

# PERIODIC TABLE

**Ex.**

Atomic No.	IUPAC Name	Symbol	Elemental Name	Symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf / <i>Kn</i>
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnilium	Uun	Darmstadtium	Ds

(viii) **Identification of group, period and block:**

(A) **When atomic number is given :**

**Step I:**  $71 \geq Z \geq 58 \Rightarrow$  Lanthanides ( $6^{\text{th}}$  Period)  $\square$  f - block  
 $103 \geq Z \geq 90 \Rightarrow$  Actinoids ( $7^{\text{th}}$  Period)  $\square$

**Group number = IIIB (largest group of periodic table)**

~~mp~~

0 = nil

1 = un

2 = Bi

3 = Tri

4 = quad

5 = Pent

6 = Hexa

7 = Sept

8 = Oct

9 = enn

101

Unnilunium

Unu

112

Ununbium

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**Step II:**  $Z = \underline{104}$  to  $\underline{118}$  (Period number = 7 )

Group number = last two digits in atomic number of element

**Example:**  $Z = \underline{\underline{104}}$

| 0 | —————

Group no. = 4

**Step III:** Group number =  $\underline{18} + \text{given atomic number} - \text{atomic number of next noble gas}$  If the value of this formula is negative then use 32 instead of 18 in formula. When electronic configuration is given

Period number (n) = number of outermost shell/highest shell number.

## Block identification :

- If np electron present then p - block ( $ns^2np^{1-6}$ ) group number =  $12 + np$  electrons
- If np electron absent then s/f/d block

Block — last  $e^-$  in  
which shell  
that is Block

$$\underline{X} = 1s^2 \ 2s^2 \ \underline{2p^1}$$

P-Block

Period — last shell  $\Rightarrow$  Period

2nd Period

$$\text{group} = \text{no of ns } e^- + \text{no of np } e^- \\ 2 + 1 = \underline{3}$$

# PERIODIC TABLE

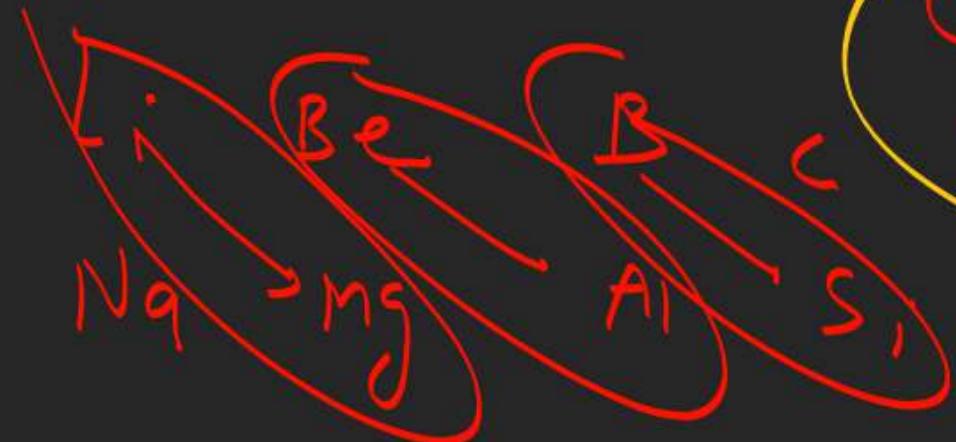
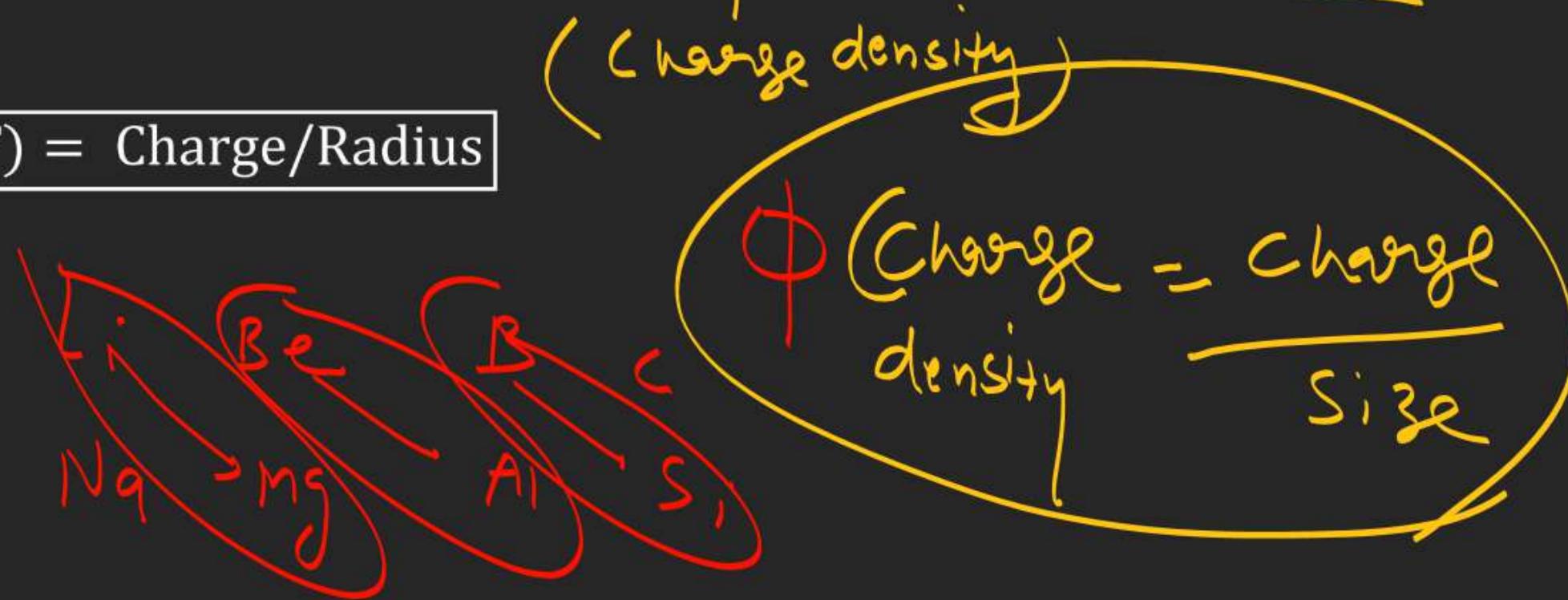
- **If  $(n - 2)f^0(n - 1)d^0 ns^{1-2} = s \text{ block}$**   
**group number = ns electrons**  
**If  $(n - 2)f^{1-14}(n - 1)d^{0-1} ns^2 = f \text{ block}$**   
**group number = IIIB**
- **If any other configuration or  $(n - 1)d^{1-10} ns^{0-2} = d - \text{block group number}$**   
 **$= (n - 1)d \text{ electron} + ns \text{ electron}$**

# PERIODIC TABLE

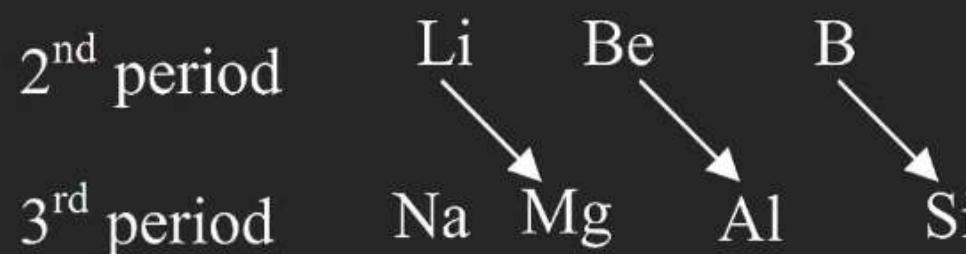
- **Some Important Points:**

- (a) **2<sup>nd</sup> period elements ( Li, Be, B) shows diagonal relationship with 3rd period elements (Mg, Al, Si). Because of same ionic potential value they show similarity in properties.**

$$\text{(Ionic potential) } (F) = \text{Charge} / \text{Radius}$$



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(b) **(3<sup>rd</sup>) period elements (Na,Mg,Al,Si,P,S,Cl)** except **inert gases** are **called typical elements** because they represent the properties of other element of their **respective group.**

(c) **Representative or normal elements.**

(i) Outermost shell of these elements is **incomplete**. The number of electrons in the outermost shell is less than eight.

(ii) **s-and p-block elements** except **inert gases** are called **normal or representative elements.**

(d) **Transuranic Elements :**



Elements having atomic number more than 92 are known as transuranic element. All transuranic elements are radioactive & artificial.

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			f-block		d-block	
Th	Pa	U	Np 93	—	Lr 103	Unq 104
90	91	92				Uub 112

**First man made element is - Tc**

**First man made lanthanoid is – Pm**

**All actinoids are radioactive but all lanthanoids are not artificial / man made (except Pm)**

(e) The group containing most **electropositive elements** - GROUP 1A.

(f) The group containing most **electronegative elements** - GROUP VIIA

(g) The group containing maximum number of gaseous elements - **GROUP ZERO**

(h) The group in which elements have generally **ZERO valency** - GROUP ZERO (18th)

(i) **Bridge elements** have similar electronic configuration and group nomenclature e.g.

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	Na	Mg	Al	Si	P	S	Cl	
K	Cu ; Ca	Zn ; Ga	Sc ; Ge	Ti ; As	V ; Se	Cr ; Br	Mn	

(j) In the periodic table

**Number of Gaseous elements - 11 ( H, N, O, F, Cl + Noble gases)** **Number of Liquid elements upto 30°C – 6 (Cs, Fr, Ga, Hg, Br, Uub)** **Bromine is the only non-metal which exists in liquid form.** **Number of Solid elements - 95 (if discovered elements are 112)**

(k) 0/18<sup>th</sup> group have all the elements in gaseous form.

(l) 2nd period contains maximum number of gaseous elements - N, O, F, Ne

(m) IIIB/3rd group is called longest group having 32 elements including 14 Lanthanides and

# PERIODIC TABLE

Actinides

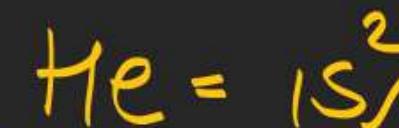
Sc

Y

La.....Lanthanides (14)

Ac.....Actinides (14)

H



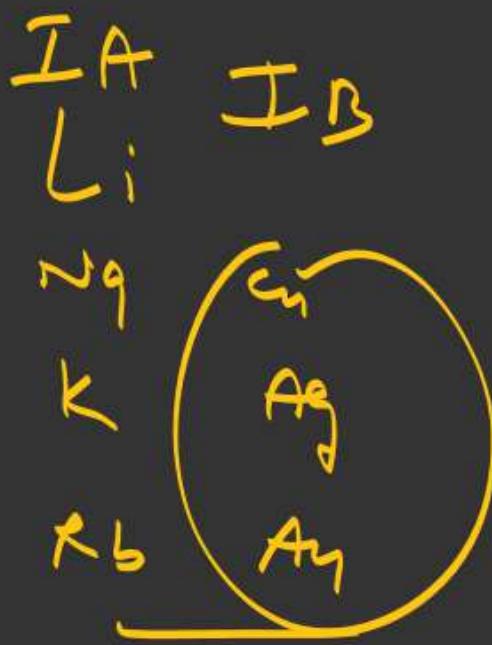
## Defects:

- (a) Position of hydrogen is not settled.
- (b) Position of helium cannot be justified. It is the only element which belongs to s-block but placed in p-block.
- (c) Lanthanides & Actinides have not been given space in the main body of periodic table.

