

Periodic Table:

Tables in which elements are arranged in order of increasing atomic number in the manner that elements with same properties fall in same vertical column.

Elements in horizontal rows → period
Vertical rows → group

Earlier first successful periodic ~~table~~ law was developed by Russian chemist

Mendeleev

acc. to he divided elements on the basis of chemical & physical properties

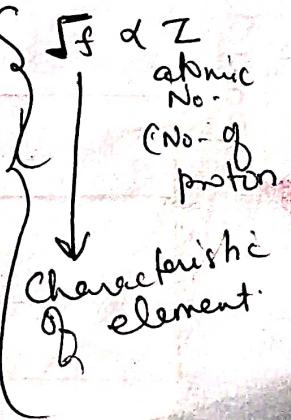
These phy. & chem. properties are periodic function of atomic weight

H/d as Periodic law

But later on, another physicist (Henry Moseley) discovered new law ~~of modern periodic law~~
Acc. to that, physical & chemical properties of elements are periodic function of its atomic number.

H/d as modern periodic law

Based on developed long-form of periodic table.



S-block (1-2 groups)

Non metals

	13	14	15	16	17	18
← P-block C_{13-18} (pp)	3, 8	10, 18, 13, 2			Hc ₂	

14 elements

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Lanthanide
Actinide

F-block

S-Block

Electronic Conf^n

S-block (Group - 1 & 2)

→ If last electron enters in s orbital \rightarrow placed in s orbital

An observation of the electronic conf. of the elements of group 1 & 2 shows that the last electron enters to the ns orbital of the last shell.

where $n =$ no. of period in which element of gp. 1 or 2 enters.

i.e. S block elements are those elements in which the last electron enters to n_s orbital / in which one orbital is being filled up successively with electrons.

Valence shell conf. of s-block $\rightarrow n_s^1$ or n_s^2

nS^{1-2}

for gp 1

for gp 2

example:

$H_{z=1} = 1s^1 \rightarrow$ 1 period, gp 1

$Ca_{z=20} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2 \rightarrow$ 4th period, gp 2

$Mg_{z=12} = 1s^2, 2s^2, 2p^6, 3s^2 \rightarrow$

3 period, gp 2

Gp. 1 \rightarrow Alkali metals

Group 2 \rightarrow Alkaline earth metals

P-block elements :

- Last electron enters into p-subshell
- elements
- Belong to 13-18 group of periodic table
- 18 group contains noble gas

→ Electronic config: $n\text{p}^{1-6}$ (for last electron)

Period II

$$2\text{s}^2 3\text{p}^{1-6}$$

8 elements

Period III

$$3\text{s}^2 3\text{p}^{1-6}$$

8 elements

Period IV

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

18 elements

Period V

$$5\text{s}^2 4\text{d}^{10} 5\text{p}^{1-6}$$

18 elements

Period VI

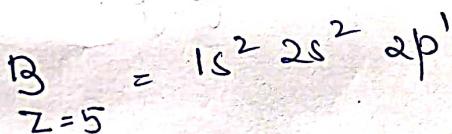
$$6\text{s}^2 4\text{f}^{14} 5\text{d}^{10} 6\text{p}^{1-6}$$

32 elements

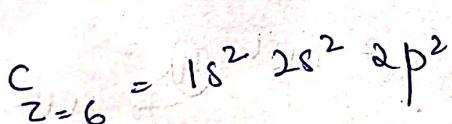
6th period contains 32 elements

7th period contains → only actinides & d-block elements

Example



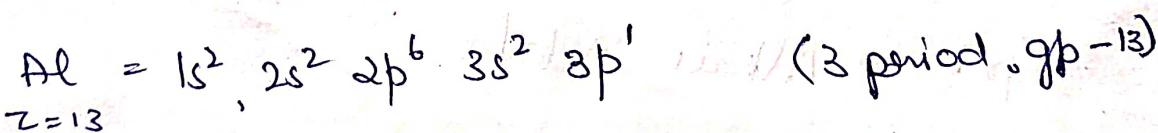
(2 period, 13 gp)



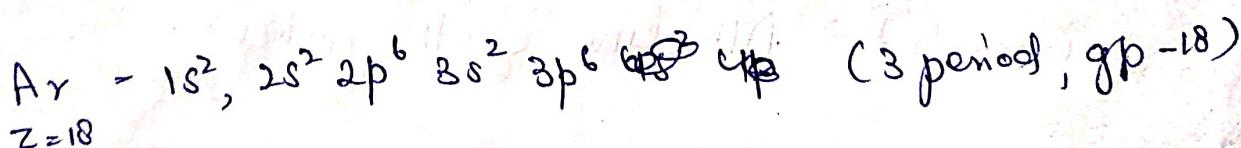
(2 period, 14 gp)



(2 period, 18 gp)



(3 period, gp-13)



(3 period, gp-18)

From above electronic config it is clear that ns orbital always contains 2 electrons (ns^2). However the electrons in np subshell is different, i.e., from 1-6 depending upon the position of element in the gp.

Group 13	$ns^2 np^1$	$\left. \begin{matrix} \\ \\ \\ \\ \\ \end{matrix} \right\}$	partially filled → completely filled noble gas
14	$ns^2 np^2$		
15	$ns^2 np^3$		
16	$ns^2 np^4$		
17	$ns^2 np^5$		
18	$ns^2 np^6$		
$ns^2 np^{1-6}$			

d-block : Electronic Configuration

- last electron enters in d subshell of penultimate shell
- General electronic config. $ns^{1-2} (n-1) d^{1-10}$

