CHAPTER NAME : CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

MODULE NO : 02/03(PERIODICITY IN PROPERTIES)
SUBJECT : CHEMISTRY
CLASS : XI

LEARNING OBJECTIVES

Periodic trends in properties of elements- trends in Physical Properties

1) Atomic size

2) Ionization energy/Ionization potential/Ionization enthalpy)

Periodicity-periodic properties

When elements are arranged in increasing order of atomic number, elements with similar properties reoccur (due to similar outer electronic configuration) at regular intervals of atomic numbers in the periodic table. This repetition of properties is called periodicity and such properties are called periodic properties.

Ex. Elements coming at intervals of 1,3,11,19,37,55 & 87 will have similar properties and thus grouped in one particular group(i.e Group 1)

Periodic trends in physical properties

1) Atomic size

(a) Atomic radii
    (i) covalent radius,
    (ii)metallic radius(or crystal radius))
(b)Inert gas radius(Van der waals' radius)
(c) Ionic radii

2) Ionization energy /Ionization potential/Ionization enthalpy

3) Electron gain enthalpy

4) Electronegativity

I) Atomic size

(a) Atomic radii

((i)covalent radius, (ii)metallic radius(or crystal radius))

(b) Inert gas radius(Van der Waals' radius)

(c) Ionic radii

(a) Atomic radii:

Atomic radius is the distance between the centre of the nucleus and the electron cloud of outer most energy level is called atomic radius

To avoid practical difficulties, Atomic radius cannot be determined directly, but measured from the internuclear distance of combined atoms, using X-ray diffraction techniques

Atomic radius refer to both covalent or metallic radius depending on whether the element is a non-metal or a metal.

(i)Covalent radius: Half of the internuclear distance of the two atoms held together by a covalent bond is called covalent radiu This radius is generally used in reference to non-metals held b covalent bond

Ex. In the chlorine molecule (Cl_2) is 198 pm and half of this distance (99 pm), is taken as the atomic radius of chlorine.

(ii)Metallic radius(or crystal radius):

Half of the internuclear distance between the adjacent atoms o solid metallic crystal is called crystal radius or metallic radius

Ex. The distance between two adjacent copper atoms in solid copper is 256 pm; hence the metallic radius of copper is assigned a value of 128 pm.

- **(b)Inert gas radius(Van der Waals' radius)**

Half of the internuclear distance between two atoms of differen molecules which are very close to each other in solid state due to Van der waals' forces is called Van der waals' radius.

- This radius is used for molecular substances and inert gases in the solid state only.
- Ex. The distance between two adjacent chlorine atoms of different Cl_2 molecules is 360pm, Van der Waals' radius of Cl is 180pm. ****Order of radii: (Decreasing order)

Inert gas radius(Van der waals' radius), Metallic radius(or cryst radius),Covalent radius

(c) Ionic radii:

It is defined as the distance between the nucleus and the electron in the outer most shell of an ion.

The ionic radius of cation is less than that of neutral atom. It is because the cation has higher effective nuclear charge.

Ex. Size of Na is bigger that Na^+

Among the cations as the positive charge increases, th ionic radius decreses.

Ex. Ionic radius of Na^+ is higher than Mg^{2+}

The ionic radius of anion is more than that of its atom, due to decrease in effective nuclear charge.

Ex. Ionic radius of Cl^- is greater than Cl

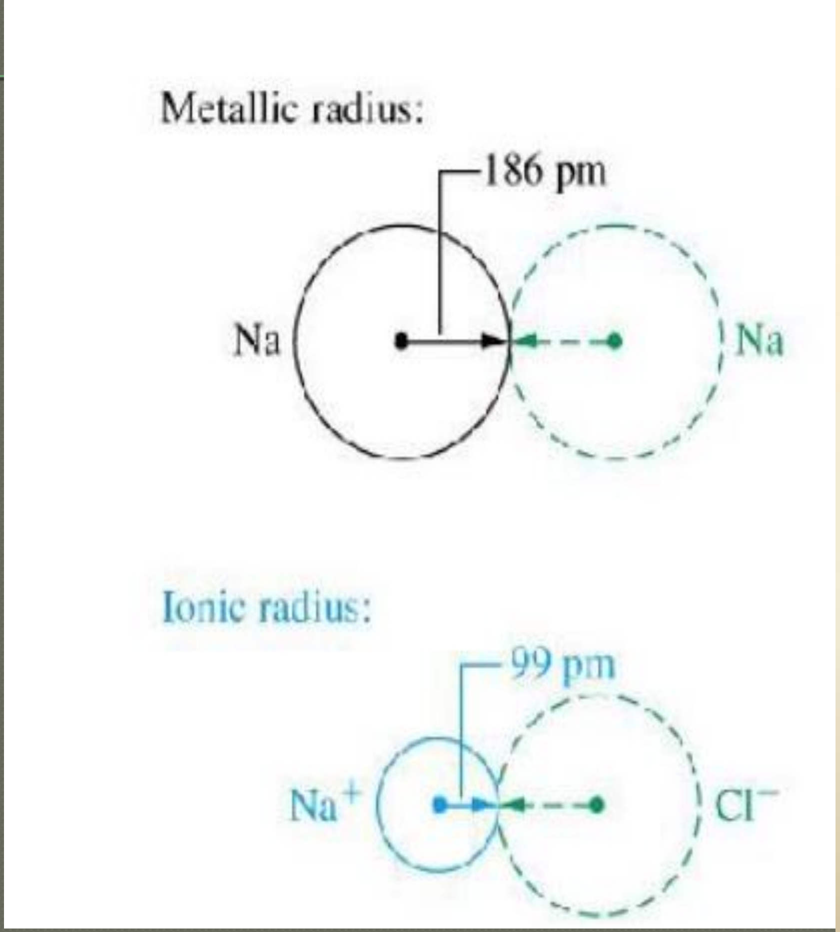
Among the anions as the negative charge increases th ionic radius increases