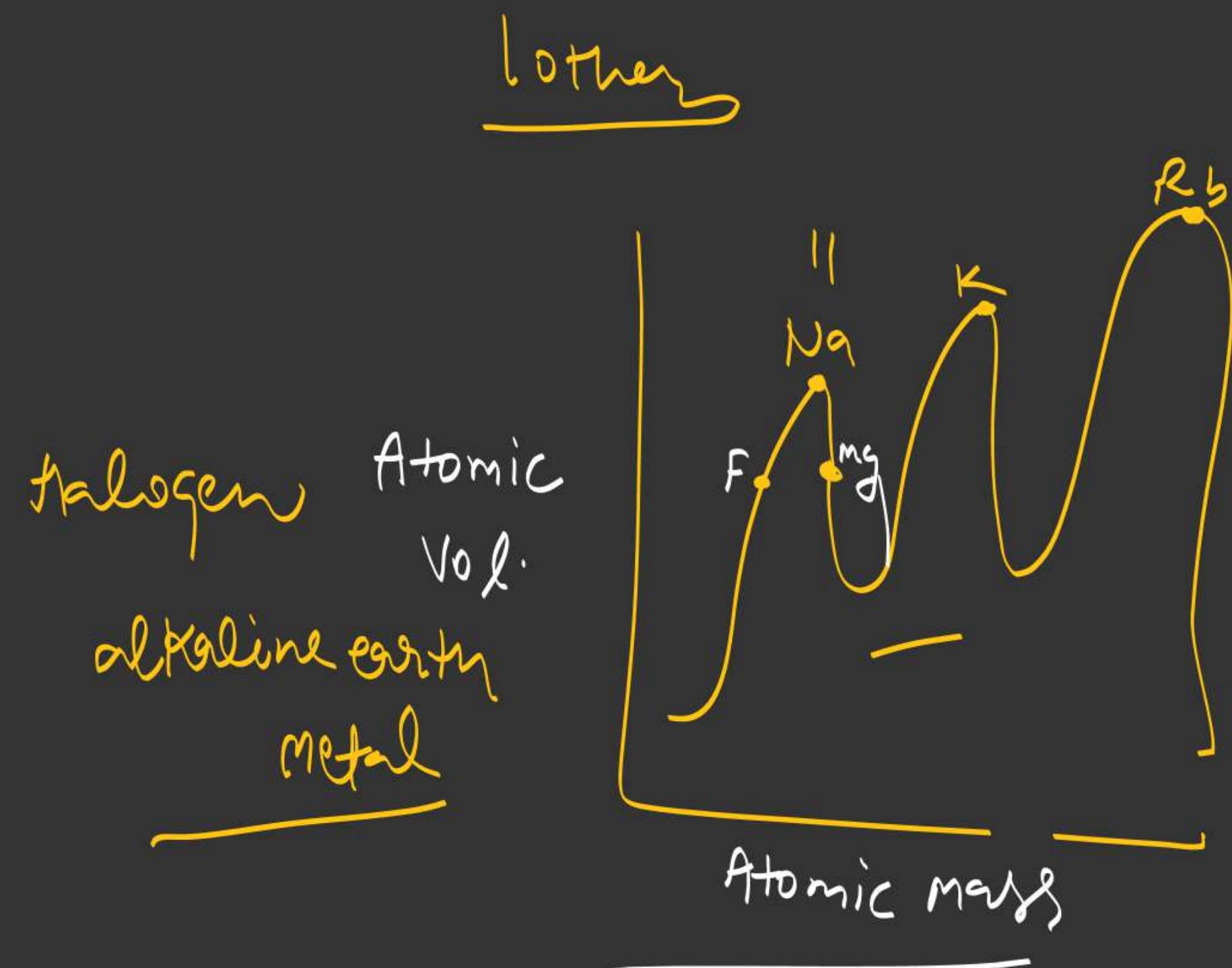
G

$$\underline{\underline{Be}} = 2^{nd}$$

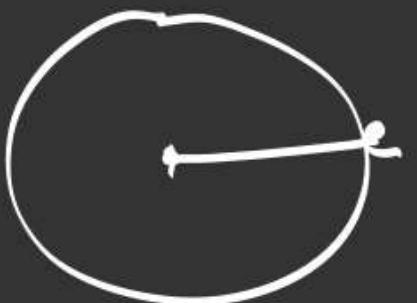
$$\underline{\underline{G_3}} \quad I_A \quad I_B$$



$$\underline{P+Au}$$



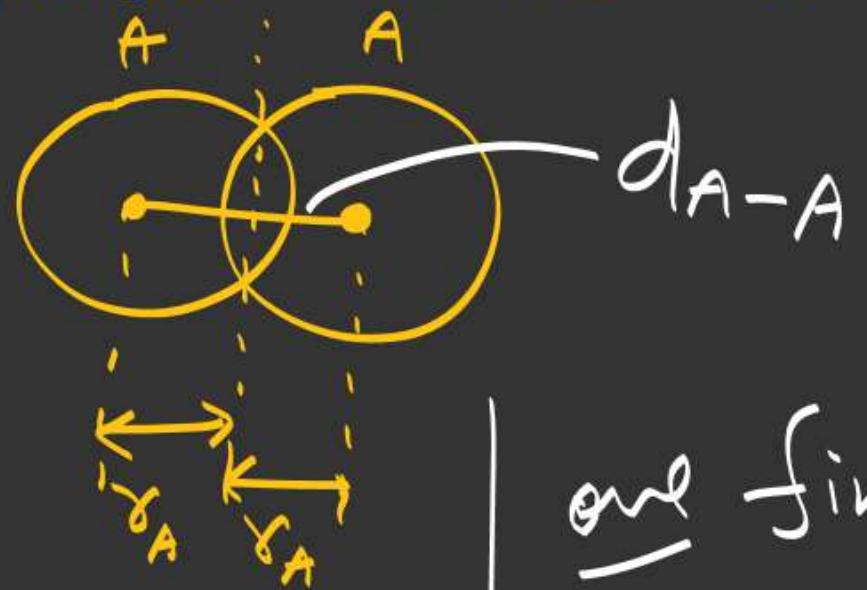
## Atomic Radii



- ① Covalent radii
- ② metallic radii
- ③ Ionic radii
- ④ v.w.R

## Covalent radii

(i) Homonuclear diatomic molecules



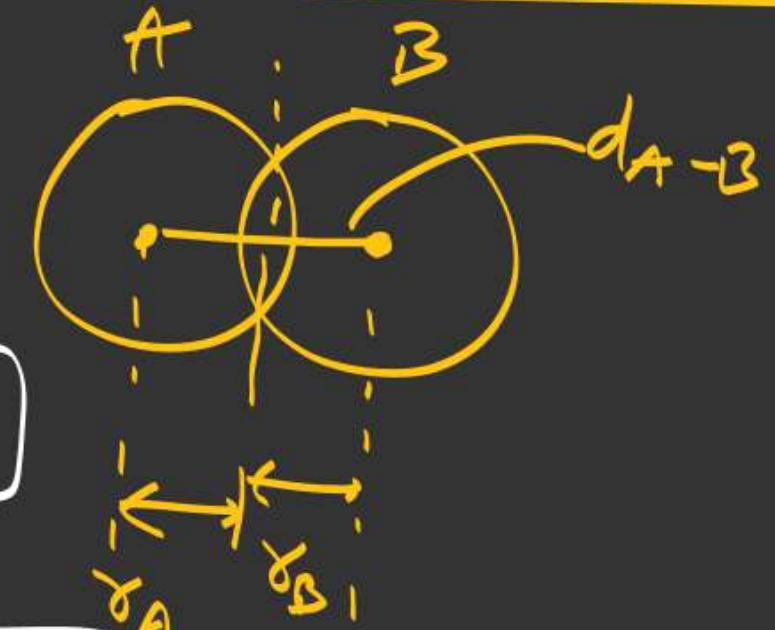
$$d_{A-A} = r_A + r_A$$

$$r_A = \frac{d_{A-A}}{2}$$

only find the  $r_{Cl}$  if  
 $d_{Cl-Cl}$  is  $1.98\text{ \AA}$

$$\frac{1.98}{2} = 0.99$$

## Heteronuclear diatomic molecule



$$d_{A-B} = r_A + r_B - 9 \Delta x \text{ pm}$$



$\epsilon \cdot N$  value

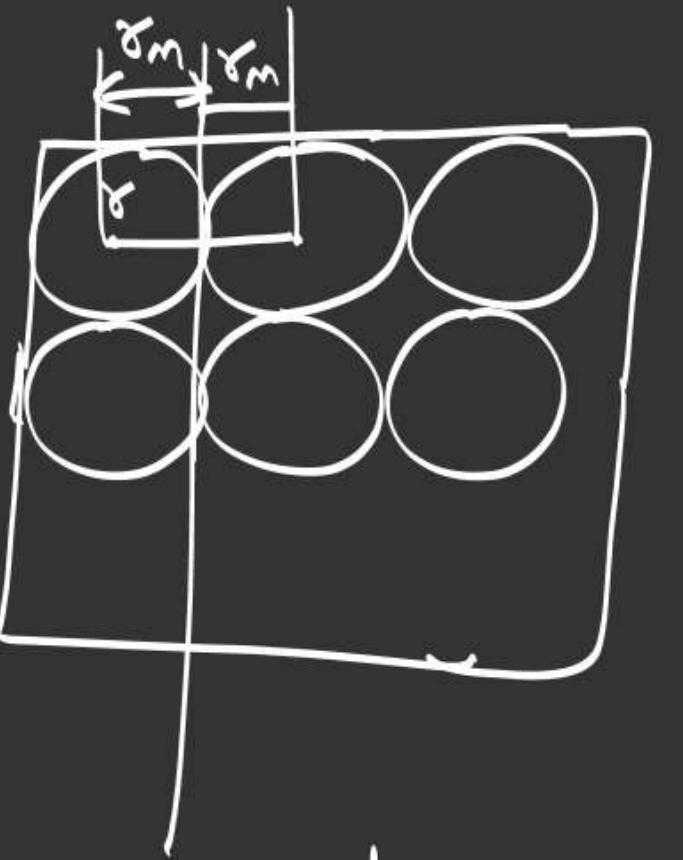
$$| 2.1 H$$

Showmaker Stevensen

$$d_{A-B} = r_A + r_B - 0.9 |\Delta x| \text{ } \text{Å}$$

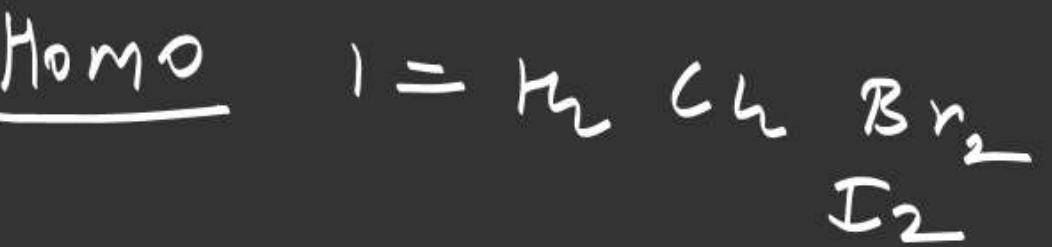
$$\Delta x = \epsilon \cdot N \text{ diff}$$