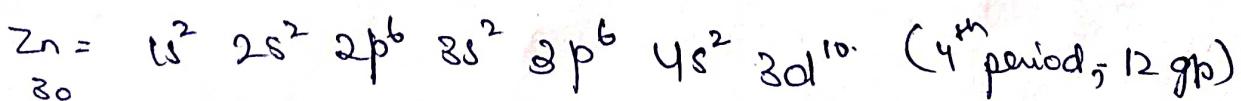
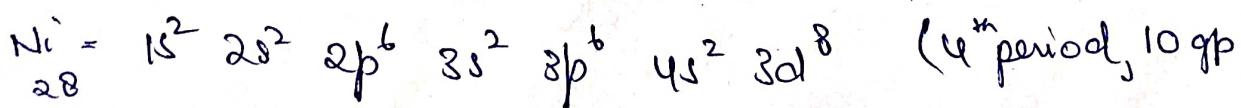
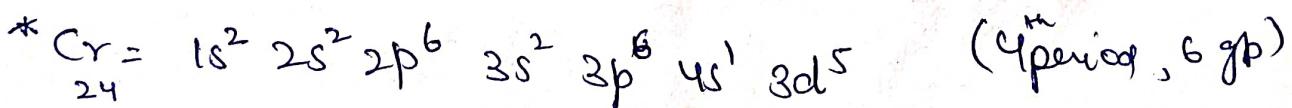
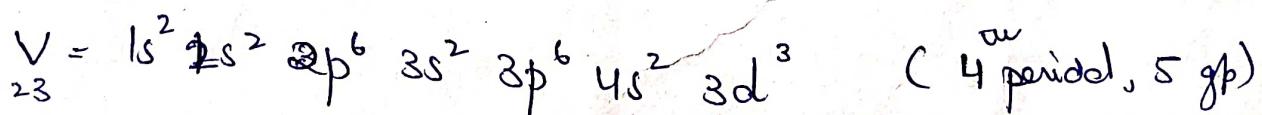
→ Elements belongs to 3-12 group

→ d-block → all metals

ultimate shell

penultimate shell

Example - ^(Transition) 12 Group element → not transition



F-block element

F block elements are Group 3 element

F' block elements

Lanthanoids (14 elements)
(4f)

Actinoids (14 elements)
(5f)

Lanthanoids \rightarrow 6 period

Actinoids \rightarrow 7 period (Radioactive)

electronic conf" \rightarrow determined by experimental calculation

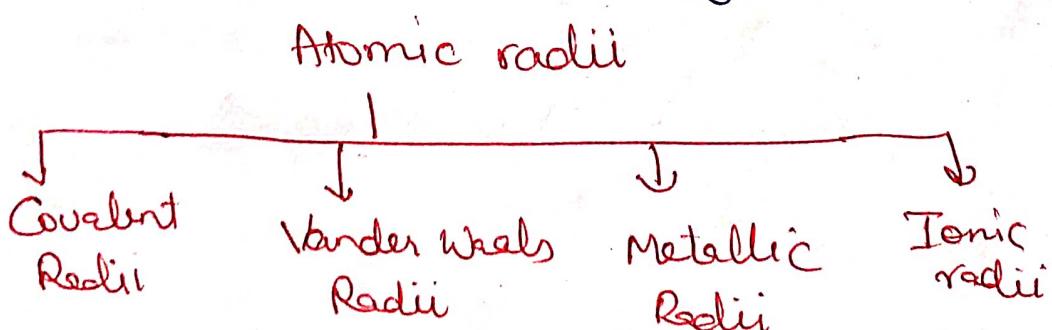
$$\boxed{n s^{1-2} (n-1) d^{0-2} (n-2) f^{1-14}}$$

↓
anti penultimate shell

Atomic Radii

- Physical properties like density & depends on the mpt & bpt size of atom.
- As it is not possible to isolate an individual atom, so to measure the atomic radii is not possible.
- Traditional defⁿ of atomic radii is the distance of valence electrons to the nucleus.
 - ↓
 - But e⁻ is not particle, it behave as wave also
 - ↓
 - So, the path of e⁻ is not defined or boundary of an atom is not defined.
- ∴ Atomic radii can be determined indirectly by measuring the distance b/w the nuclei of two bonded atoms in a gaseous molecule or b/w the nuclei of two ions in a crystal.
 - ↓
 - i.e. in bonded form

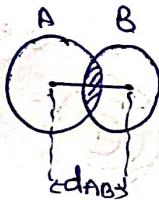
So, atomic radii is classified as:



Covalent Radii

↓
one half of the distance b/w the two nuclei of two atoms bonded through a single covalent bond

↓
Covalent bond formed by sharing of e-



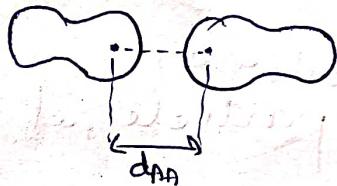
$$r_{AB} = d_{AB}/2$$

ORDER of RADII

Vander Waal's Radii

Molecules held together by Vander waal's forces

↓
half the distance b/w 2 atoms within two molecules refined as
Vander Waal for radii

