

Periodicity - Periodic Properties, Atomic Radius

1. When elements are arranged in increasing order of atomic number, elements with similar properties reoccur (due to similar outer electronic configuration) at regular intervals in the periodic table. This repetition of properties is called periodicity and such properties are called periodic properties
2. Elements coming at intervals of 2, 8, 8, 18, 18, 32 will have similar properties and thus grouped in one particular group.

Ex: Elements with atomic number 1, 3, 11, 19, 37, 55 & 87(IA). Elements with atomic number 4, 12, 20, 38, 56 & 88(IIA) will have similar properties.

2. Properties like valency, effective atomic number, screening effect, atomic radius, ionic radius, ionisation potential, electron affinity, electronegativity, metallic nature, oxidation and reduction ability, acidic or basic nature of the oxides, etc. are periodic properties.
3. Properties like specific heat, refractive index, colour etc., are not periodic properties as they are not related to the electronic configuration of elements.

Atomic Radius

4. The distance between the centre of the nucleus and the outer most orbit of an atom is called atomic radius. It can be determined by using X-ray diffraction techniques. It is expressed in angstrom, nanometers, picometer units.

$$1 \text{ \AA} = 10^{-1} \text{ nm}; 1 \text{ \AA} = 10^2 \text{ pm}$$

5. Atomic radius depends on
 - a) Nature of bonding
 - b) Bond order i.e. Number of bonds (multiplicity of bonding)
 - c) Oxidation states

- d) Co-ordination number of atom
- e) Bond character etc.
6. based on the nature of bonding atomic radii are of three types , they are
- a) Crystal radius
 - b) Van der Waals radius
 - c) Covalent radius
7. **Crystal Radius (Atomic Radius)** - It is half of the internuclear distance between the adjacent atoms of a solid metallic crystal.
- Ex:** Distance between two nuclei of adjacent sodium atoms is 3.72\AA ,
crystal radius of Na = $3.72 / 2 = 1.86\text{\AA}$.
8. **Covalent Radius:** It is half of the internuclear distance of the two atoms held together by a covalent bond. it is used for non-metals.
- Ex:** a) in H_2 molecule H - H bond distance (Internuclear distance) is 0.68\AA .
Covalent Radius of H = 0.34\AA .
9. **Van der Waals radius:** It is half of the internuclear distance between two atoms of adjacent molecules which are held by Van der Waals forces in solid state. It is used for molecular substances and inert gases in the solid state. Van der Waals radius is 40% greater than covalent radius.
- Ex:** The distance between two adjacent atoms of H_2 molecules in solid state is 2.4\AA Vander Waals radius of H is 1.2\AA .
10. In general Vander Waals radius > crystal radius > Covalent Radius
11. As the bond order i.e number of bonds between two atoms increases the covalent radius decreases.

Ex: The covalent radius of carbon decreases with increase in the number of

bonds between carbon atoms.

$\text{C}-\text{C}$	$\text{C}=\text{C}$	$\text{C}\equiv\text{C}$
(1.54Å)	(1.34Å)	(1.20Å)

13. In a period from left to right, atomic radius decreases upto halogens and increases to inert gases as the effective nuclear charge increases.

In each period alkali metal is the largest and halogen atom is the smallest in size.

14. In each period the radius of an inert gas is nearly twice than halogen atom as only Van der waal radius is applicable for atoms of inert gases

15. In *a group* from top to bottom, the atomic radius increases gradually due to the increase in the number of orbits and screening effect.

Atomic radius is least for hydrogen and is highest for Caesium among the available elements.

Ex: Variation of atomic radius in IA group is $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$

16. in transition elements the decrease in size in a period across a particular transition series is less than in case of representative elements due to screening effect of (n-1)d-electrons.

Thus the atomic radius decreases slightly as we move from left to right in a transition series.

* Covalent radii of $\text{Zn} > \text{Cu}$ due to repulsions among 3d electrons.

17. In Lanthanides (La, Ce, Pr, Nd, Pm, Sm, Eu, Gd, Tb, Dy, Ho, Er, Tm, Yb & Lu),the atomic and ionic radii decreases steadily due to poor shielding effecting of f-orbitals because of their diffused shape. This is known as “Lanthanide Contraction”

Consequences of Lanthanide contraction

a) Atomic sizes of 4d and 5d transition elements of IVB(Zr and Hf), VB(Nb and Ta), VIB(Mo and W) are almost equal.

b) The crystal structure and other properties of lanthanides are very similar. Hence Separation of lanthanides is difficult from their mixture.

c) Heavy metals of p- block in IIIA, IVA, VA group i.e Tl, Pb, and Bi respectively exhibit inert pair effect.