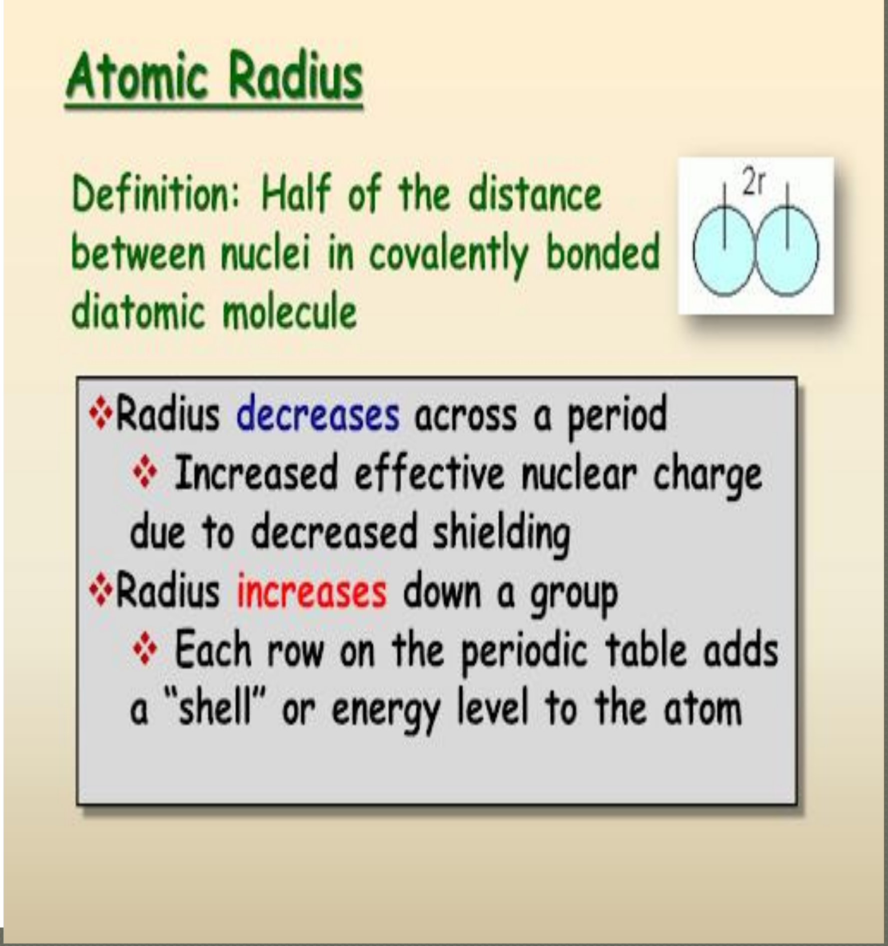
Ex. O^{2-} is begger than F^-

The decreasing order of the radii is

Anion, Atom, Cation

Ex. I^-,I,I^+





Atomic Radius

Definition: Half of the distance between nuclei in covalently bonded diatomic molecule

- Radius decreases across a period
  - Increased effective nuclear charge due to decreased shielding
- Radius increases down a group
  - Each row on the periodic table adds a “shell” or energy level to the atom