Metallic radii



$$d_{M-M} = r_m + r_m$$

$$r_m = \frac{d_{M-M}}{2}$$



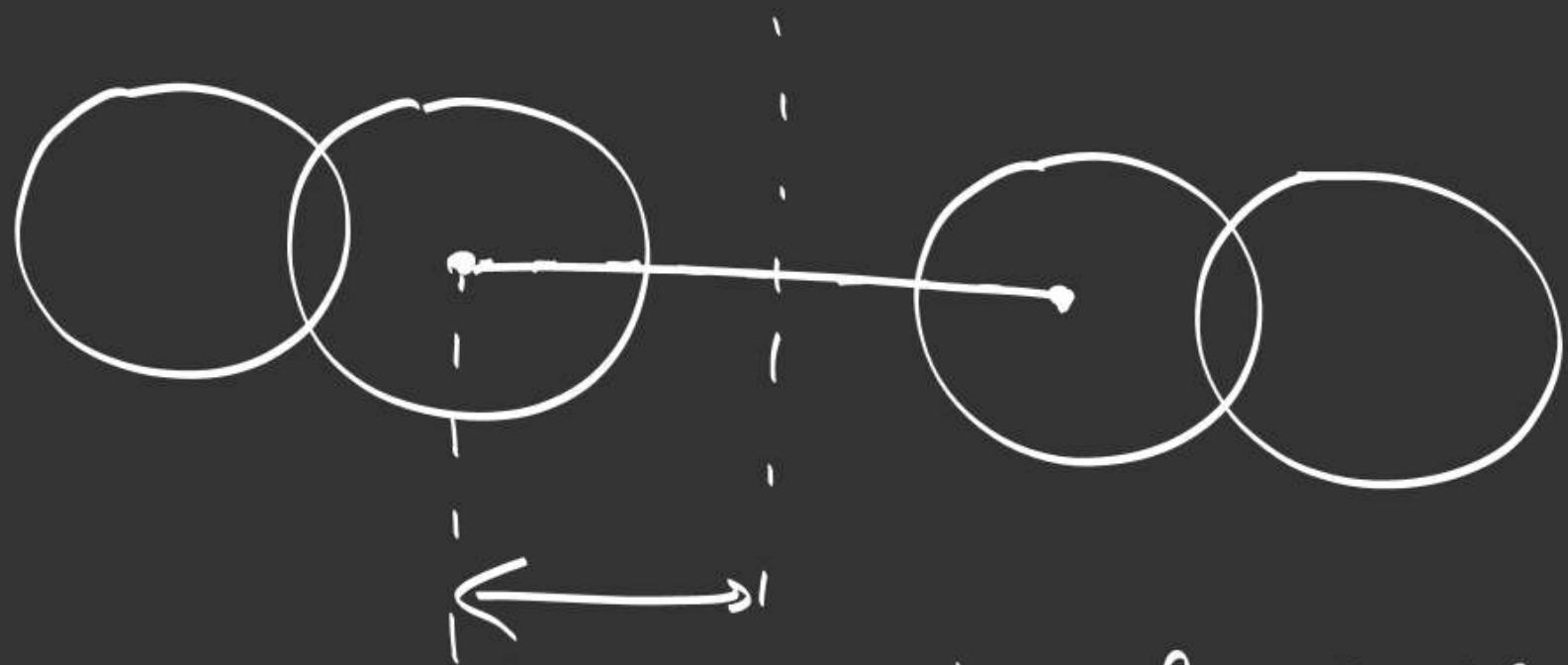
Hetero HF HCl HBr HI

and  $r_m = ?$

if  $d_{H-I} = 129 \text{ pm}$

$$r_H = \frac{129}{2} = 64.5 \text{ pm}$$

$$\frac{V \cdot w \cdot R}{}$$



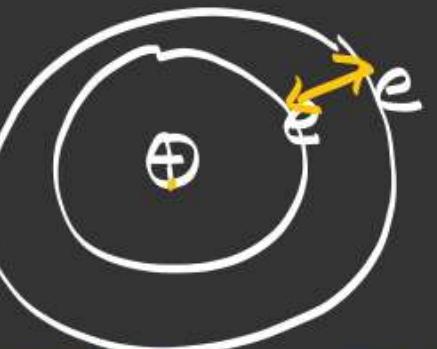
$$\frac{V \cdot w \cdot R > C \cdot R}{V \cdot w \cdot R > M \cdot R}$$

$V \cdot w \cdot R =$  gaseous molecule  
and Noble gas

factor affecting

- ①  $Z \uparrow \gamma \downarrow$
- ②  $Z_{eff} \uparrow \gamma \downarrow$
- ③  $n \uparrow \gamma \uparrow$
- ④  $\sigma \uparrow \gamma \uparrow$

$n = \text{Sheel}$



$$Z_{eff} = Z - \sigma$$

$S \cdot F$   
 $S > P > d > f$

## Ionic radii

Ion → cation  
 anion

## Cationic radii



$$\frac{e}{P} = \frac{10}{10} \quad \frac{9}{10} \quad \frac{8}{10} \quad \frac{7}{10}$$

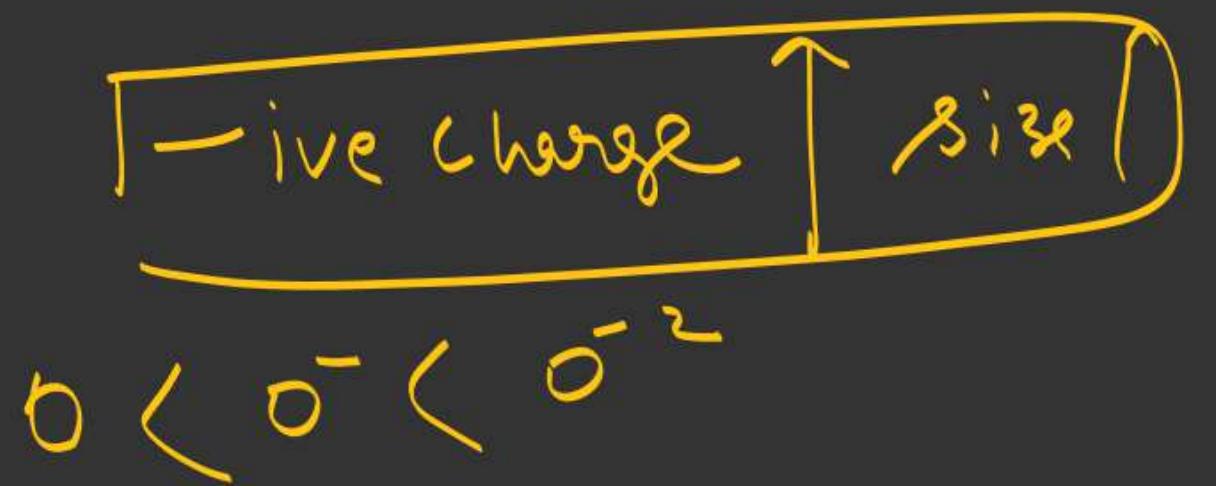


+ive charge  $\uparrow$  att  $\uparrow$   $\gamma \downarrow$

Anionic radii

$$\frac{e}{r} = \frac{10}{10} \quad \frac{11}{10} \quad \frac{12}{10}$$

$$N < N^- < N^{-2} < N^{-3}$$



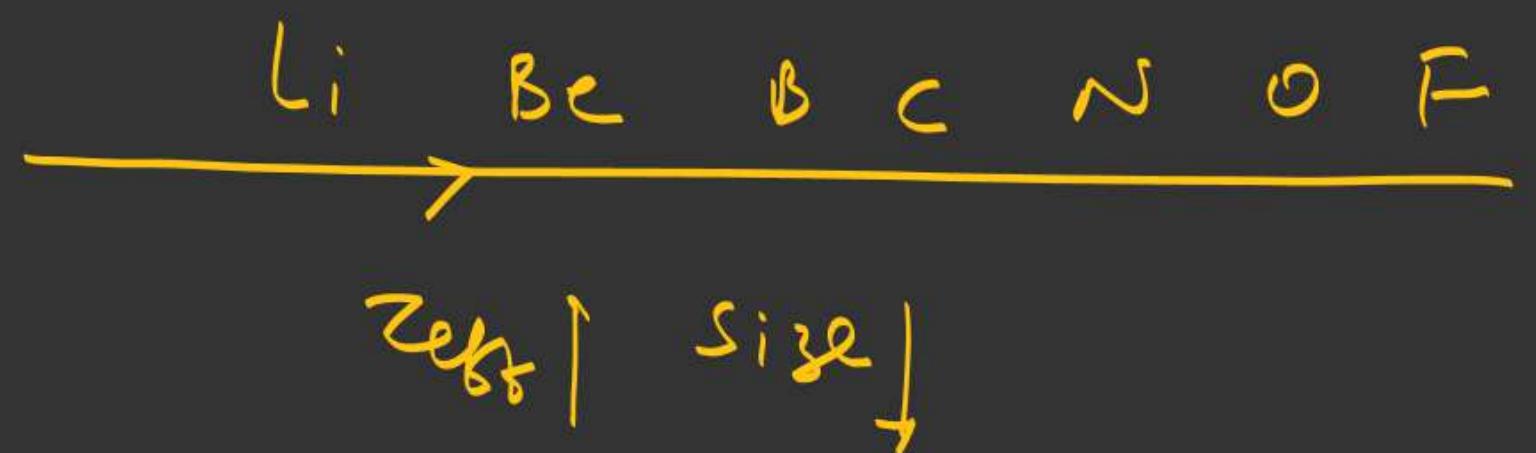
## iso electronic species

order of size

	$N^{-3}$	$O^{-2}$	$F^-$	$Na^+$	$Mg^{+2}$	$Al^{+3}$
$\frac{e}{r}$	$\frac{10}{7}$	$\frac{10}{8}$	$\frac{10}{9}$	$\frac{10}{11}$	$\frac{10}{12}$	$\frac{10}{13}$

-ine Charge ↑ size ↑

along the period



but Noble gas Higher size  
due to their V.W.R



Gmp'

along the group.

