★ Periodicity and cause of Periodicity

The properties of elements are repeated after regular interval. This is called periodicity.

Reason for Periodicity - Valance shell electronic configuration in a group is same.

★ Periodic Properties

The properties which are repeated at regular intervals are called Periodic Properties. These properties mainly include:-

1. Atomic and ionic radii
2. Ionisation enthalpy.
3. Electron Gain Enthalpy
4. Electronegativity
5. Valency
6. nature of oxides, hydrides, oxoacids etc.

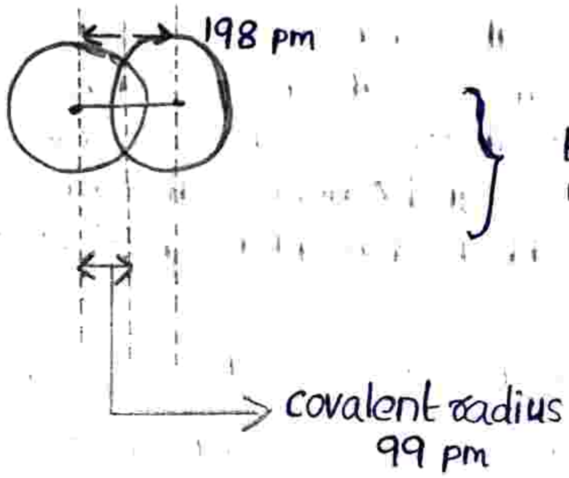
1. • Atomic Radius

It is the distance from the centre of the nucleus to the outermost shell of electrons. [imaginary definition]

The radius of a non-metal is expressed as covalent radius. And that of a metal as metallic radius.

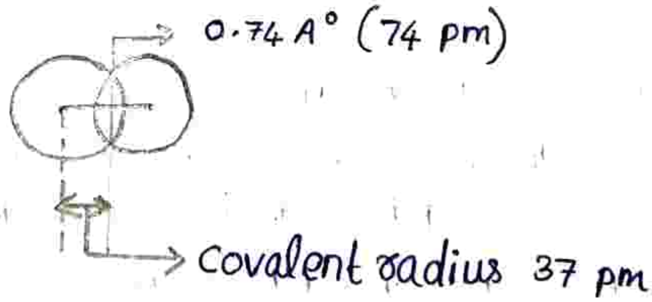
Covalent radius - half of the internuclear distance between two identical atoms bonded by a single covalent bond.

eg ① :-

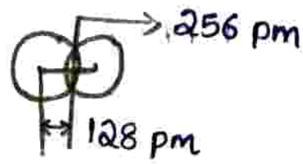


} Bond distance in Cl_2 molecule

eg ② :- Bond distance in H_2 molecule



Metallic radius - Half of the internuclear distance between two adjacent atoms in a metallic crystals.



[radius of copper]

The order of different atomic radius is

vanderwaal's radius > metallic radius > covalent radius

Variation of Atomic Size

1. In a period - on moving from left to right in a period atomic size decreases. This is due to ↑ in the effective nuclear charge (and the added electrons go to same principal shell.) .

eg :	Li $z-3$	Be $z-4$	B $z-5$	C $z-6$	N $z-7$	O $z-8$
Config^n :	2,1	2,2	2,3	2,4	2,5	2,6
Radius :	1.34 A°	0.90	0.82	0.77	0.75	0.73

2. In a group - on moving down the group the number of Principal shells around the nucleus ↑ses. and the force of attraction between nucleus and outermost electron get decreases.

- Ionic radius

The distance between the centre of the nucleus of an ion and its orbit.

cation is $\underset{2,8,1}{Na} \longrightarrow \underset{2,8}{Na^+} + e^-$ [formation of cation] formed by lose of e^-.

Here the number of e^- decreases while the nuclear charge remains the same. Hence the effective nuclear charge ↑ses and the size of cation is less than that of parent atom.