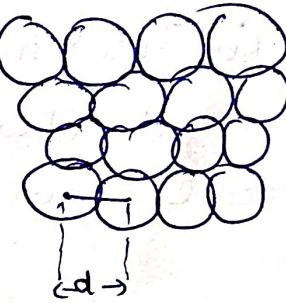


$$r_{AA} = d_{AA}/2$$

Found in Nobel gases

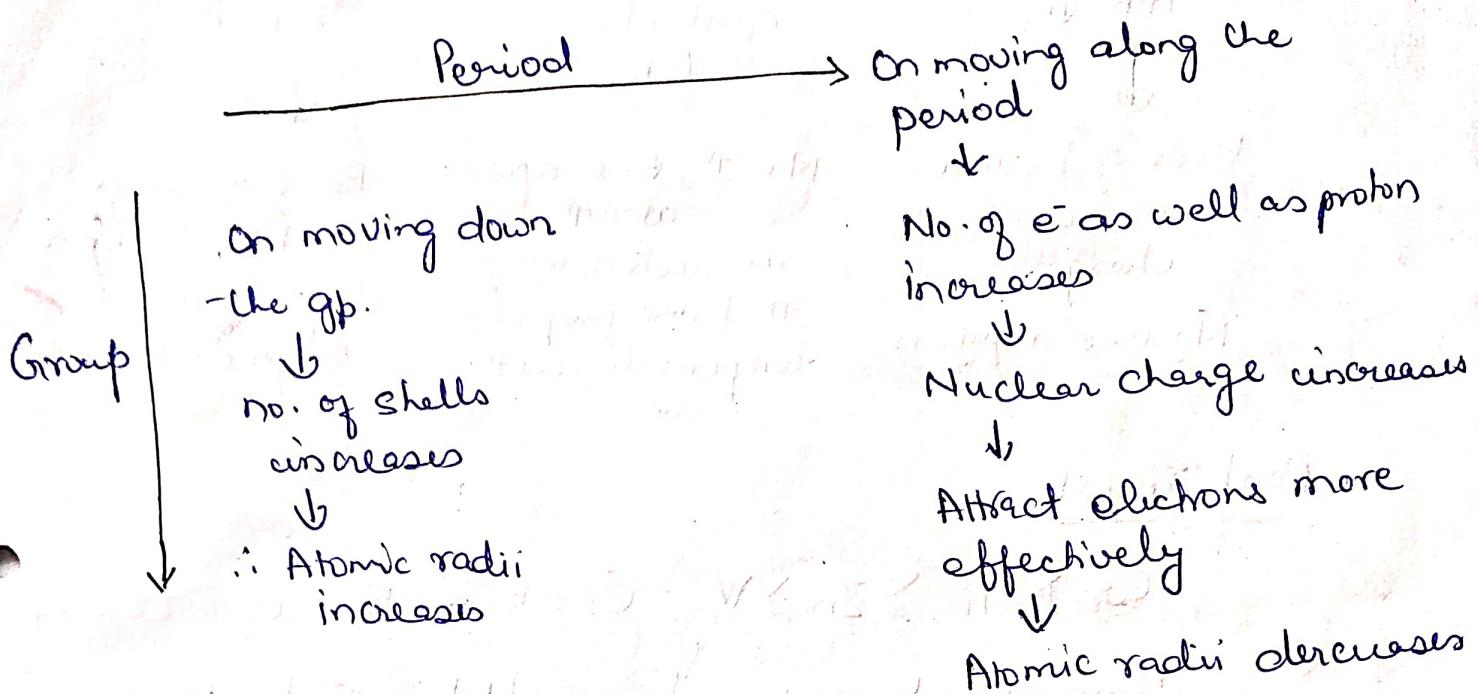
Metallic Radii

↓
One half the distance b/w two nuclei in two adjacent metal atoms.



Vander Waal's radii \rightarrow Metallic radii \rightarrow Covalent Radii

Trends or Periodic Variations of Atomic Radii



Exception :

Along period

Li Be B C N O F Ne

expected trend \rightarrow Li > Be > B > C > N > O > F > Ne

Reality \rightarrow ① $\underbrace{\text{Ne} > \text{Li} > \text{Be} > \text{B} > \text{C} > \text{N} > \text{O} > \text{F}}$ covalent radii

Vander waal's radius Ne radii expected to be least

but in actual the atomic radii of Ne is largest bcos of Vanderwaal radius

Similar trend will be found in -

② Na Mg Al Si P S Cl Ar

Ar > Na > Mg > Al > Si > P > S > Cl

③ 3d elements -

Sc	Ti	V	Cr	Mn	Fe	Ni	Cu	Zn
A.R ↓			constant A.R			A.R ↑		
↓			↓			↓		

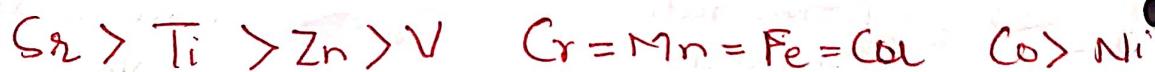
Bcos of increase
in nuclear
charge

No e-e repulsion

N.C ↑, E-e repulsion
also ↑
in such a way
that one property
compensate another
one.

Bcos of large
electro-electron
repulsion as
compared to N.C

Real Trend



Same trend is found in 4d & 5d elements

④

3d	Sc _{a1}	Ti ₂₂	V ₂₃
4d	Y ₃₉	Zr ₄₀	Nb ₄₁
5d	La ₅₇	Hf ₇₂	Ta ₇₃

In case of Sc

Y
La

A.R increases

same trend

i.e. Sc < Y < La

But Ti₂₂ < Zr₄₀ < V₂₃

Zr₄₀ & Nb

Hf₇₂ & Ta

Real trend is Ti < Zr ≈ Hf due to
lanthanide contraction

V < Nb ≈ Ta

due to

actinide contraction

Both contraction

↓

Bcos of less shielding effect

of f-orbital

↓

Outer e- will be attracted
more towards nucleus

↓

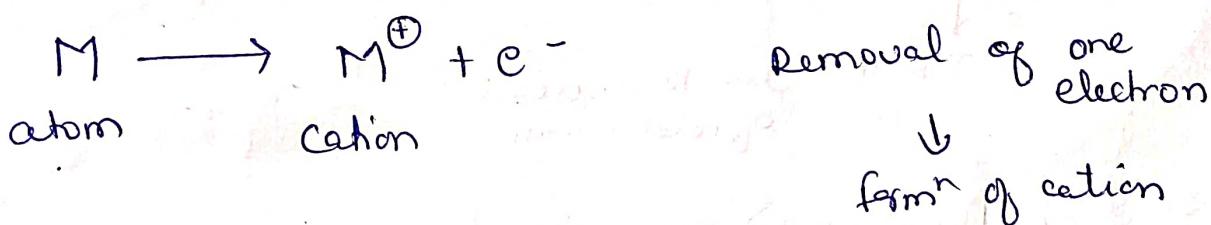
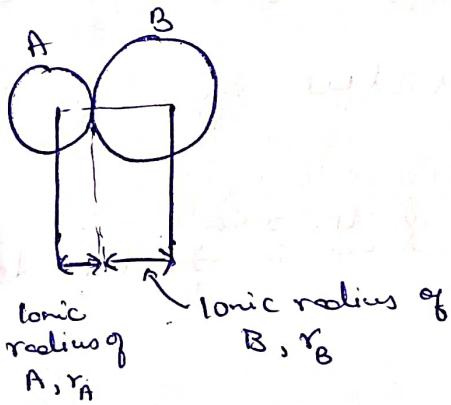
As a result the atomic radii of Ta ≈ Nb & Hf ≈ Zr