B

$$\text{B} < \text{Ga} < \text{Al} < \text{In} < \text{Tl}$$

143 pm Al

Al =  $1s^2 2s^2 2p^6 3s^2 3p^1$

135 pm Ga

due to poor S.E of 3d subshell

In

$$\text{In} = \text{1s}^2 \text{2s}^2 \text{2p}^6 \text{3s}^2 \text{3p}^6 \text{4s}^2 \text{3d}^{10} \text{4p}^1$$

Tl



# d-block

Sc Ti V Cr Mn Ti Co Ni Cu Zn  
 $Z_{eff} > S.E$        $Z_{eff} \approx S.E$        $Z_{eff} < S.E$   
 $\gamma \downarrow$   
 Size ↑

7	4	1	1	1	1	1	1	1	1
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# Group

3d series      Sc   Ti   . . . . .   Zn

4d series      Y   Zr   - - -   Cd

5d series      57 La — Hf   - - -   Mg  
 89 Ac      |  
 104      |  
 KLu/Rf

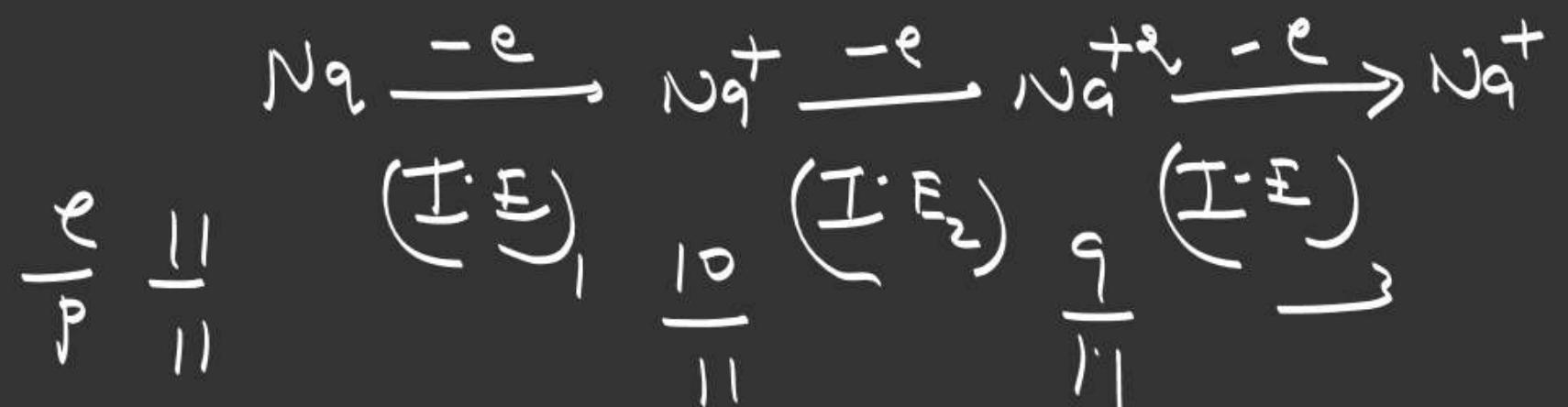
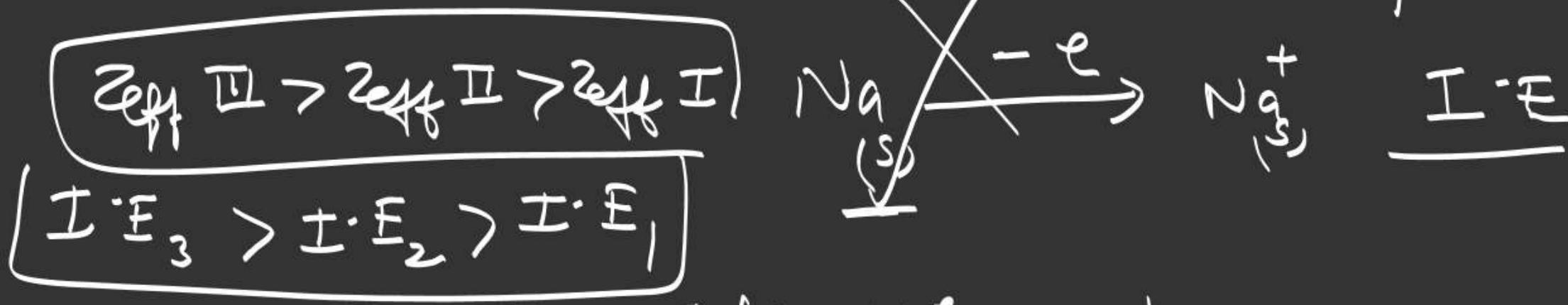
Ce — Lu   4f series  
 58      71

Th — Lr   5f series  
 90      103

3d series < 4d series

Element      Element  $\approx$  5d series [ due to poor S.E. of 4f element  
 Dub shell / Lanthanide cont. ]

I.E  $\rightarrow$  amount of req. energy for removal of an  $e^-$  from gaseous atom.

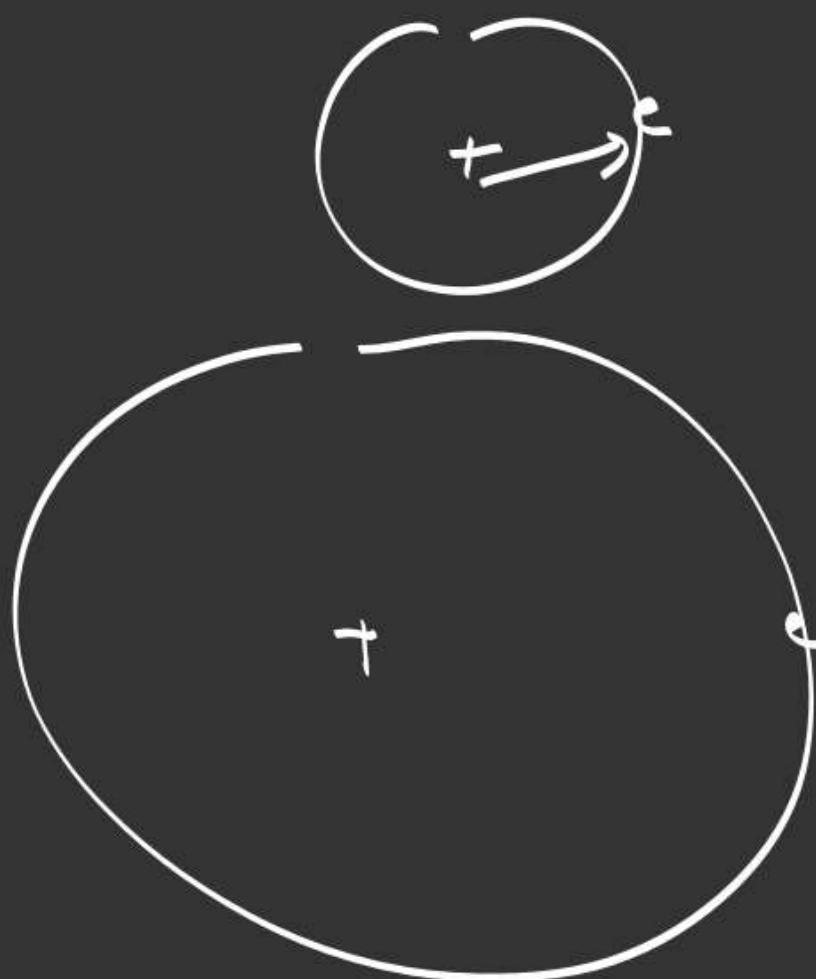


Unit

ev/atom

$$1 \text{ ev/atom} = 96.4 \text{ KJ/mole}$$

$$1 \text{ ev/atom} = 23.1 \text{ Kcal/mole}$$



factors

- ①  $Z \uparrow I \cdot E \uparrow$
- ②  $r_{\text{eff}} \uparrow I \cdot E \uparrow$
- ③  $n(\text{shell}) \uparrow I \cdot E \downarrow$
- ④  $\sigma \uparrow I \cdot E \downarrow$

Half filled / fully filled.

$$\begin{matrix} N & > & 0 \\ 2p^3 & & 2p^4 \end{matrix}$$



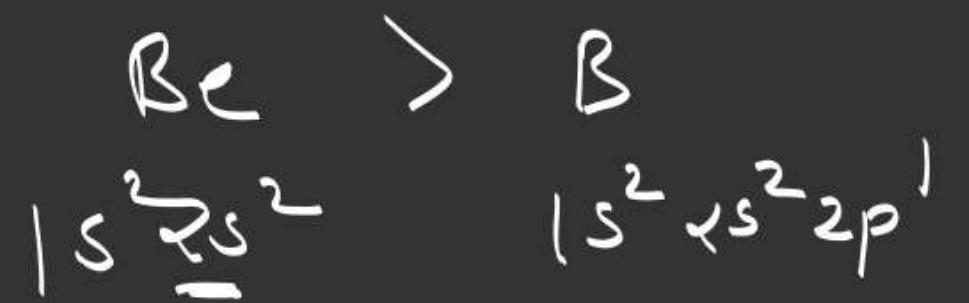
$$P > S$$

$$As > S_b$$

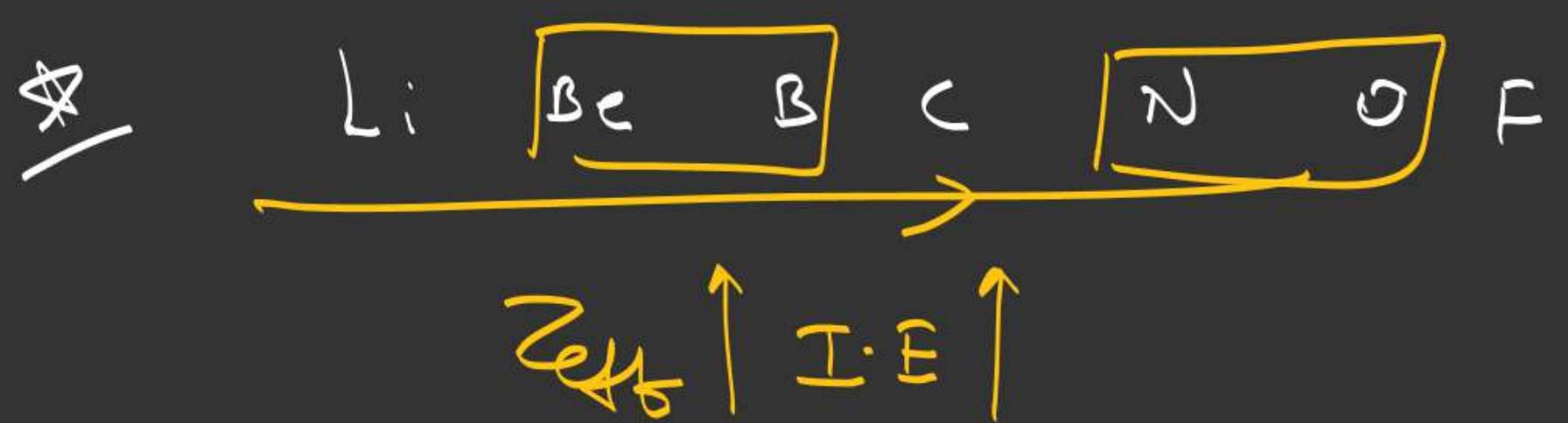
Penetration { Closeness towards nucleus }

s p d f      +      nucleus

⊕ ) ) ) )



along the period



Note  $\Rightarrow$  Noble gases have higher I.E. in their respective period.

