elements of II A group, both covalent and ionic radii of M^{2+} ions increase when we pass from Be to Ba.

Elements of II A Group	Be	Mg	Ca	Sr	Ba
Atomic radii	0.90	1.36	1.74	1.91	1.98
Ionic radii	0.31	0.65	0.99	1.13	1.35

Explanation

(i) On proceeding downwards in a group, the electrons are added to higher main energy level which are farther from the nucleus. This effect decreases the electrostatic attraction between the nucleus and valance-shell electrons and this decreased electrostatic attraction increases the atomic and ionic ra

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IONISATION POTENTIAL OR IONISATION ENERGY

Definition

The amount of energy required to remove the most loosely bound electron (i.e the outermost electron) from an isolated gaseous atom of an element in its lowest energy state (i.e ground state) to produce a cation is known as ionisation potential or ionisation energy of that element.

It is represented as I or IP and is measured in electron volts (eV) or Kilo calories (K. calories) per gram atom.

The ionisation potential can be expressed as:

$$M_{(g)} + \underbrace{\text{energy supplied}}_{\text{(Ionisation energy)}} \longrightarrow \underbrace{M^+_{(g)}}_{\text{cation}} + e^-$$

Factors affecting the Magnitude of ionisation potential and its periodic variations.

The magnitude of ionisation potential depends on the following factors:

(i) charge on the nucleus and variation in a period

The greater charge on the nucleus of an atom the more difficult it would be to remove an electron from the atom and hence greater would be the value of ionisation potential. Thus the value of ionisation potential generally increases in moving from left to right in a period, since the nuclear charge of elements (i.e atomic number) also increases in the same direction.

(ii) The degree of penetration of electron

For a given value of n, the degree of penetration of electrons will decrease in the order: $s > p > d > f$. This means that a s-electron will approach the nucleus more closely than p-electron, a p-electron more closely than d and d more closely than f. Thus other factors being equal, an s-electron will be harder to remove than p-electron, a p-electron harder to remove than a d-electron and so on.

<u>example</u>

The ionisation potential corresponding to the removal of $2p^1$ electron of boron ($B \longrightarrow 2s^2 p^1$) is 8.3 eV while those corresponding to the removal of two 2s electrons are 25.1 and 37.9 eV

(iii) Completely - filled and half - filled orbitals

According to Hund's rule, atoms having half - filled or completely filled orbitals are comparatively more stable and hence more energy is needed to remove an electron from such atoms.

Example

Be and N in the second period.

Be ($Be \longrightarrow 2s^2$)

N ($N \longrightarrow 2s^2 p^3$)

(iv) The shielding effect (or) screening effect of the inner electrons on the valance electrons.

It has been observed that a valance - electron in a multi - electron atom is attracted by the nucleus and repelled by the electrons of inner - shells.

The combined effect of this attractive and repulsive force acting on the valance electron is that the valance - electron experiences less attraction from the nucleus. This is known as screening effect. Thus larger the number of electrons in the inner - shell, lesser is the attractive force holding the valance electron to the nucleus and consequently the lower will be the value of ionisation potential.

(v) Atomic radius and variation in a group.

The ionisation potential decreases with the increase in atomic radius. This is because of the fact that in case of larger atoms, the attraction between the nucleus and outer - most electron is less.

Moving from

ELECTRONEGATIVITY

The Measure of the capacity or tendency of an atom to attract the shared pair of electrons of the covalent bond towards itself is called electronegativity of that atom.

Electronegativity is a relative value that indicates the tendency of an atom to attract shared electrons more than the other atom bonded to it. Therefore it does not have any unit. Pauling was the first scientist to put forward the concept of electronegativity.

The numerical value of electronegativity of an atom depends on its ionisation potential and electron affinity values.

Factors affecting Electronegativity

Atomic size

Electronegativity of a bonded atom decreases with increase in its size.

$$\boxed{\text{Electronegativity} \propto \frac{1}{\text{Atomic size}}}$$

Periodic variations

(1) In a period

Electronegativity increases on moving in a period (left to right) in the periodic table. This is due to increase in nuclear charge.

(2) In a group

In moving down through a group of the Periodic table since the nuclear charge again increases.

ELECTRON AFFINITY [EA]

The energy released on adding up one mole of electron to one mole of neutral atom (A) in its gaseous state to form an anion (A^-) is called electron affinity of that atom. In general electron affinity is associated with an exothermic process.

$$A_{(g)} + e^- \longrightarrow A^-_{(g)}, \quad \Delta H_{eg} = -$$

$$A^-_{(g)} + e^- \longrightarrow A^{2-}_{(g)}$$

Factors affecting Electron Affinity

* **Effective nuclear charge:**

When effective nuclear charge is more, then the atomic size less. Hence EA increases.

$$\text{Electron affinity} \propto \text{Effective nuclear charge}$$

* **Atomic size or Atomic Radius:**

When the size of an atom increases the electron entering the outermost orbit is more weakly attracted by the nucleus and the value of electron affinity is lower

$$\boxed{\text{Electron affinity} \propto \frac{1}{\text{Atomic size}}}$$

* **shielding effect:**

Shielding effect is directly proportional to atomic size and atomic size is inversely proportional to electron affinity.

$$\boxed{\text{Electron affinity} \propto \frac{1}{\text{shielding effect}}}$$