order of Screening effect

s > p > d > f

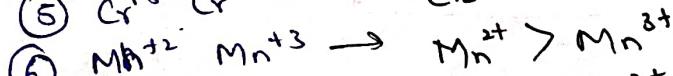
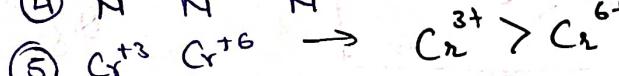
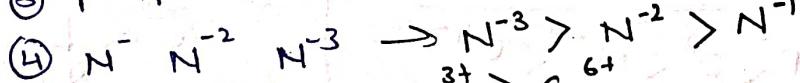
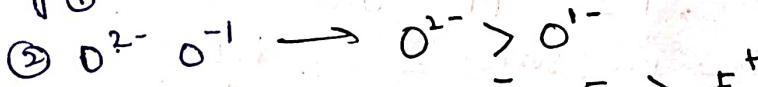
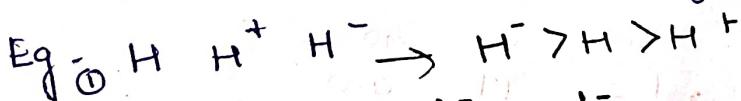
due to their shape

Ionic Radii

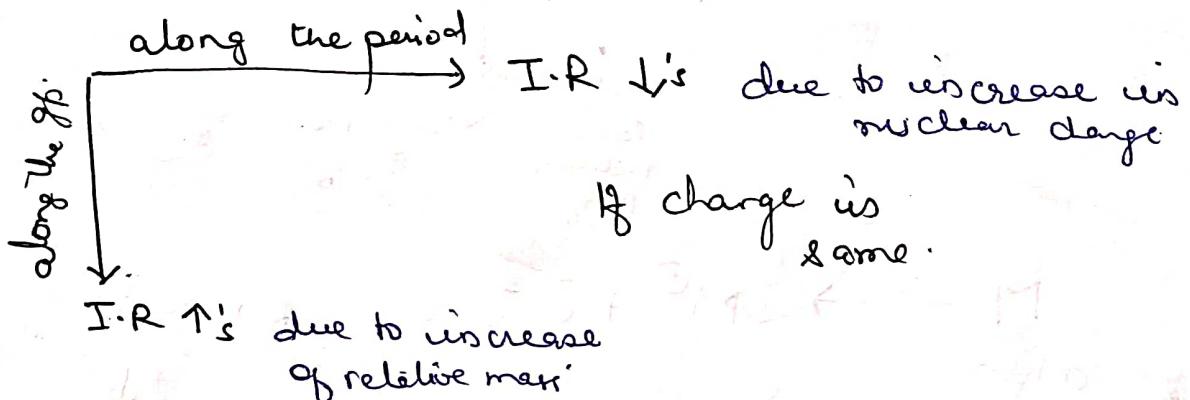


Anion > Atom > Cation

Atom	Anion	Cation
no. of e^-	more e^-	less e^-
no. of Proton	no. of proton same	no. of proton same
		↓
	Effect of Nuclear charge decreases on valence e^- /more e^-e^- repulsion	Bcos of more N.C as compared to e^-
	↓	↓
	Size increases	Size decreases



Trends in value of Ionic Radii Same trend as Atomic radii



Size of Isoelectronic ions:

Isoelectronic species / ions → species / ions with same number of electrons

for such species

$$\text{size} \propto \frac{1}{\text{Atomic no.}}$$

or Higher the value of Z/e or $\frac{\text{Atomic no.}}{\text{no. of e}^-}$

i.e., for a series of ions in same arrangement of electrons smaller is the size

size of ion decreases as Atomic no. increases



Nuclear charge increases

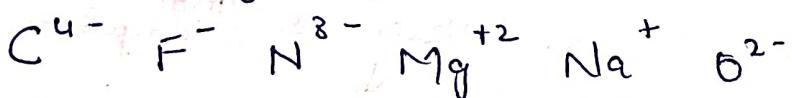


attract e^- more strongly



therefore size decreases

Example:



AN	6	9	7	12	11	8
----	---	---	---	----	----	---

no. of electrons	10	10	10	10	10	10
------------------	----	----	----	----	----	----

Z/e	0.6	0.9	0.7	1.2	1.1	0.0
-------	-----	-----	-----	-----	-----	-----

∴ Isoelectronic series & size $\propto \frac{1}{\text{At. no.}}$



or $Z/e \propto \frac{1}{\text{size}}$

③ Penetration power of Valence electrons

→ Effect or power to remove e^- from s, p, d or f subshell of same energy level.

∴ In shell or the penetration power of the electrons in same principal energy shell is

$$s > p > d > f$$



ns e^- 's are more firmly held by the nucleus than np e^- 's and so on.



I.E reqd to remove e^- from s-orbital is more as compared to other orbitals

∴ Order of I.E reqd

$$ns > np > nd > nf$$

④ Stable Electronic configuration: * Check for different atoms

As we know half filled / full filled orbitals are more stable than partially filled.

So difficult to remove electrons from these orbitals as compared to partially filled orbitals

Ionization energy will be higher as expected

E.g Period II Be (half filled) and Ne (full filled)

Period III

Mg (half filled) and P (full filled)

Less stability
as compared to
fully filled

but more I.E reqd
as compared to other
elements of same period as expected

Extra stability

Higher I.E

Noble gases

He,
Ne
Ar
Kr
Xe
Rn

} → octet configⁿ → highest I.E

Li⁺
Na⁺
K⁺
Rb⁺

} also requires more I.E, bcos they have noble gas config ($1s^2 p^6$)

So, requires greater energy to pull out an electron

⑤

Shielding effect of inner electrons:

In poly electrons atom, the valence electrons are attracted by nucleus and repelled by inner electrons.

As a result, valence e⁻'s / outermost e⁻'s experiences less attraction from the nucleus

↓
K/d as Screening effect / shielding effect.