As effective nuclear charge increases in period, atomic size decreases

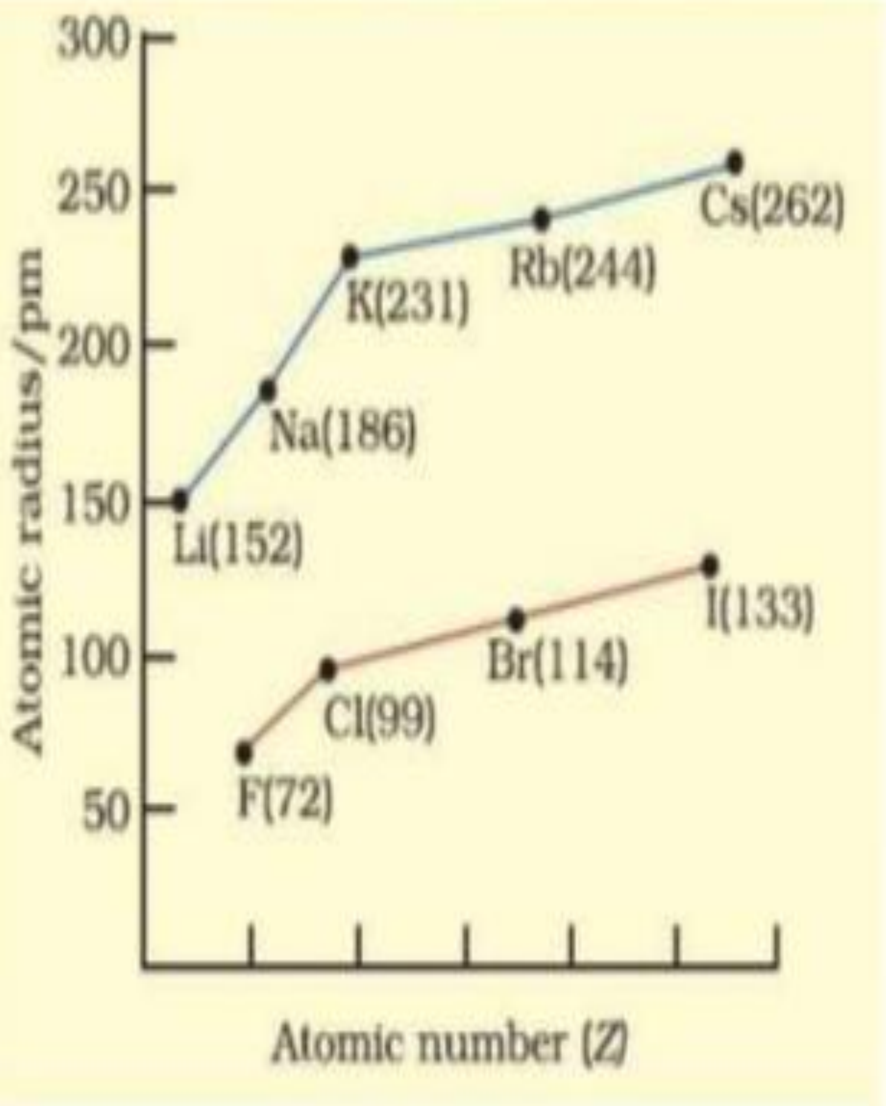


Fig. 3.4 (b) *Variation of atomic radius with atomic number for alkali metals and halogens*