Q. Which is larger ? why

(a) Na or Na^+

(b) K or K^+

(c) Mg or Mg^+

Anion is formed by gaining of e^- to a neutral atom.

$$\underset{2,8,7}{Cl} + e^- \longrightarrow \underset{2,8,8}{Cl^-}$$

Here the effective nuclear charge decreases and the electrostatic attraction exerted by the nucleus is spread over a large number of e^-s. so the anions become larger than the neutral atom.

Q. Which is larger ?

ⓐ Cl or Cl^-

ⓑ Br or Br^-

Iso electronic species

Atoms or ions containing same number of e^-s.

eg :- N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+}, Al^{3+}

Q. Arrange the above ions in the increasing order of atomic size.

Ans)

	Al^{3+} <	Mg^{2+} <	Na^+ <	F^-	< O^{2-}	< N^{3+}
no. of e^-s	10	10	10	10	10	10
no. of protons	13	12	11	9	8	7

On ↑ing the effective nuclear charge atomic size decreases. Thus Al^{3+} has least atomic size among the ions given.

2. Ionization Enthalpy

It is defined as the minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom. Unit of ionization Enthalpy is $KJ\ mol^{-1}$.

$$X_{(g)} \longrightarrow X^{+}_{(g)} + e^{-} \qquad \Delta_i H = x\ KJ\ mol^{-1}$$

Second ionization enthalpy is the energy required to remove the second most loosely bound electron.

$$X^{+}_{(g)} \longrightarrow X^{2+}_{(g)} + e^{-} \qquad \Delta_i H_2$$

$$X^{2+}_{(g)} \longrightarrow X^{3+}_{(g)} + e^{-} \qquad \Delta_i H_3 \ldots$$

Successive Ionization Enthalpy follows the order.

$$\Delta_i H_3 > \Delta_i H_2 > \Delta_i H_1$$

Factors affecting Ionization Enthalpy

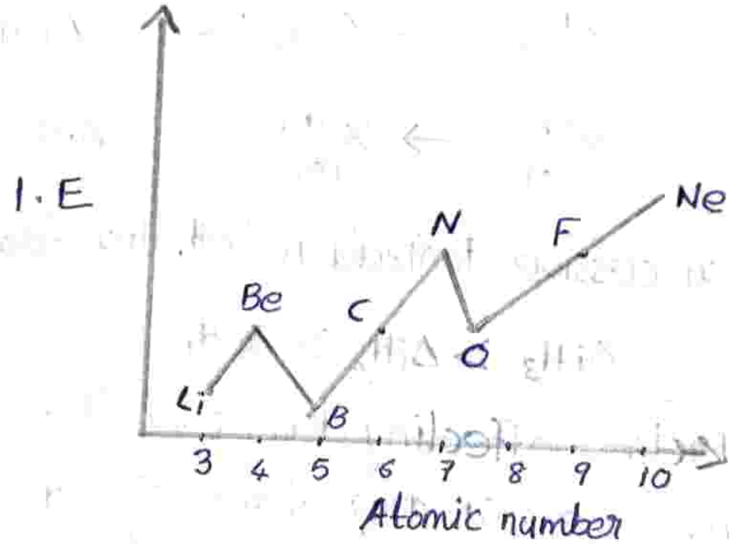
1. Atomic size - On ↑ing atomic size the force of attraction between nucleus and outermost e^- decreases and I.E decreases.
2. Nuclear charge - On increasing the nuclear charge I.E ↑ses. [Reason: the force of attraction between nucleus and e^- ↑ses.]
3. Screening effect or shielding effect - The phenomenon which occurs when the nucleus reduces its force of attraction on the valence electrons due to the presence of electrons in the inner-shell. As the no. of inner electron ↑ ses, shielding effect ↑ ses, and I.E decreases.
4. Electronic configuration - Half filled and completely filled orbitals acquire extra stability ($s^2, p^3, p^6, d^5, d^{10}, f^7, f^{14}$..) and hence it is difficult to remove electrons from these stable configuration.