→ Within a given shell, the screening effect of inner electrons decreases due to their shapes and follows the order $s > p > d > f$

More no. of inner electrons

↓
Act as shield / Screen b/w nucleus & outermost e⁻
lower I.E

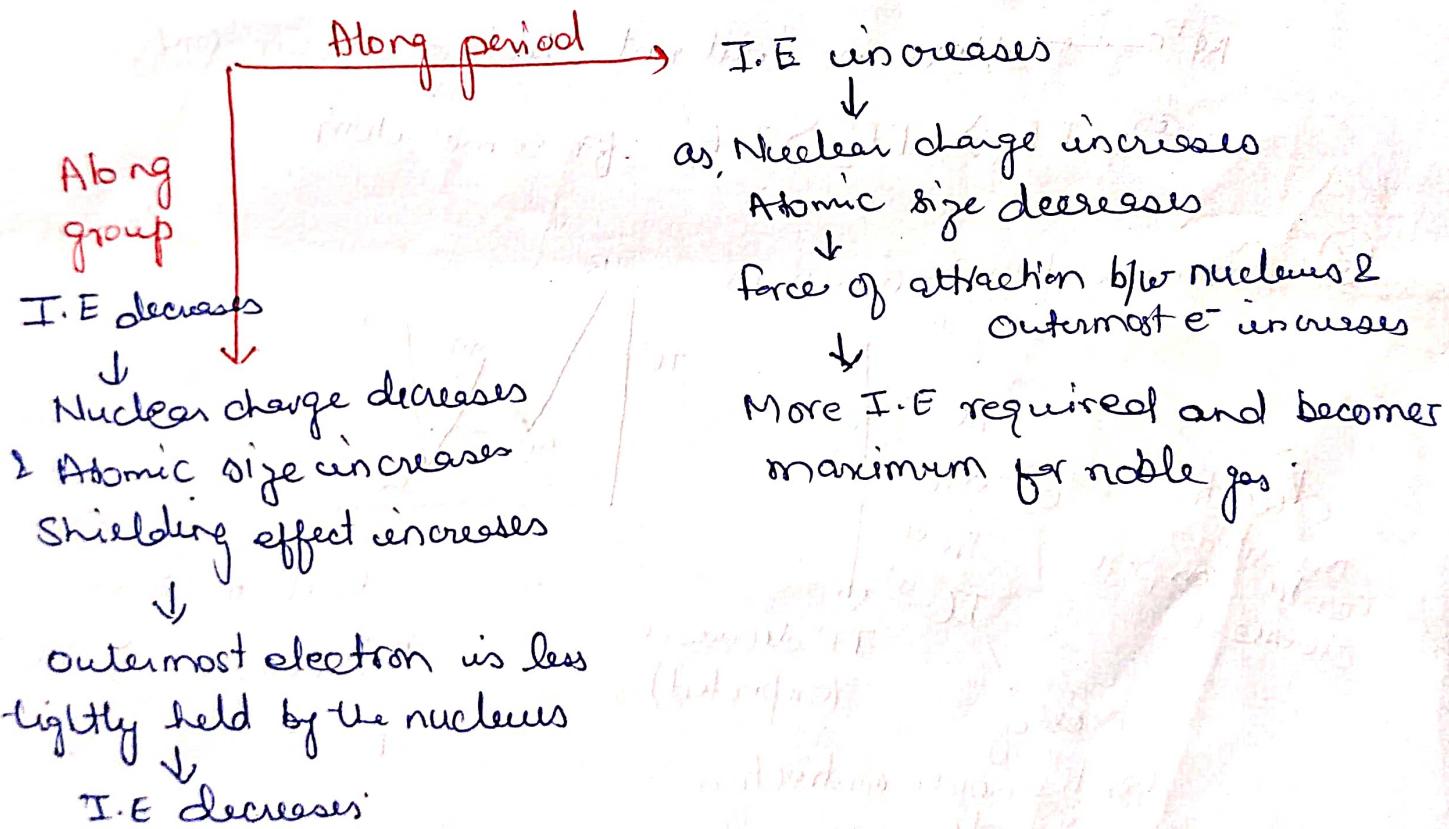
⑥ Principal Quantum no.

Distance of outermost electron from the nucleus is determined by principle quantum number (n)

As n increases
 ↓

Distance increases → Nuclear charge on outermost e⁻ decreases
 ↓
 I.E decreases

Trends of Ionization energy in periodic table



Exceptions: I.E along a period is not smooth, some irregularities are distinctly observed.

Expected trend

$\text{Li} < \underline{\text{Be}} < \underline{\text{B}} < \underline{\text{C}} < \underline{\text{N}} < \underline{\text{O}} < \underline{\text{F}} < \text{Ne}$ → I.E