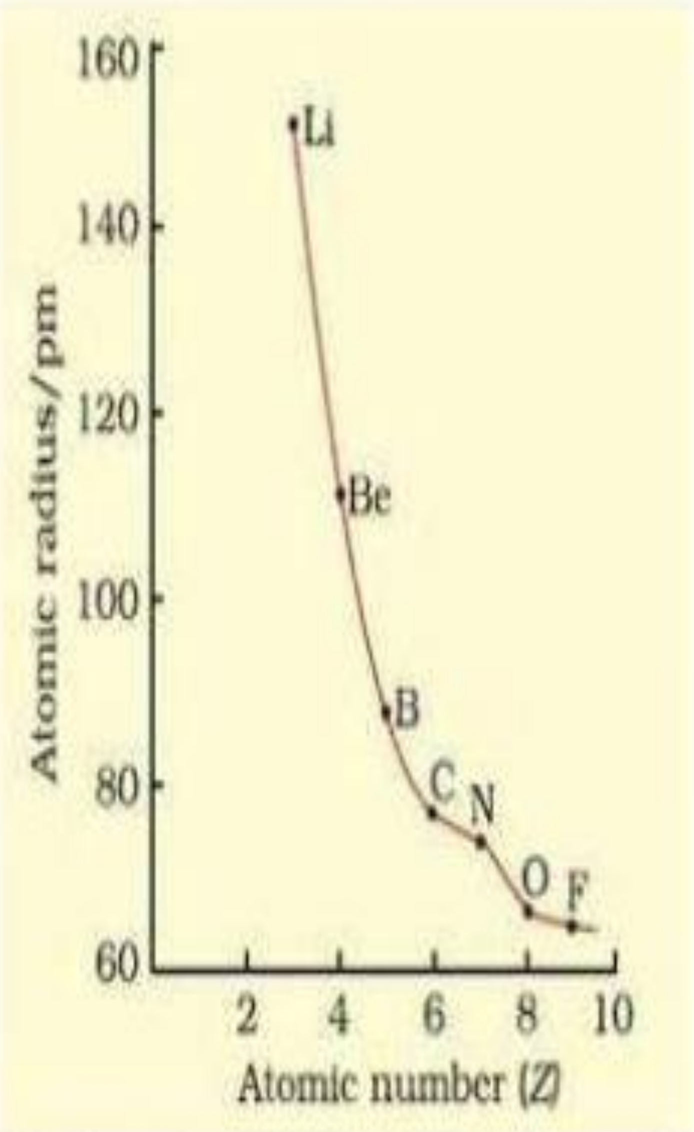
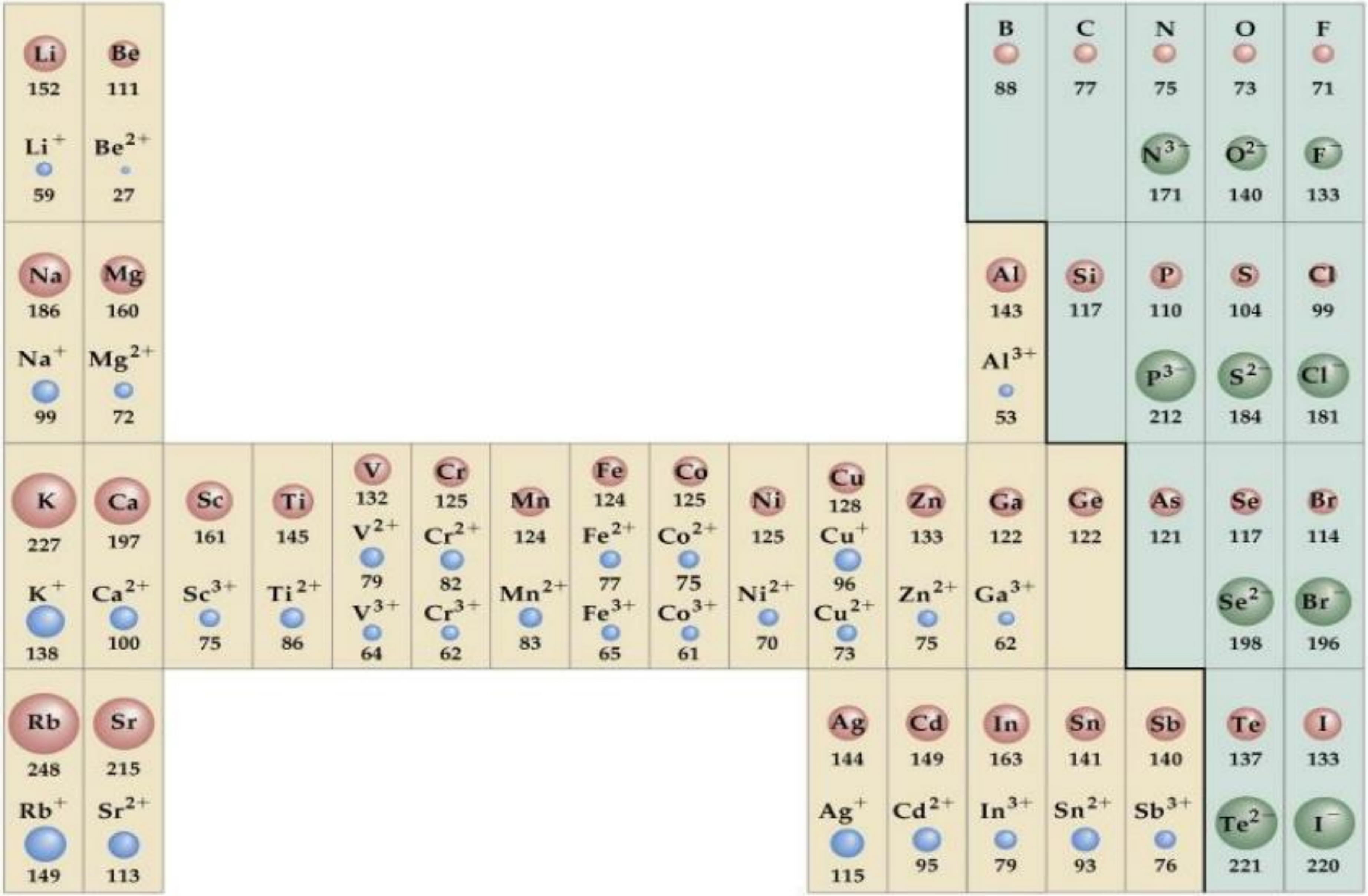