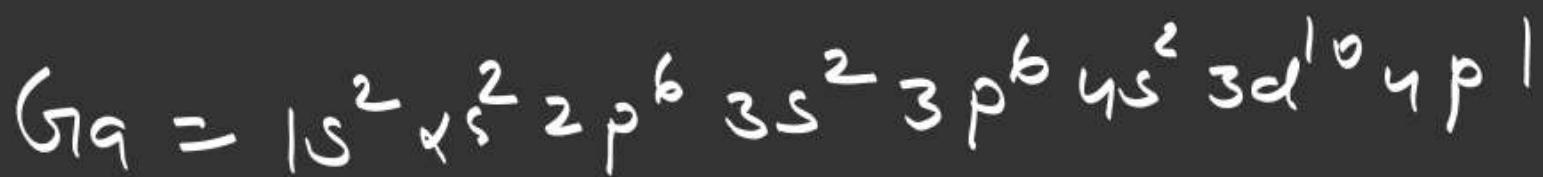
group 1

due to poor SE of 4f subshell

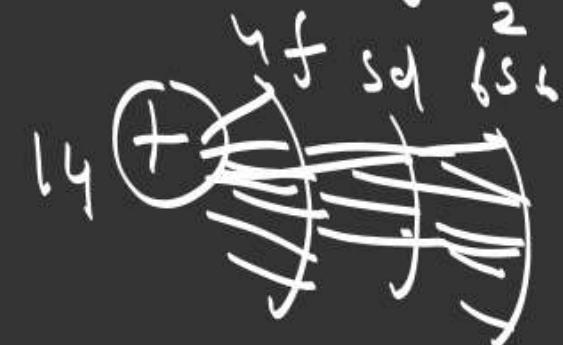
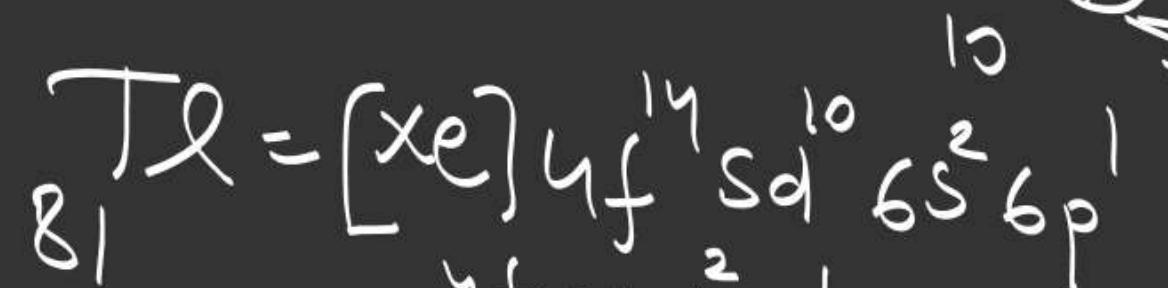
B

{ Al  
Ga }

In



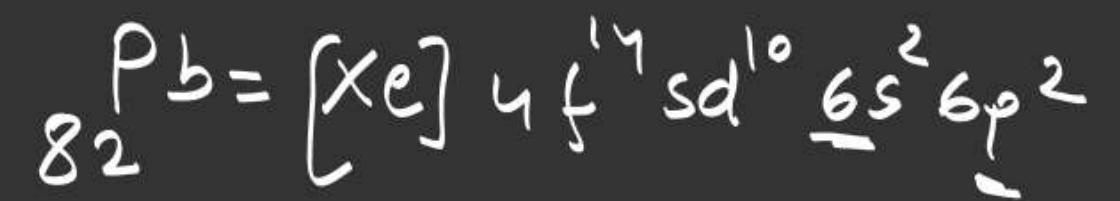
Te





Inert pair effect

The tendency of inertness of  
ns e<sup>-</sup> towards bonding  
is called inert pair effect.



C  
Si  
Ge  
Pb

C > Si > Ge > Pb > Sn

Li  
Na  
K  
Rb  
Cs  
Fr  
87

$\text{Li} > \text{Na} > \text{K} > \text{Rb} > \text{Fr} > \text{Cs}$

$\text{Be} > \text{Mg} > \text{Ca} > \text{Sr} > \text{Rb} > \text{Ba}$

due to poor

S.E of  $n_f$  subshell



3d series	Sc	Ti	- - -	Zn
4d series	Y	Zr	- - -	Cd
5d series	La	Hf	- . -	Hg
	Ac	Ku		

3d series > 4d series < 5d series  
 element                    element                    element

order of  $\pm E$

$N < O^+$

$2p^3 \quad 2p^3$

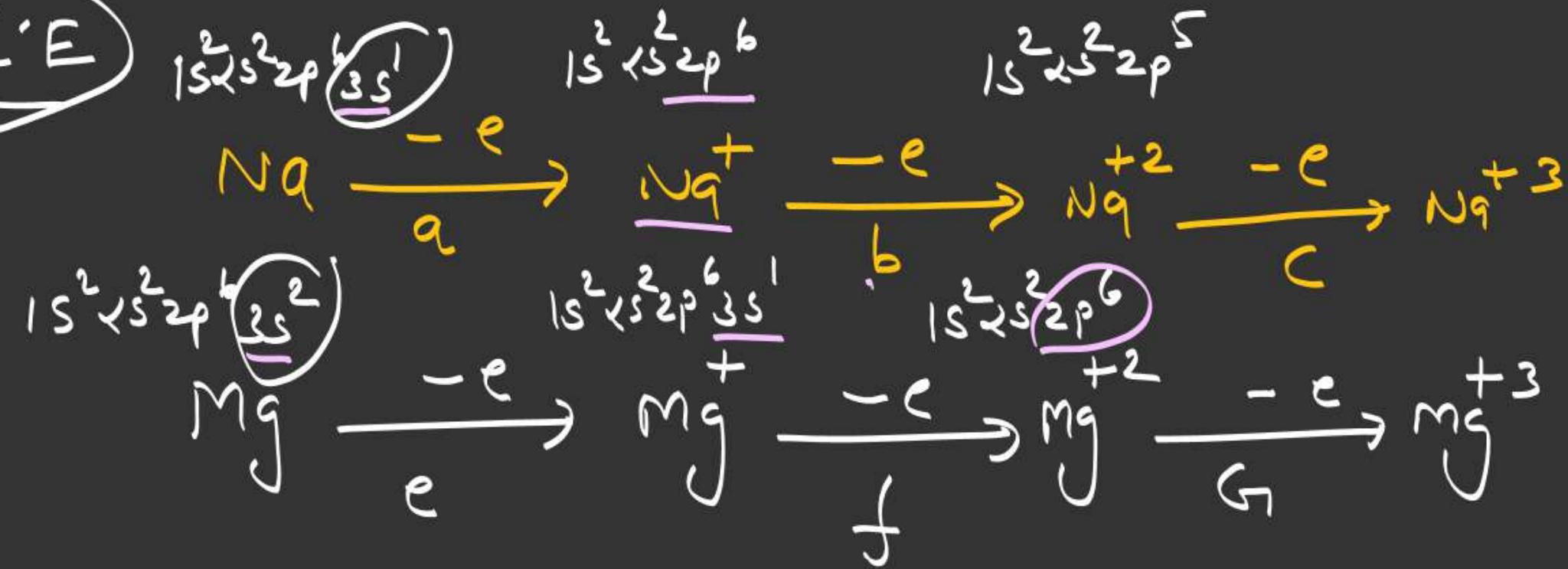
(+ive Charge)

Successive I·E

$$a < b < c$$

$e < f < g$

I·E



$a < e$   
 $b > f$   
 $g > c$

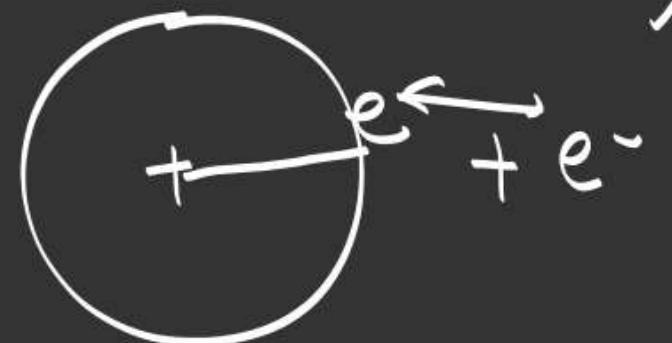
Compare I·E

$a$	$b$	$b$	$c$	$e$	$f$	$f$	$g$
$b$	$c$	$f$	$g$				
$a$	$e$						

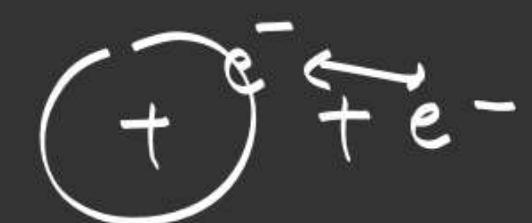
$$\underline{\Delta H_{\text{eg}}}$$

amount of release energy

When one electron is added in to  
gallows atom



most of the case att. dominant



but in II  $\Delta H_{\text{eg}}$  electronic recip.  $\uparrow$   
 So energy neg. for addition of  $e^-$