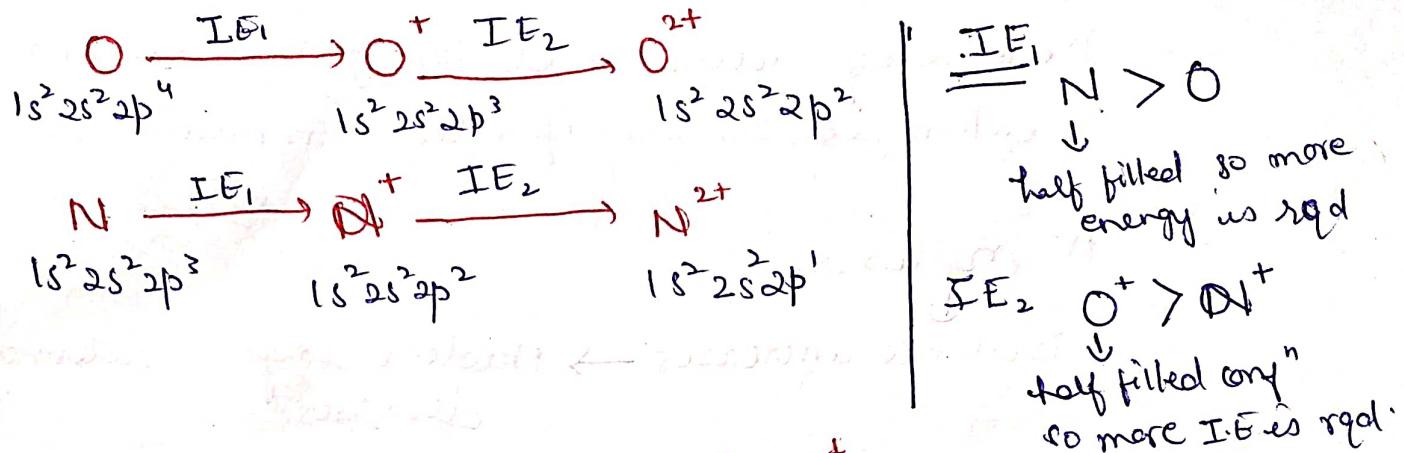
Real trend

$\text{Li} < \underline{\text{B}} < \underline{\text{Be}} < \underline{\text{C}} < \underline{\text{O}} < \underline{\text{N}} < \underline{\text{F}} < \text{Ne}$

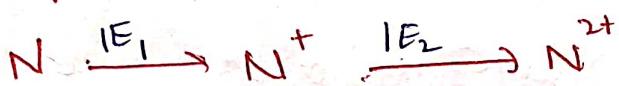
$1s^2 2s^2 2p^1$
 ↓ partially filled
 1s², 2s²

$1s^2 2s^2 2p^4$
 ↓ partially filled
 1s² 2s² 2p³

① Compare IE_1 , $2IE_2$ of O & N



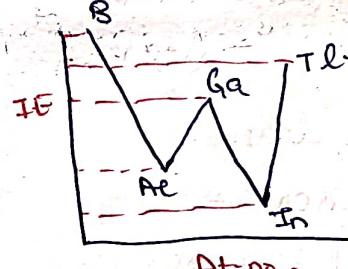
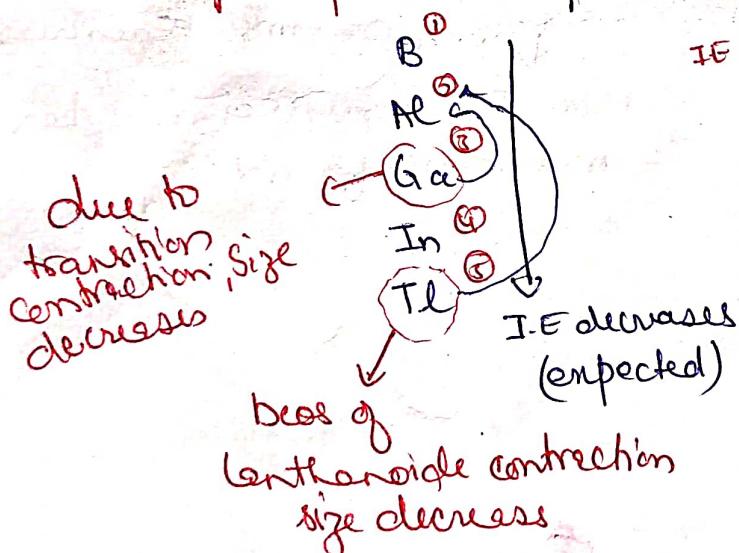
② Compare IE_1 , $2IE_2$ of N & N^+



~~here we will not check electronic confⁿ~~

as $IE_3 > IE_2 > IE_1$ for same atom

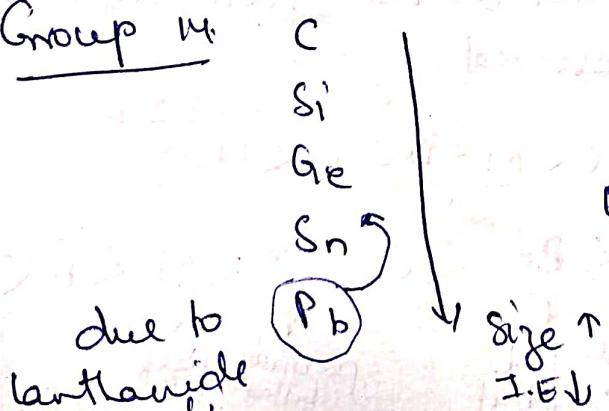
③ Group exception Group 13



Real order \rightarrow

$B > Tl > Ga > Al > In$

Group 14



Exact order

$C > Si > Ge > Pb > Sn$

$IE \downarrow \rightarrow e^-$ removal $\xrightarrow{\text{easy}}$ Metal

$IE \uparrow \rightarrow$ difficult e^- removal \rightarrow Non metals

→ When an e^- is removed from a neutral atom



Repulsion among the remaining electrons decreases
and nuclear charge increases

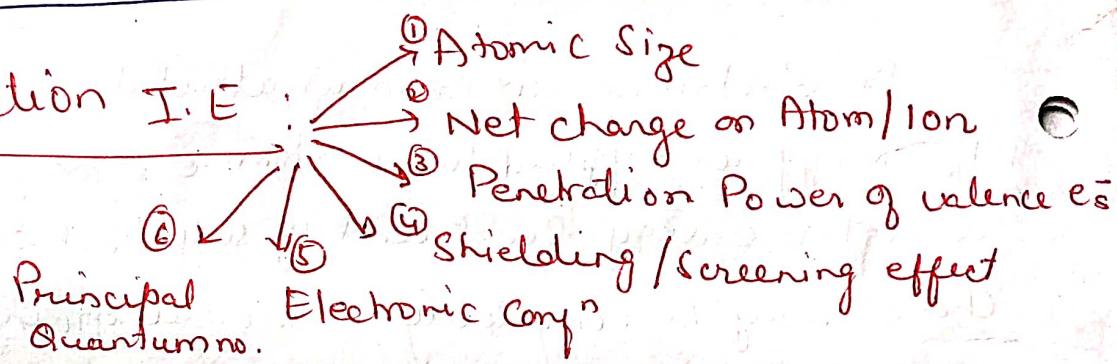


Ionization enthalpy increases

$$I_1 < I_2 < I_3 < I_4 \dots I_n$$

* for same atom.

factors affecting I.E.:



① Atomic Size: Larger the size of atom

↓
less will be attraction b/w valence e^- 's and nucleus

↓
 e^- will be removed easily

↓
less Energy reqd / less I.E.