Fig. 3.4 (a) *Variation of atomic radius with atomic number across the second period*

Atomic and Ionic Radii



Li 152 Li^+ 59	Be 111 Be^{2+} 27											B 88	C 77	N 75 N^{3-} 171	O 73 O^{2-} 140	F 71 F^- 133
Na 186 Na^+ 99	Mg 160 Mg^{2+} 72											Al 143 Al^{3+} 53	Si 117	P 110 P^{3-} 212	S 104 S^{2-} 184	Cl 99 Cl^- 181
K 227 K^+ 138	Ca 197 Ca^{2+} 100	Sc 161 Sc^{3+} 75	Ti 145 Ti^{2+} 86	V 132 V^{2+} 79 V^{3+} 64	Cr 125 Cr^{2+} 82 Cr^{3+} 62	Mn 124 Mn^{2+} 83	Fe 124 Fe^{2+} 77 Fe^{3+} 65	Co 125 Co^{2+} 75 Co^{3+} 61	Ni 125 Ni^{2+} 70	Cu 128 Cu^+ 96 Cu^{2+} 73	Zn 133 Zn^{2+} 75	Ga 122 Ga^{3+} 62	Ge 122	As 121	Se 117 Se^{2-} 198	Br 114 Br^- 196
Rb 248 Rb^+ 149	Sr 215 Sr^{2+} 113									Ag 144 Ag^+ 115	Cd 149 Cd^{2+} 95	In 163 In^{3+} 79	Sn 141 Sn^{2+} 93	Sb 140 Sb^{3+} 76	Te 137 Te^{2-} 221	I 133 I^- 220

ISO-ELECTRONIC SPECIES

- The species(atoms or molecules ions) having the same number 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

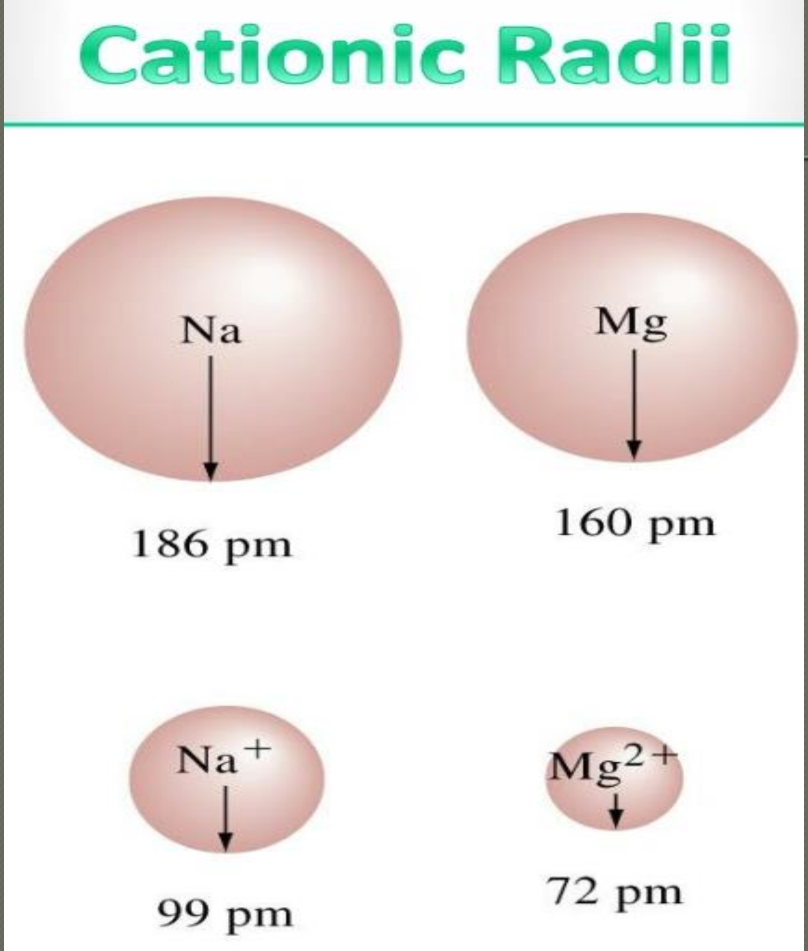
Q. Arrange the following iso-electronic species in the decreasing order of size