We can not compare

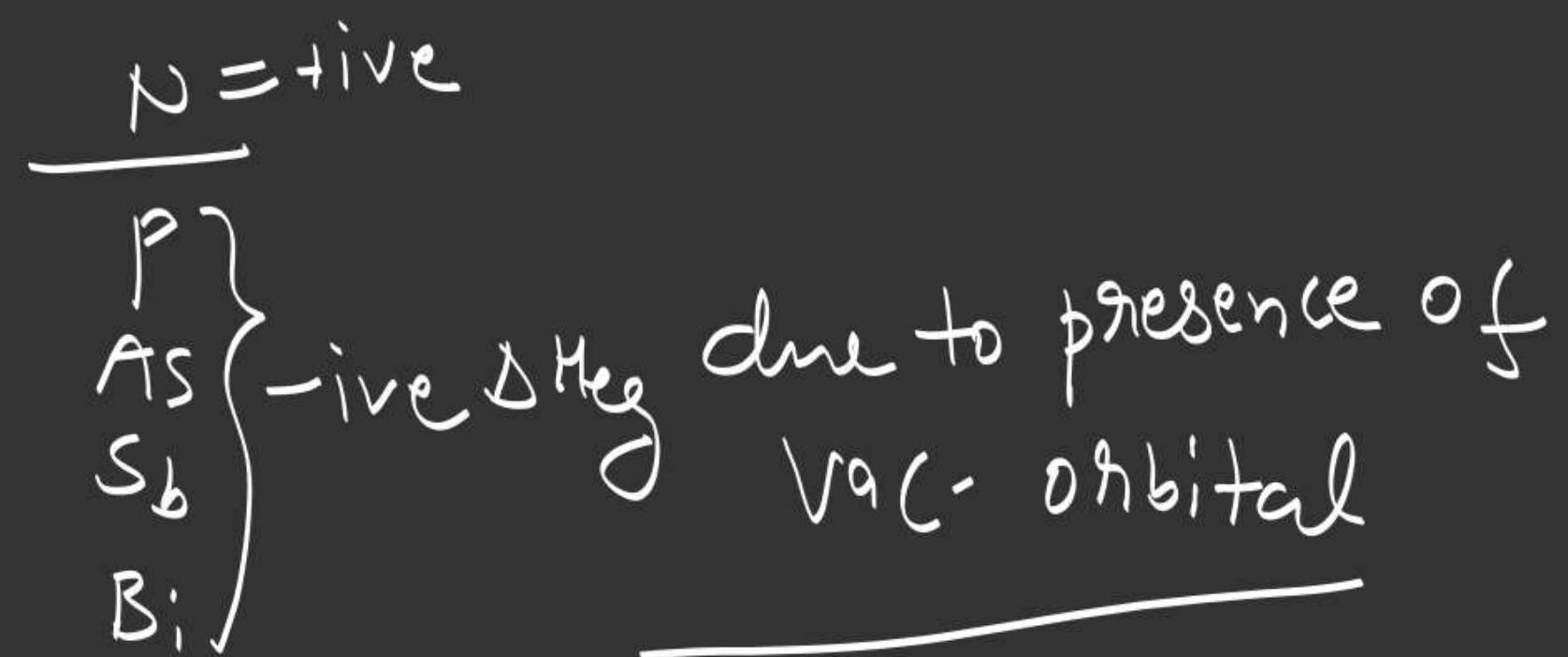
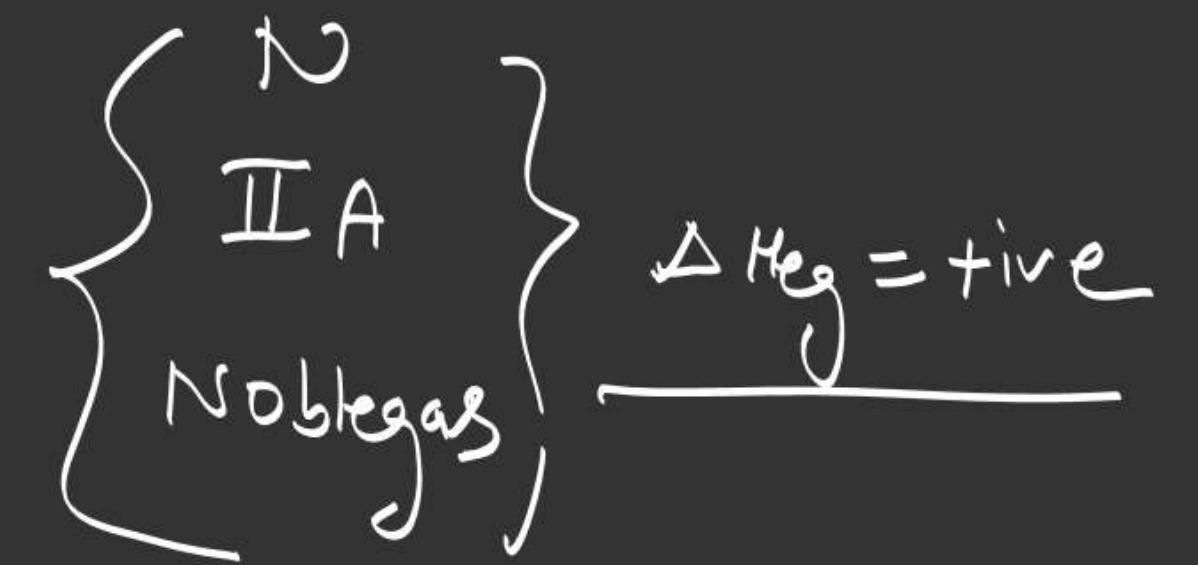
I and II  $\Delta H_{eg}$

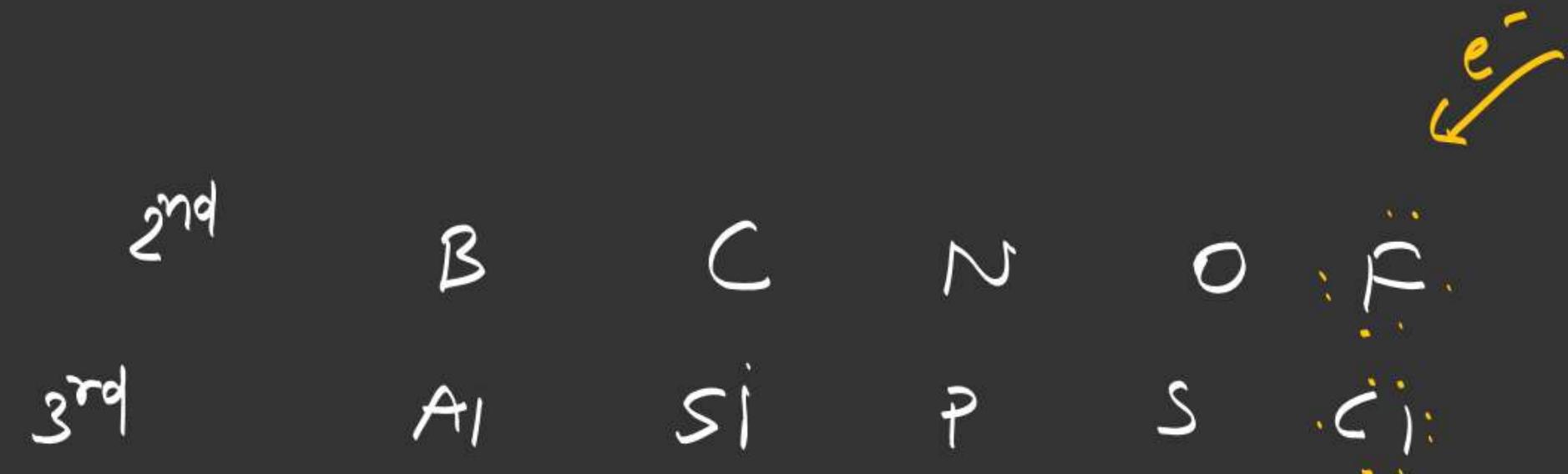
However

$$\left| \Delta H_{eg} II \right| > \left| \Delta H_{eg} I \right|$$

$$\underline{\Delta H_{eg_2} + \Delta H_{eg_1} > 0}$$

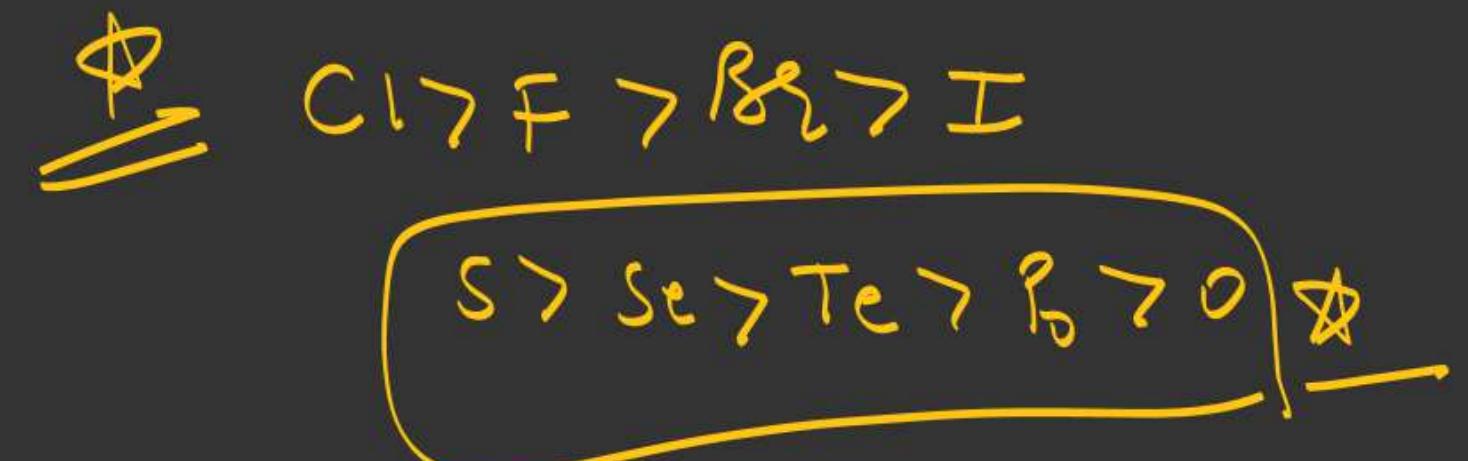
$$\frac{\text{neg. } \Delta H_{\text{eg}}}{\text{neg. } \Delta H_{\text{eg}}} \begin{array}{c} \textcircled{1} \quad z \uparrow \quad \text{neg. } \Delta H_{\text{eg}} \uparrow \\ \textcircled{2} \quad z_{\text{eff}} \uparrow \quad \text{neg. } \Delta H_{\text{eg}} \uparrow \\ \textcircled{3} \quad n \uparrow \quad \text{neg. } \Delta H_{\text{eg}} \downarrow \\ \textcircled{4} \quad \bar{n} \uparrow \quad \text{neg. } \Delta H_{\text{eg}} \downarrow \end{array}$$





$2^{\text{nd}} \text{ period} < 3^{\text{rd}} \text{ period}$

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highest neg  $\Delta H_{eg} = Cl$

$\Delta H_{eg}$

PYLS

$K_e = 48$   
 $N_c = 116$   
 $A_R = 96$   
 $K_S = 96$   
 $X_e = 78$

# PERIODIC TABLE

(A) Covalent radius

(B) Ionic radius

(C) Metallic radius

(D) Vander Waal's radius

(A) Covalent Radius: (SBCR -Single Bonded Covalent Radius)

(a) Covalent bonds are formed by overlapping of atomic orbitals.