② Net charge on atom/ion

As net charge on the atom/ion increases



Attraction effect b/w nucleus & outermost e^- increases



Ionization energy also increases

Eg: $Na^+ \rightarrow$ difficult to remove one electron from Na^+ , as now remaining electrons are more lightly held by the nuclear attraction.

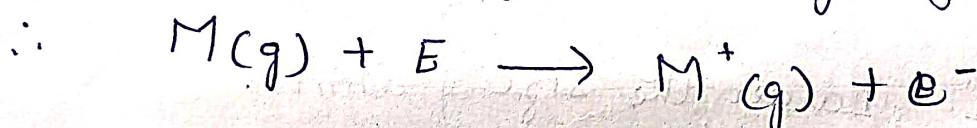
Another example of iso-electronic species

	S^{2-}	P^{3-}	Cl^-	
At.no -	16	15	17	order of size
e^-	18	18	18	
z/e	$16/18$	$15/18$	$17/18$	

$$Cl^- < S^{2-} < P^{3-}$$

IONIZATION ENERGY / IONIZATION ENTHALPY

An electron in an atom is attracted by +ve nucleus. If energy is supplied to an atom, e^- 's may be promoted to a higher energy level and if sufficient energy is supplied the e^- may be completely removed giving a +ve ion.



Minimum amt of energy ($\text{in } \text{kJ mol}^{-1}$) reqd to remove an electron from an isolated gaseous atom is called Ionization energy.

→ Magnitude of Ionization energy is a measure how tightly an e^- is held in the atom.

Higher I.E. → more difficult to remove the e^- .

→ Energy for removal of 1st e^- → First I.E. (I_1)

" reqd to remove 2nd e^- → Second I.E. (I_2)

" reqd to remove 3rd e^- → Third I.E. (I_3)

" reqd to remove n^{th} e^- → n^{th} I.E. (I_n)

→ When an e^- is removed from a neutral atom



Repulsion among the remaining electrons decreases
and nuclear charge increases

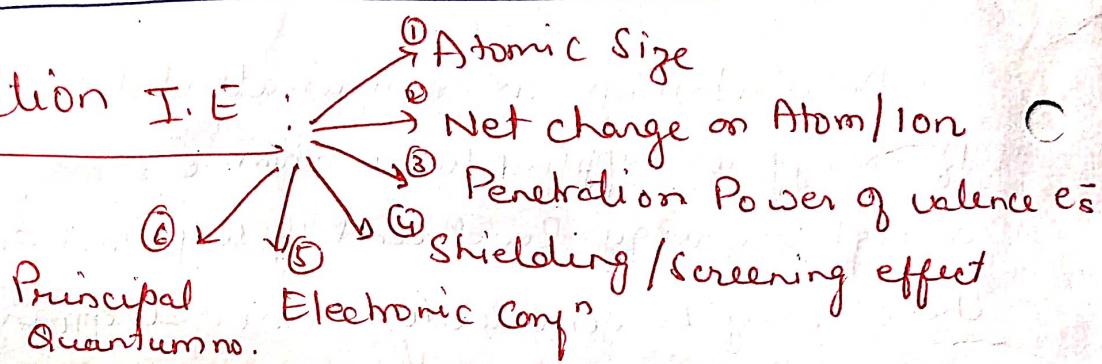


Ionization enthalpy increases

$$I_1 < I_2 < I_3 < I_4 \dots I_n$$

* for same atom.

factors affecting I.E:



① Atomic Size:

Larger the size of atom



less will be attraction b/w valence e^- 's and nucleus



e^- will be removed easily



less Energy reqd / less I.E

② Net charge on atom/ion:

As net charge on the atom/ion increases



Attraction b/w nucleus & outermost e^- increases

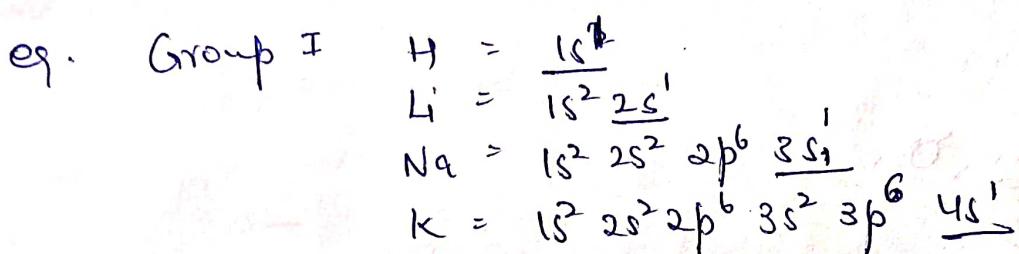


Ionization energy also increases

Eg: $Na^+ \rightarrow$ difficult to remove one electron from Na^+ as now remaining electrons are more lightly held by the nuclear attraction.

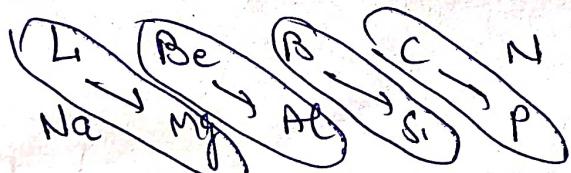
Periodicity

- ⇒ ~~Repet~~ Regular / gradual increase or decrease after ~~in~~ physical or chemical properties after fixed interval in periodic table
- ⇒ Cause of periodicity → similar electronic config.



- ⇒ Typical elements Na Mg Al Si P

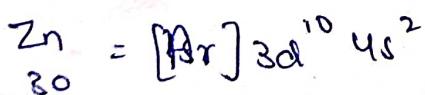
⇒ Diagonal relationship



→ Diagonal elements show similar properties

- ⇒ Transition elements (from group 3-11)

Group no. ↓
elements have vacant d orbital.