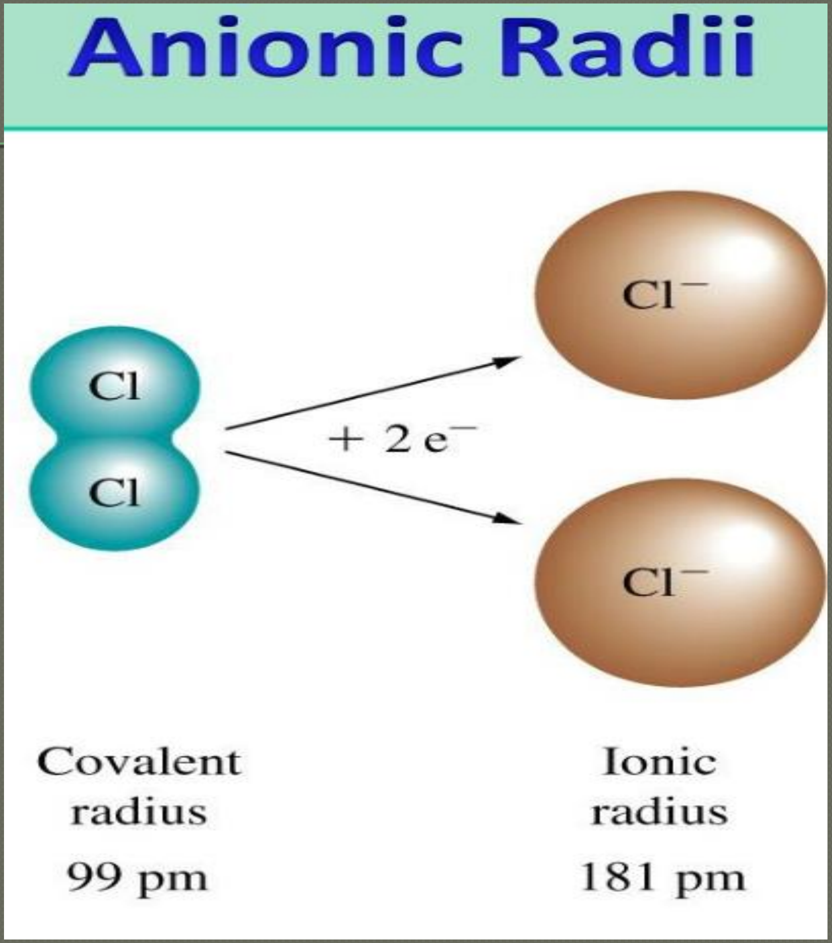
C^{4-}, O^{2-},Ne,F^{-},N^{3-},Mg^{2+},Al^{3+},Si^{4+},Na^{+}

(iso electronic species with 10 electrons)

Ans: C^{4-}, N^{3-}, O^{2-}, F^{-}, Ne, Na^{+}, Mg^{2+}, Al^{3+}, Si^{4+}

- (Hint $$\text{size} \ \alpha \ \frac{1}{\text{effective nuclear charge}}$$)


Cationic Radii


Anionic Radii

2) Ionization energy /Ionization potential/Ionization enthalpy:

- The minimum amount of energy required to remove the most loosely bound electron(i.e outer- most shell e^-)from an isolated neutral gaseous atom is known as Ionization energy/Ionization potential/Ionization enthalpy(IE/$\Delta_i H$)
  (Isolated -> without any bonding with other atom)
- $M(g) \; + \; IE_1 \rightarrow M^+ \,(g) \;\; + 1e^-$
- It is an endothermic process
- It is measured in eV/atom or KJ/mol or Kcal/mol
  $1eV/atom=23.06K.cal/mol=96.45KJ/mol=1.6.02x10^{-19}J/atom$
- II Ionization Enthalpy is the energy required to remove an electron from unipositive ion to convert it into dipositive ion(IE_2)
  $M^+(g) \; + \; IE_2 \rightarrow M^{++} \,(g) \;\; + 1e^-$
  Here IE_2 greater than IE_1