5. Penetration effect - For a particular energy ~~tel~~ level the order of energy required to remove an electron from different subshell is $s > p > d > f$

Variation of ionization enthalpy in a period

In a period on moving from left to right I.E increases.
This is due to :- 1. Increase in the effective nuclear charge.
2. decrease in atomic size.

Variation of I.E of 2nd period elements



Observations

I. I.E. of Be is more than that of B.
This is due to :- 1. stable configuration of Be ($1s^2 2s^2$) compared to Boron ($1s^2 2s^2 2p^1$)
2. Penetration effect (energy required to remove 2s electron is more that that of 2p electron.)

II. I.E of Nitrogen is more than that of oxygen.
This is because N has stable half filled electronic configuration.

$N (z-7) - 1s^2 2s^2 2p^3$

$O (z-8) - 1s^2 2s^2 2p^4$

Variation of I.E in a group

on moving down the group I.E decreases. This is due to increase in atomic size. on increasing atomic size distance between nucleus and outermost electron increases. moreover in multielectron atoms screening effect also observed, which leads to decrease in I.E

3. Electron gain Enthalpy ($\Delta_{eg}H$)

It is the amount of energy released when an isolated gaseous atom gains an extra electron to form a negatively charged ion in the gaseous state.

$$X_{(g)} + e^- \longrightarrow X^-_{(g)} \quad \Delta_{eg}H = x \text{ KJmol}^{-1}$$

$$Cl_{(g)} + e^- \longrightarrow Cl^-_{(g)} \quad \Delta_{eg}H = -340 \text{ KJmol}^{-1}$$

Factors affecting $\Delta_{eg}H$

1. Atomic size - electron gain enthalpy become less -ve with increase in atomic size.

2. Nuclear charge - on increasing the nuclear charge electron gain enthalpy become more negative.

3. Electronic configuration - stable configuration (half filled or completely filled) do not have any charge to take up extra electron. In such case they possess +ve electron gain enthalpy.

   eg:- Noble gases have positive electron gain enthalpy. Here electron has to enter next higher principal quantum number leading to a very unstable configuration.

Note: If e^- is added to an anion $\Delta_{eg}H$ is +ve, because energy is needed to overcome the repulsion between anion and the electron. eg:- $O_{(g)} + e^- \longrightarrow O^-_{(g)} \quad \Delta_{eg}H \; -141 \text{ KJmol}^{-1}$

$$O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)} \quad \Delta_{eg}H \; +780 \text{ KJmol}^{-1}$$

Variation of $\Delta_{eg}H$ along a period and also in a group

on moving from left to right in a period electron gain enthalpy become more negative

exceptions -

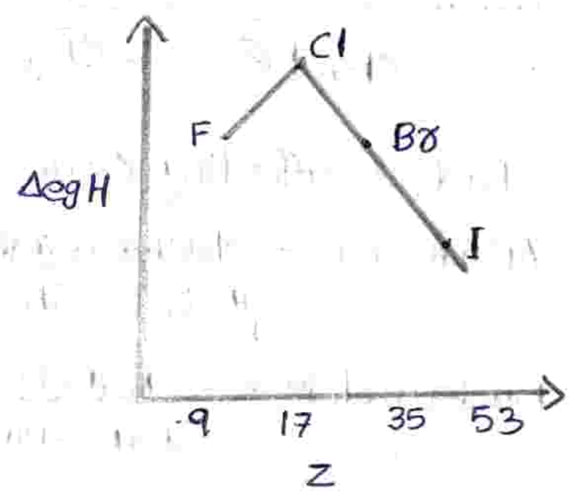
① Be, N both have +ve electron gain enthalpy due to stable configuration.

② Noble gases have positive electron gain enthalpy.

Electron gain enthalpy become less negative on moving down the group.