No vacant d orbital
Not transition metal

due to charge
same size
(charge to size ratio)

or
same charge density.

- ⇒ All transition element are d-block elements
but all d block elements are not transition elements.

→ E.A of half filled orbitals → Be, Mg, Al & P

{ Be
Mg } fully filled
Ca S orbitals
N - half filled p orbital

Fully filled / half filled orbitals → stable

← Conf

less tendency to add extra e-
↓
Hence E.A is very low or zero.

→ E.A of metals < E.A of non-metals
low high.

→ E.A decreases from top to bottom. In case of F, Cl, Br, I
The E.A of F is unexpected low due to smaller size
of F & consequently greater e-e repulsion.

→ Energy released to add 1st e- → first EA (E_1)

→ Energy released to add 2nd e- → second EA (E_2)

→ Energy released to add 3rd e- → third EA (E_3)

Addition of 1st electron → atom becomes -ve charged M⁻

↓ Add¹ of 2e⁻ to M⁻
opposed by columbic force of repulsion

↓
∴ Energy required to add 2nd e- E_2 is always -ve

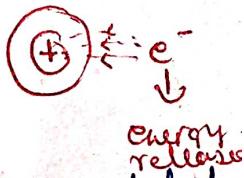
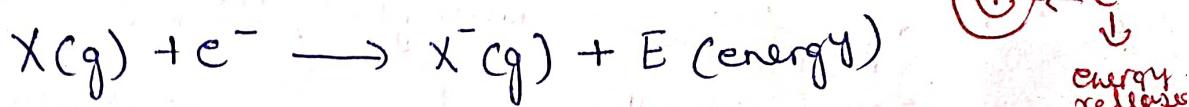
→ If atom has a spontaneous tendency to add/gain an extra electron

↓
E.A is said to be positive

If an atom is reluctant to gain an e- or -ve tendency to gain an e-

↓
negative E.A

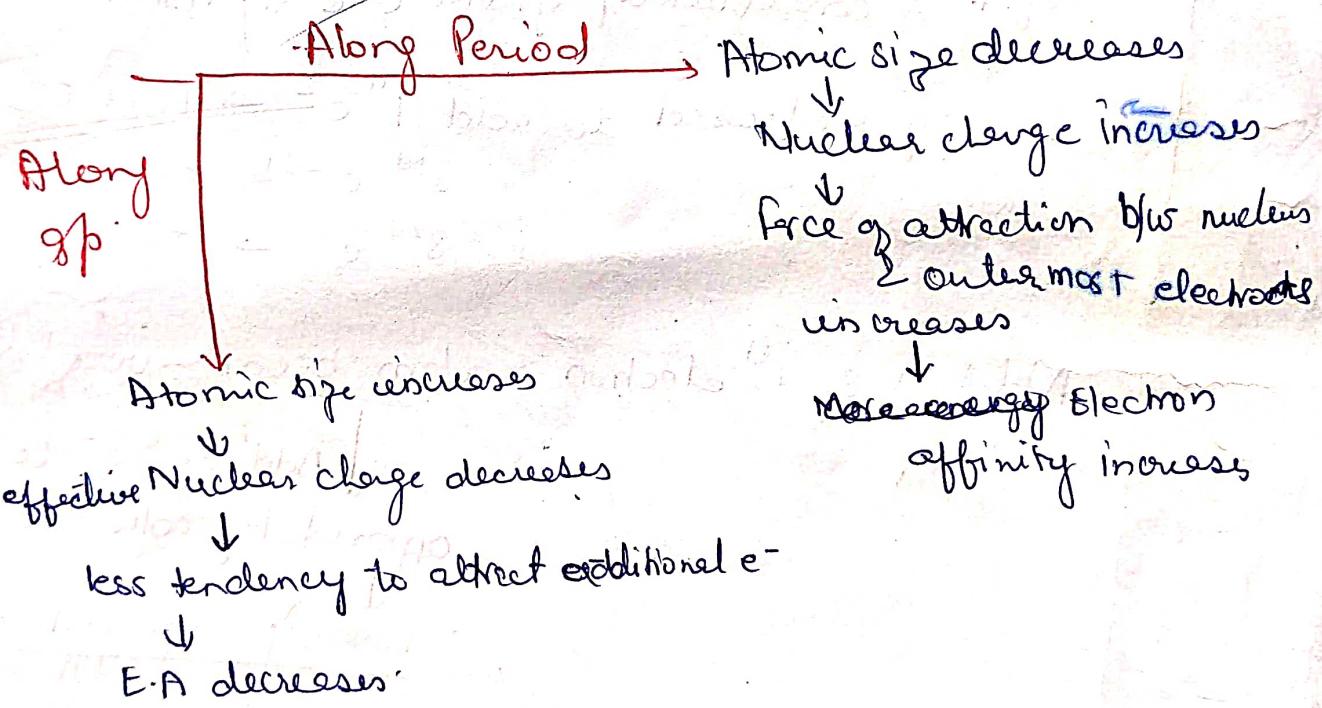
Electron gain affinity: measures the tendency of an atom (to gain ^{an} electron) to change into anion



- Amt of energy released, when an electron is added to a gaseous isolated atom or ion.
- Greater the energy released \rightarrow greater is E.A. in process of taking up extra electron.

→ Factors affecting \rightarrow atomic size, Effective nuclear charge, stable electronic configⁿ.

Variation / Trends of E.A. in periodic table



→ E.A. for noble gases \rightarrow is positive, bcos of stable electronic config.

↓

less tendency to add ^{an} extra electron

→ E.A. of halogens: is very high bcos of config $\sim s^2 p^5$

↓

Have strong tendency to add ^{an} extra electron to complete their octet.

Oxygen - $1s^2 2s^2 2p^4$

EA_1 ($1s + EA$) is positive, bcs energy is released when a mole of O atom get converted to O^- ions

↓
But (EA_2) IInd E.A is -ve, bcs energy has to be supplied to convert 1 mole of O^- ions to O^{+2} ions

Similarly for others also EA_1 is positive & EA_2 is -ve.

Exceptions