Q. Why IE_2 is greater than IE_1

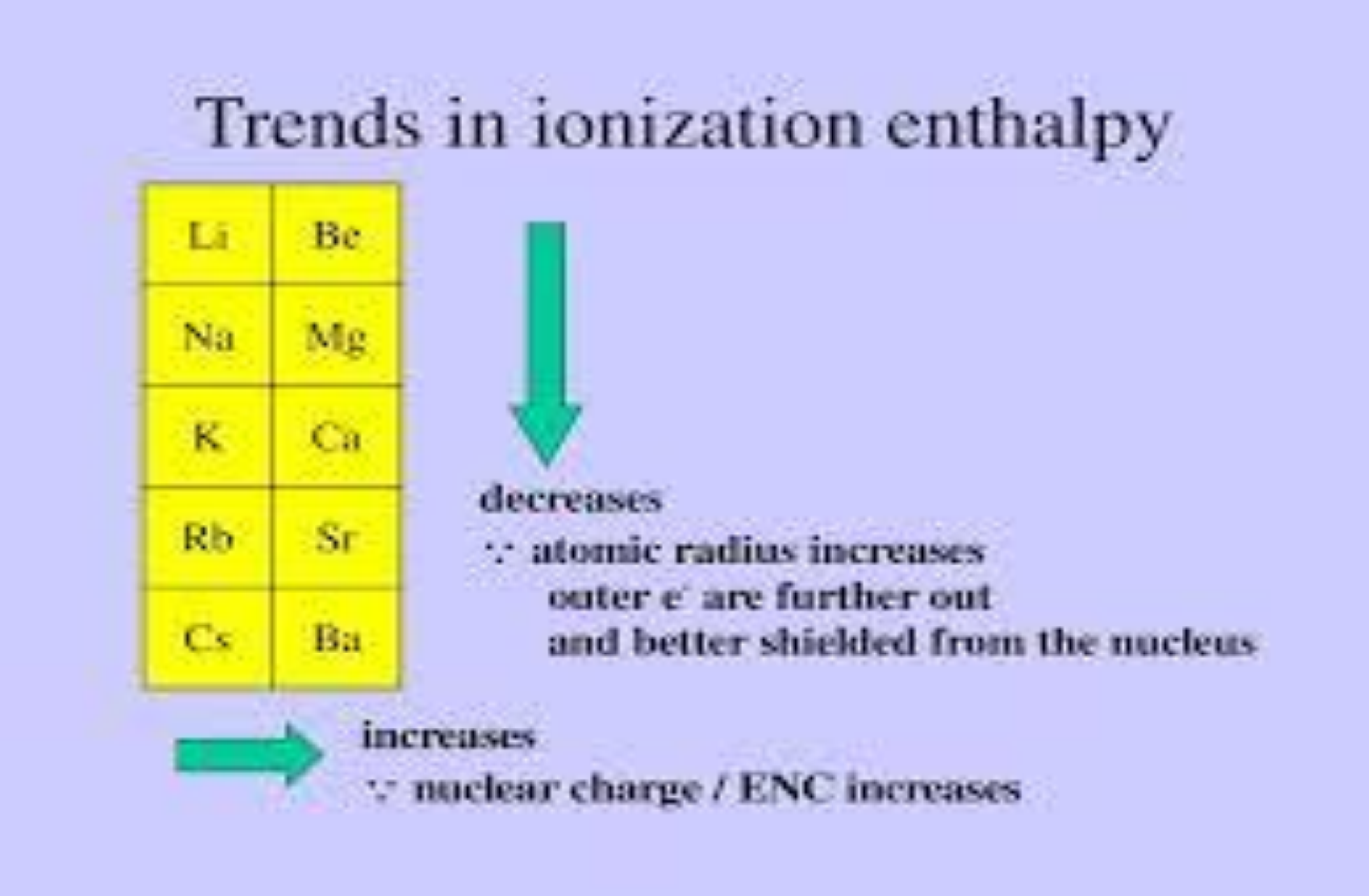
The second ionization enthalpy will be higher than the first ionization enthalpy because it is more difficult to remove an electron from a positively charged ion than from a neutral atom because a cation has greater ENC than a neutral atom.

In the same way the third ionization enthalpy will be higher than the second and so on.

Q. Why IE_2 shows sudden jump for alkali metals whereas IE_3 for alkaline earth metals?

- Ans. As they get stable noble gas electronic configurations (i.e Ne)
- Alkali metals
  ex Na-$1s^1,2s^2 2p^6 3s^1$
- Alkaline earth metals.
  Ex Mg-$1s^2 2s^2 2p^6 3s^2$

Periodic Trend of Ionisation Enthalpy ($\Delta_i H$)



Trends in ionization enthalpy

Li	Be
Na	Mg
K	Ca
Rb	Sr
Cs	Ba

decreases

∵ atomic radius increases
outer e^- are further out
and better shielded from the nucleus

increases

∵ nuclear charge / ENC increases

Factors affecting Ionisation Enthalpy ($\Delta_i H$)

(a)Atomic size

(b)Effective nuclear charge (Z_{eff}) :

(c) Screening effect :