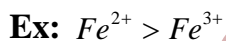
18. The distance between the nucleus and the electron in the outer most shell of an ion is called Ionic Radius.

19. When a neutral atom loses one (or) more electrons, a positive ion called cation is formed.



The radius of cation is less than that of neutral atom as the cation has higher effective nuclear charge. **Ex:** $Na > Na^{+}$

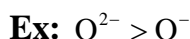
Among the cations as the positive charge increases, the ionic radius decreases.



20. When a neutral atom gains one (or) more electrons a negative ion called anion is formed. The radius of anion is more than that of its atom, due to decrease in effective nuclear charge.



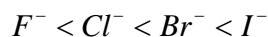
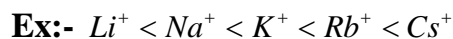
Among the anions as the negative charge increases the ionic radius increases.



21. The decreasing order of the radii is

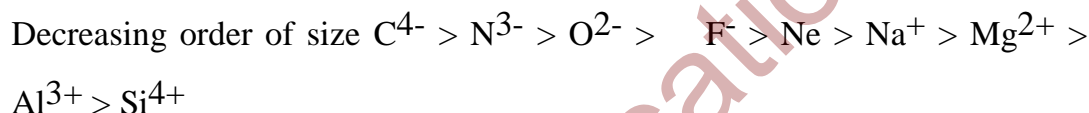


In a group, the radius ions (cations or anions) increase on moving from top to bottom due to increase in number of shells.



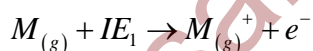
22. H^- & I^- are the smallest and largest anions respectively. Smallest atom is He & largest atom is Fr.

23. The species (atoms or ions) having the same numbers of electrons are known as iso - electronic species. In iso electronic species, the size increases with increase of negative charge and decreases with increase of positive charge.



Ionization Energy (Ionization Enthalpy)

1. The minimum amount of energy required to remove the most loosely bound electron from an isolated neutral gaseous atom is called First ionization potential (IE_1).



It is an endothermic process. It is measured in eV/atom or kJ/mole or K.cal/mole.

$$1 \text{ eV / atom} = 23.06 \text{ K.Cal/mole} = 96.45 \text{ KJ/mole} =$$

2. Energy required to remove an electron from unipositive ion to convert it into dipositive ion is IE_2 . $M_{(g)}^+ + IE_2 \rightarrow M_{(g)}^{2+} + e^-$

3. For a given element $IE_2 > IE_1$ as the effective nuclear charge in uni positive ion is more than in a neutral atom

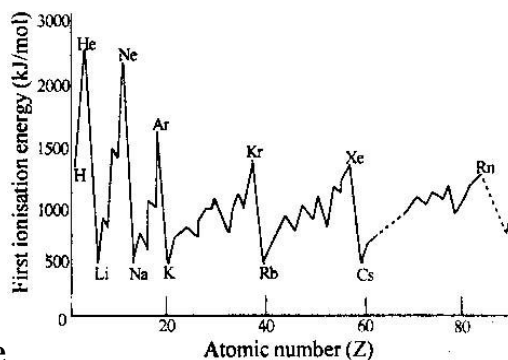
4. For a given element the number of IE values possible is equal to its Atomic number (Z)
5. Ionization energy can be determined by discharge tube experiment.
6. Ionization potential depends on
 - a) **Size of atom:** "IP" decreases with increase in the atomic size due to decrease in attractive force of nucleus on outer most orbit electrons.
 - b) **Nuclear charge:** IP increases with increase in the effective nuclear charge.
 - c) **Shielding capacity:** IP decreases as the number of electrons in the inner shells are more due to shielding capacity of the inner electrons on the outer electron increases.
 - d) In a given orbit the order of screening power of orbitals is $s > p > d > f$
 - e) IP increases as the positive charge on cation increases.
 - f) IP increases as the valency electrons are more penetrated into inner shells, Penetration power of different orbitals is in the order of $s > p > d > f$
 - g) Atoms with half filled or completely filled orbitals have relatively higher IE value.