(b) Internuclear distance is minimum in this case.

(c) Covalent radius is the half of the internuclear distance between two singly bonded homo atoms.

Ex. If internuclear distance of  $A_2$  molecule is (A-A) and covalent radius is r then

$$\overline{d_{A-A}} = \overline{r_A} + \overline{r_A} \text{ or } 2\overline{r_A}$$

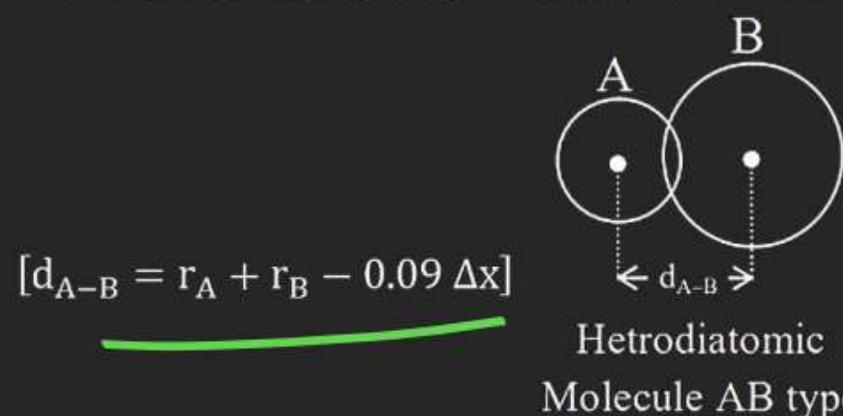
$$\overline{r_A} = \frac{\overline{d_{A-A}}}{2}$$

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**Ex.** In  $\text{Cl}_2$  molecule, internuclear distance is  $1.98\text{\AA}$  so  $r_{\text{Cl}} = \frac{1.98}{2} = 0.99\text{\AA}$ .

**Note:** In a heterodiatomic molecule AB where the electronegativity of atoms A and B are different, the experimental values of internuclear distance  $d_{A-B}$  is less than the theoretical values ( $r_A + r_B$ ).

According to Schomaker and Stevenson –



Where  $\Delta x$  is the difference of electronegativities of the atoms A and B ·  $r_A$  and  $r_B$  are in Å.

**Example:** X – X bond length is  $1.00\text{\AA}$  and C – C bond length is  $1.54\text{\AA}$ . If electronegativities of X and C are 3.0 and 2.0 respectively, then C – X bond length is likely to be? (using Stevenson & schomaker formula).

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Sol.  $r_{C-X} = r_C + r_X - 0.09\Delta x$

$$\begin{aligned}&= \frac{1.00}{2} + \frac{1.54}{2} - 0.09 \quad [\Delta x = 1] \\&= 1.27 - 0.09;\end{aligned}$$

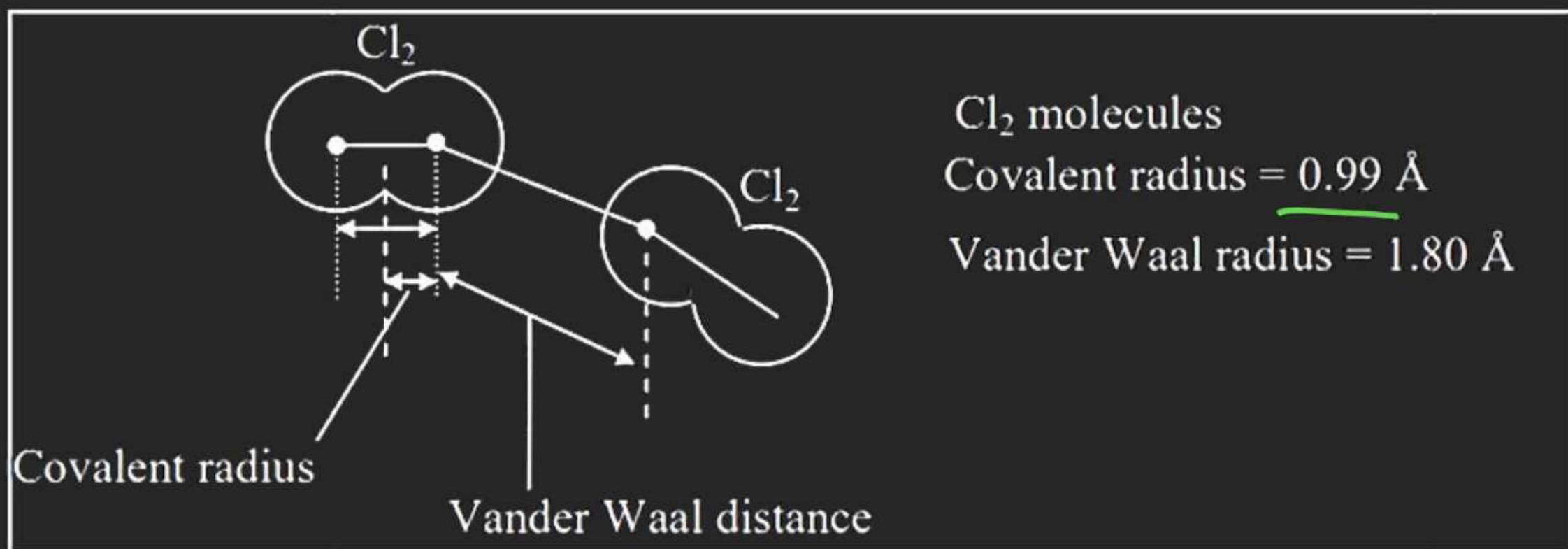
Ans. C – X bond length = 1.18 Å

(B) Ionic Radius:

(i) **Cationic Radius:**

- (a) When a neutral atom loses electron it converts into a cation (positive charged ion)
- (b) Cationic radius is always smaller than atomic radius because after losing electron number of electron reduces, but number of protons remains same, due to this  $Z_{eff}$  increases, hence electrons get pulled towards nucleus and atomic radius decreases, moreover after losing all the electrons from the outer most shell, penultimate shell becomes ultimate shell which is nearer to nucleus so size decreases.

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**van der Waal's radius =  $2 \times$  covalent radius**

**van der Waal's radius > Metallic radius > Covalent radius**

**Factors affecting atomic size are:**

**(a)  $Z_{\text{eff}}$  increases, atomic radius decreases**

**Ex.  $\text{Li} > \text{Be} > \text{B} > \text{C} > \text{N} > \text{O} > \text{F}$**

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(ii) In a group : It increases from top to bottom in a group as number of shell increases

Ex.  $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$

**Exceptions: Transition elements**

**From left to right in a period:**

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In the first transition series the atomic radii slightly decreases from Sc to Mn because effect of effective nuclear charge is more prominent than the shielding effect. After that it remains almost the same because both the effects balance each other. The atomic size from Cu to Zn slightly increases because shielding effect is more prominent than effective nuclear charge due to  $d^{10}$  configuration of Cu and Zn. The atomic radii of the elements of d transition series are as under:

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## Difference between electronegativity and Electron Affinity :

<u>Electronegativity</u>	<u>Electron Affinity</u>
 <ul style="list-style-type: none"><li>• Tendency of an atom in a molecule to attract the bonded electrons</li><li>• It is not an energetic term.</li><li>• It regularly increases in a period because it does not depends on stable electronic configuration</li><li>• It has no unit</li><li>• It is a relative term</li></ul>	$X + e^- \rightarrow X^-$ <ul style="list-style-type: none"><li>• Energy released when an electron is added to neutral isolated gaseous atom</li><li>• It is an energetic term.</li><li>• It does not increase regularly in a period because it depends on stable electronic configuration.</li><li>• It is measured in eV/atom or KJ/mol or K cal mole<sup>-1</sup></li><li>• It is an absolute term</li></ul>

# PERIODIC TABLE

(a) Charge on atom: The cation will be more electronegative than parent atom. As the positive charge on the cation increases electronegativity increases.

Order of E.N.

Ex.



The anion will be less electronegative than atom. As the negative charge on the anion Increases electronegativity decreases.

Ex.



(b)

Ionization energy and electron affinity: Higher the ionization energy and electron affinity of an atom, greater will be the its electronegativity.

(c)

Effect of Substitution: The electronegativity of an atom depends upon the nature of substituent attached to that atom. For example - Carbon atom in  $\text{CF}_3\text{I}$  acquires greater positive charge than in  $\text{CF}_3\text{I}$ . Hence, C atom in  $\text{CF}_3\text{I}$  is more electronegative than in  $\text{CH}_3\text{I}$ .

# PERIODIC TABLE

## OTHER SCALES OF ELECTRONEGATIVITY :

**Mulliken scale:** According to Mulliken electronegativity is average value of ionisation potential and electron affinity of an element,

$$X_m = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{2}$$

$$X_p = \frac{X_m}{2.8}$$

$X_p$  = Pauling's EN where (EN = electronegativity)

$X_m$  = Mulliken's EN

$$X_m = \frac{I\cdot E + E\cdot A}{2} \text{ ev/atom}$$

# PERIODIC TABLE

$Z =$  The actual number of charge present in the nucleus i.e. number of protons and  $\sigma =$  shielding constant.

Note:- (1)  $Z_{\text{eff}}$  to be calculated for  $(x + 1)^{\text{th}}$  valence  $e^-$ . (2)  $X_P = X_{\text{AR}} + 0.744$

## APPLICATIONS OF ELECTRONEGATIVITY

(a) Bond strength : If the electronegativity difference of covalently bonded atoms ( $\Delta X$ ) increases, the bond energy of the covalent bond also increases. For example the order of the H – X bond strength is



As the bond strength decreases, the acidic strength increases.