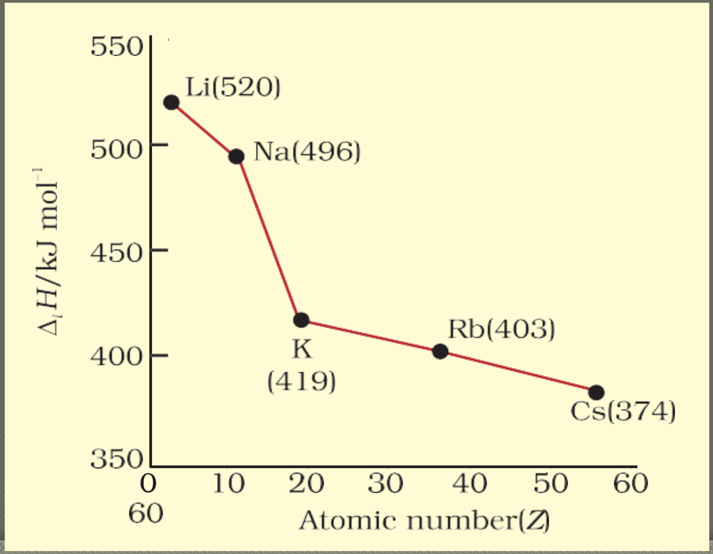
(d) Penetration power of sub shells:

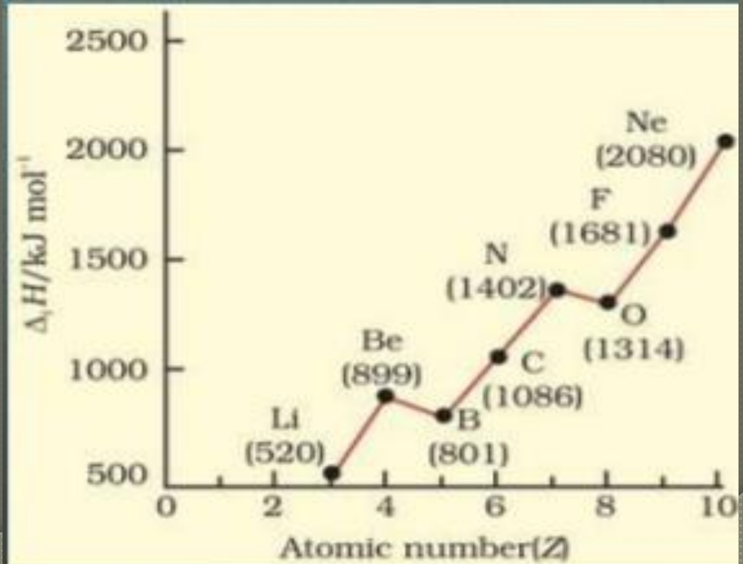
(e)Stability of half filled and fully filled orbitals :

(a)Atomic size

(a)Atomic size: Larger the atomic size, smaller in the Ionisation potential. It is due to that the size of atom increases the outermost electrons e^- farther away from the nucleus and nucleus loses the attraction on that electrons and hence that can be easily remove

$$\text{Ionisation potential} \;\alpha\; \frac{1}{\text{Atomic size}}$$





b)Effective nuclear charge (Z_{eff}) : Ionisation potential increases with the increase in nuclear charge between outermost electrons and nucleu

Ionisation potential α effective nuclear charge

(c) Screening effect :Higher is the screening effect on the outer most electrons causes less attraction from the nucleus and can be easily removed , which is leading to the lower value of Ionisation potential .

$$\text{Ionisation potential} \quad \alpha \quad \frac{1}{\text{Screening effect}}$$

(d) Penetration power of sub shells:

(i) Order of attraction of subshells towards nucleus (Penetration power) is $s>p>d>f$

(ii) As s- subshell is more closer to nucleus so more energy will be required to remove e^- in comparision to p,d,&f.

Ex.

	Be	B
	$1s^{2,}\ 2s^2$	$1s^2\ 2s^2\ 2p1$
Ionisation potential(IE_1)	Be >	B

After loosing one e^- B attains electronic configuration of Be , so IInd ionisation potential of B is more than Be

IInd Ionisation potential of B > Be

Ionisation potential α Penetration power of sub shells

(e)Stability of half filled and fully filled orbitals :

- Half filled p^3 ,d^5 ,f^7 or fully filled s^2 , p^6, d^{10} , f^{14} are most stable than others so it requires more energy .
  Ex : N O
  $1s^2$ $2s^2$ $2p^3$ $1s^2$ $2s^2$ $2p^4$
  1st Ionisation potential order is O < N because of half filled p- orbitals in N , its ionisation energy (due to stability) is higher than O.

- Ionisation potential α atoms with half filled and fully filled orbitals

Ionization potential:
Group 1 Vs Group 0(18)