Eg: IE_1 of N $>$ IE_1 of O & IE_1 of Be $>$ IE_1 of 'B' in II Period

 IE_1 of P $>$ IE_1 of S & IE_1 of Mg $>$ IE_1 of 'Al' in III Period

Eg: IE_2 of O $>$ IE_2 of F & IE_2 of B $>$ IE_2 of C' in II Period

 IE_2 of S $>$ IE_2 of Cl & IE_2 of Al $>$ IE_2 of Si in III Period
 - h) Atoms of inert gases have highest IP values due to the presence of completely filled orbitals.
 - i) In the graph showing relation between IP and atomic number, the inert gases appear at the peaks and alkali metals appear at the bottom



Ionization energy curve

- 7) In any period an Alkali metal atom has lowest IP and inert gas element has highest IP.
- 8) In periods from left to right side IP increases, due to decrease in atomic size and increase in effective nuclear charge.
- 9) In groups from top to bottom, IP decreases due to the increase in the atomic size and increase in the screening effect. Variation of IE_1 in I A group
 $Li > Na > K > Rb > Cs$
10. Variation of IE among II period elements.
 $IE_1 Li < Be > B < C < N > O < F < Ne$ &
 $IE_2 Li > Be < B > C < N < O > F < Ne$
- 11) Variation of IE among III period elements.
 $IE_1 - Na < Mg > Al < Si < P > S < Cl < Ar$ &
 $IE_2 - Na > Mg < Al > Si < P < S > Cl < Ar$
- 12) Cs has Lowest IP and He has highest IP elements in periodic table.
- 13) Knowledge of successive IE can be used to find the number of valence electrons
 - a) For alkali metals the IE_2 shows sudden jump.
 - b) For alkaline earth metals, the IE_3 shows sudden jump.
 - c) For IIIA elements, the IE_4 shows sudden jump
- 14) I.E of coinage metals is $Cu > Ag < Au$.

15) **Effective nuclear charge (Z_{eff})**

Due to screening effect the valency electron experiences less attraction towards nucleus. This decrease the nuclear charge (Z) actually present on the nucleus. The reduced nuclear charge is called effective nuclear charge and is represented by Z_{eff} .

16) Z_{eff} is related to actual nuclear charge (Z) by the following formula $Z_{\text{eff}} = Z - \sigma$: where σ is screening constant.

The magnitude of effective nuclear charge increases in a period from left to right and remains almost the same for elements in a subgroup.

17) The effective nuclear charge (Z_{eff}) can be calculated

$$1) Z_{\text{eff}} = \sqrt{\frac{(E)(n^2)}{1312 \text{ KJ/mol}}} \text{ Where,}$$

E = The energy necessary to remove an electron from an atom

n = The principal quantum number of the electron (Or)

$$2) Z_{\text{eff}} = Z - S \text{ (or) } Z_{\text{eff}} = Z - \sigma$$

Where, Z = Nuclear charge

S = Shielding or Screening constant

18. Rules to calculate the shielding or screening constant for an np or ns- orbital

i) Write out the electronic configuration of the element in the following order and groupings: (1s) (2s, 2p) (3s, 3p) (3d) (4s, 4p) (4d) (4f) (5s, 4p), etc.

ii) Electrons in any group to the right of the (ns, np) group contribute nothing to the shielding constant

iii) All of the other electrons in the (ns, np) group, shield the valence electron to an extent of 0.35 each

iv) All electrons in the (n-1) shell shield to an extent of 0.85 each

v) All electrons (n - 2) or lower shell shield completely, if their contribution is 1.00 each.

Examples: (i) Let us consider the valence electron in the Oxygen (${}_8\text{O}$)

atom = $1s^2 2s^2 2p^4$. Grouping of the orbitals gives $(1s^2)(2s^2 2p^4)$

$$\sigma = (2 \times 0.85) + (5 \times 0.35) = 3.45$$

$$Z_{\text{eff}} = Z - \sigma = 8.0 - 3.45 = 4.55$$

19. IN an isoelectronic series as +ve charge increases, IE increases and -ve charge increases, IE decreases
20. In case of IIA group elements, Cu and Cr atoms their is nearly twice to
- 21) In case of IIIB group elements of $\text{Sc} > \text{Y} > \text{La}$.
- 22) In case of IVB group elements of $\text{Ti} > \text{Zr} < \text{Hf}$, of $\text{Zr} < \text{Hf}$ due to lanthanide contraction.