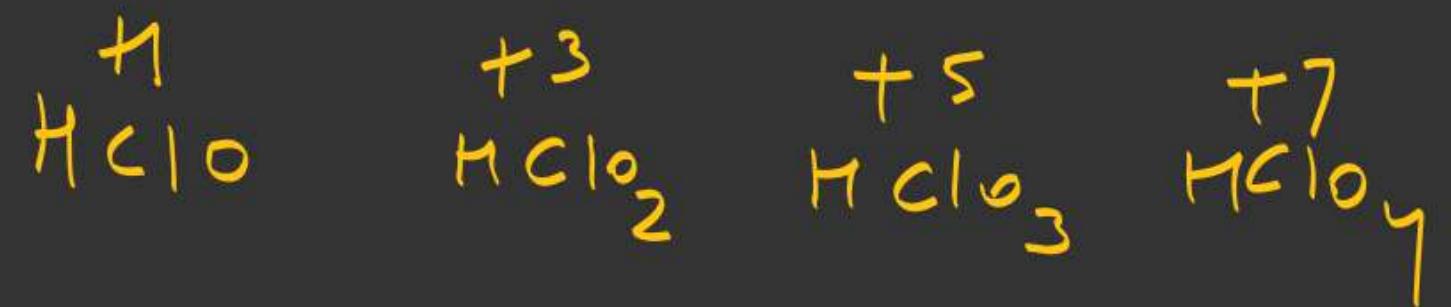
Order of increasing acidic strength is:-





Bond strength > >

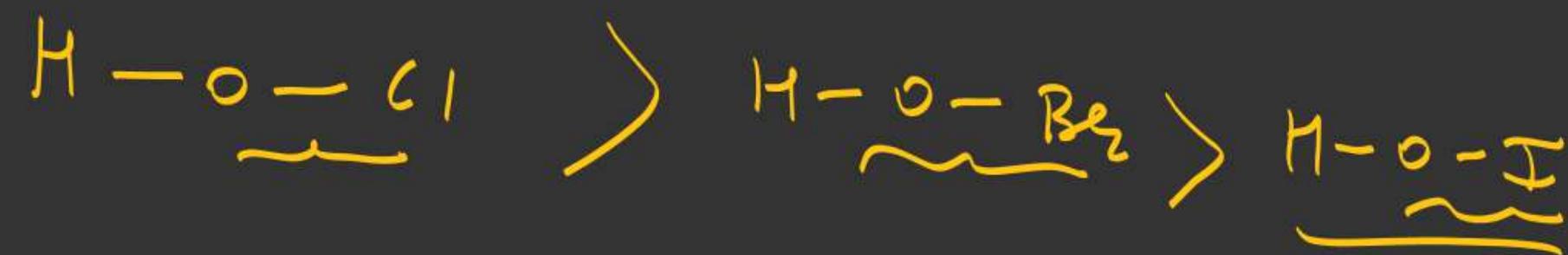
titration ↑ & n↑



% of Ionic Ch.  
Henry and Smith

$$\frac{\% \text{ I.C.}}{\% \text{ I.C.}} = 16\Delta + 3.5\Delta^2$$

$$\Delta = \epsilon \cdot N \text{ diff}$$



## PERIODIC TABLE

2023

1. Inert gases have positive electron gain enthalpy. Its correct order is

(A)  $Xe < Kr < Ne < He$

(B)  $He < Ne < Kr < Xe$

~~(C)  $He < Xe < Kr < Ne$~~

(D)  $He < Kr < Xe < Ne$

Ans

$He = 18$

$Ne = 16$

$Ar = 16$

$Kr = 16$

$Xe = 78$

## PERIODIC TABLE

2. Which of the following represents the correct order of metallic character of the given elements?

(A) ~~Si < Be < Mg < K~~

(2) Be < Si < Mg < K

(3) K < Mg < Be < Si

(4) Be < Si < K < Mg



# PERIODIC TABLE

3. The bond dissociation energy is highest for

(A)  $\text{Cl}_2$

(B)  $\text{I}_2$

(C)  $\text{Br}_2$

(D)  $\text{F}_2$



# PERIODIC TABLE

**4. Match List – I with List – II**

List – I (Atomic number)	List – II (Block of periodic table)
<u>Rb</u> <u>(A) 37</u>	I. p-block
<u>(B) 78</u>	II. d-block
<u>(C) 52</u>	III. f-block
<u>(D) 65</u>	IV. s-block

**Choose the correct answer from the options given below:**

- (A) A – II, B – IV, C – I, D – III
- (B) A – I, B – III, C – IV, D – II
- (C) A – IV, B – III, C – II, D – I
- (D) A – IV, B – II, C – I, D – III

## PERIODIC TABLE

5. The correct increasing order of the ionic radii is

(A)  $\text{Cl}^- < \text{Ca}^{2+} < \text{K}^+ < \text{S}^{2-}$       (B)  $\text{K}^+ < \text{S}^{2-} < \text{Ca}^{2+} < \text{Cl}^-$

(C)  $\text{S}^{2-} < \text{Cl}^- < \text{Ca}^{2+} < \text{K}^+$       (D)  $\text{Ca}^{2+} < \text{K}^+ < \text{Cl}^- < \underline{\text{S}^{2-}}$

- ionic charge ↑ size ↑

## PERIODIC TABLE

6.  $\text{Nd}^{2+} = \underline{\hspace{2cm}}$

- (A)  $4f^26s^2$       (B)  $4f^4$       (C)  $4f^3$       (D)  $4f^46s^2$

(4)



## PERIODIC TABLE

7. Given below are two statements: one is labelled as Assertion(A) and the other is labelled as Reason (R)

**Assertion (A) :** The first ionization enthalpy of 3d series elements is more than that of group 2 metals

**Reason (R) :** In 3d series of elements successive filling of d-orbitals takes place.

In the light of the above statements, choose the correct answer from the options given below:

- (A) Both (A) and (R) are true and (R) is the correct explanation of (A)
- (B) Both (A) and (R) are true but (R) is not the correct explanation of (A)
- (C) (A) is false but (R) is true
- (D) (A) is true but (R) is false

$3d^1 4s^2$   $3d^2 4s^2$   $3d^3 4s^2$   $3d^5 4s^1$   
Sc Ti V Cr Mn Fe Co Ni Cu Zn

## PERIODIC TABLE

9. For electron gain enthalpies of the elements denoted as  $\Delta_{eg}H$ , the incorrect option is :

(A)  $\Delta_{eg}H(\text{Cl}) < \Delta_{eg}H(\text{F})$

(C)  $\Delta_{eg}H(\text{I}) < \Delta_{eg}H(\text{At})$

(B)  $\Delta_{eg}H(\text{Se}) < \Delta_{eg}H(\text{S})$

(D)  $\Delta_{eg}H(\text{Te}) < \Delta_{eg}H(\text{Po})$

$S > \text{Se} > \text{Te} > \text{Po} > \text{O}$

# PERIODIC TABLE

10. Metals generally melt at very high temperature. Amongst the following, the metal with the highest melting point will be

- (A) Hg
  - (B) Ag
  - (C) Ga
  - (D) Cs
- Handwritten notes:
- (A) Hg: liq. at room temp.
  - (B) Ag: d-block
  - (C) Ga: melt at summer
  - (D) Cs: liq. at 27°C

## PERIODIC TABLE

11. The correct order of electron gain enthalpies of Cl, F, Te and Po is

(A) F < Cl < Te < Po

(B) Po < Te < F < Cl

(C) Te < Po < Cl < F

(D) Cl < F < Te < Po

## PERIODIC TABLE

12. Which of the following elements is considered as a metalloid?

- (A) Sc
- (B) Pb
- (C) Bi
- ~~(D) Te~~

## PERIODIC TABLE

13. Given below are two statements: one is labelled as Assertion (A) and the other is labelled as Reason(R)

**Assertion (A):** The ionic radii of  $O^{2-}$  and  $Mg^{2+}$  are same.

**Reason (R) :** Both  $O^{2-}$  and  $Mg^{2+}$  are isoelectronic species

In the light of the above statements, choose the correct answer from the options given below

- (A) Both (A) and (R) are true and (R) is the correct explanation of (A)
- (B) Both (A) and (R) are true but (R) is not the correct explanation of (A)
- (C) (A) is True but (R) is false
- (D) (A) is false but (R) is true

## PERIODIC TABLE

14. The correct order of increasing ionic radii is

- (A)  $\text{Mg}^{2+} < \text{Na}^+ < \text{F}^- < \text{O}^{2-} < \text{N}^{3-}$
- (B)  $\text{N}^{3-} < \text{O}^{2-} < \text{F}^- < \text{Na}^+ < \text{Mg}^{2+}$
- (C)  $\text{F}^- < \text{Na}^+ < \text{O}^{2-} < \text{Mg}^{2+} < \text{N}^{3-}$
- (D)  $\text{Na}^+ < \text{F}^- < \text{Mg}^{2+} < \text{O}^{2-} < \text{N}^{3-}$

isoelectronic series

-ive charge ↑ size ↑

## PERIODIC TABLE

15. Element "E" belongs to the period 4 and group 16 of the periodic table. The valence shell electron configuration of the element, which is just above 'E' in the group is
- (A)  ~~$3s^23p^4$~~       (B)  $3d^{10}4s^2, 4p^4$   
(C)  $4d^{10}5s^2, 5p^4$       (D)  $2s^2, p^4$

$3s^2 3p^4$

# PERIODIC TABLE

16. Match List-I with List-II.

## List-I (Oxide)



## List-II (Nature)

(I) Amphoteric

(II) Basic

(III) Neutral

(IV) Acidic

Choose the correct answer from the options given below:

(A) (A) - (IV), (B) - (III), (C) - (I), (D) - (II)

(C) (A)-(II), (B) - (IV), (C) - (III), (D) (I)

(B) (A) - (IV), (B) - (II), (C) - (I), (D) - (III)

(D) (A)-(I), (B) - (II), (C) - (III), (D) - (IV)

$\text{N}_2\text{O}$   $\text{NO}$   $\text{CO}$   $\text{H}_2\text{O}$

neutral

S-block element  $\Rightarrow$  Basic  
oxide except

$\text{BeO}$   
Amphoteric

$\text{Pb} \quad \text{Zn} \quad \text{Be} \quad \text{Al} \quad \text{Ga} \quad \text{Sn} \quad \text{Cr}$

+3 | +4

all the oxides and hydroxides are  
Amphoteric

$\text{V}_2\text{O}_5 \quad \text{As}_2\text{O}_3 \quad \text{Sb}_2\text{O}_3$

## PERIODIC TABLE

17. Given below are two statements. One is labelled as Assertion A and the other is labelled as Reason R.

**Assertion A:** The first ionization enthalpy for oxygen is lower than that of nitrogen.

**Reason R:** The four electrons in 2p orbitals of oxygen experience more electron-electron repulsion.

$N > O$

In the light of the above statements, choose the correct answer from the options given below.

- (A) Both A and R are correct and R is the correct explanation of A.
- (B) Both A and R are correct but R is Not the correct explanation of A.
- (C) A is correct but R is not correct.
- (D) A is not correct but R is correct

## PERIODIC TABLE

18. Identify the correct order of standard enthalpy of formation of sodium halides.

(A)  $\text{NaI} < \text{NaBr} < \text{NaF} < \text{NaCl}$

~~(B)  $\text{NaI} < \text{NaBr} < \text{NaCl} < \text{NaF}$~~

(C)  $\text{NaF} < \text{NaCl} < \text{NaBr} < \text{NaI}$

(D)  $\text{NaCl} < \text{NaF} < \text{NaBr} < \text{NaI}$