Variation of $\Delta_{eg}H$ of Halogens

Cl > F > Br > I



Electron gain enthalpy of fluorine is lower than that of chlorine. In the case of F the added electron goes to the smaller n = 2, quantum level and suffers significant repulsion from the other electrons present in this level. [or due to small size of fluorine atom]. No such repulsion in Cl atom due to its large size compared to F atom.

Q. Electron gain enthalpy of sulphur is more than that of Oxygen. why?

Ans) In oxygen added e^- goes to the smaller n = 2 quantum level and suffers significant repulsion from the other e^-s present in this level. In sulphur atom the added e^- goes n = 3 quantum level. Here no such repulsion arises.

4. ## Electronegativity

The ability of an atom to attract the shared pair of electron to itself is called electronegativity. Linus Pauling scale is widely used to measure electronegativity. Most electronegative element is fluorine. Second most electronegative element is oxygen.

5. **Valency** :- Valency is the combining capacity of an element. i.e number of e^-s lost, gained or shared by an atom during chemical reaction is its valency.

Sl. No	Compound	constituent elements	No. of e^-s shared	Valency
1	MgO	Mg - 2,8,2 O - 2,6	2 2	2 2
2	HF	H - 1 F - 2,7	1 1	1 1
3	Na_2O	Na - 2,8,1 O - 2,6	1 2	1 2
4	OF_2	O - 2,6 F - 2,7	2 1	2 1

Anomalous Properties of second period elements

The first element of each group differ from rest of the elements of the same group due to.

1) Small in size
2) High electronegativity and I.E
3) Absence of d orbitals

Diagonal relationship

Diagonally placed elements in adjacent group shows resemblence in certain properties. This is called diagonal relation.

	Gr. 1	2	13	14
2nd period	Li	Be	B	C
3rd period	Na	Mg	Al	Si

Li shows resemble with Mg
Be " " " Al
B " " " Si

Periodic Trends in chemical reactivity

1. Metallic character - metallic character decreases from left to right in a period and increases from top to bottom in a group.

2. Reducing and Oxidising character :-

Reducing agent $\longrightarrow$ Provide e^-

The reducing power of element decreases across the period and increases down the group.

Oxidising agent $\longrightarrow$ accept e^-s

oxidising power of elements increases across the period and decreases down the group.

3. Nature of oxides -

Oxides of metal are basic
oxides of non-metal are acidic
semi metal oxides are ampholeric

Acidic nature of oxides increases from left to right in a period.

$\underbrace{Na_2O \quad MgO}_{\text{Basic}} \quad \underbrace{Al_2O_3 \quad SiO_2}_{\text{ampholeric}} , \quad \underbrace{P_2O_5 \quad SO_2 \quad Cl_2O_7}_{\text{Acidic}}$

Problems

1. 1.0 g of magnesium is burnt with 0.56 g O_2 in a closed vessel. Which reactant is left in excess and how much?

$$Mg + \frac{1}{2}O_2 \longrightarrow MgO$$

Ans) ie. 24 g Mg reacts with 16 g O_2

Thus 1.0 g Mg react with

$$\frac{16}{24} O_2 = 0.67 \text{ g } O_2$$

But only 0.56 g O_2 is available which is less than 0.67 g.

Thus O_2 is limiting reagent

16 g O_2 reacts with 24 g Mg

0.56 g O_2 will react with $\frac{24}{16} \times 0.56$

$= \underline{0.84 g}$

Amount of Mg left unreacted $= 1.00 - 0.84 = \underline{0.16 \text{ g}}$ Mg

2. The first and the second ionization enthalpies (in $KJ\,mol^{-1}$) and the electron gain enthalpy of a few elements are given below.

Elements	$\Delta_i H_1$	$\Delta_i H_2$	$\Delta_{eg} H$
I	520	7300	-60
II	419	3051	-48
III	1681	3374	-328
IV	1008	1846	-295
V	2372	5251	+48
VI	738	1451	-40

Which of the above elements is likely to be

a) the least reactive element
b) the most reactive metal
c) the most reactive non metal
d) the least reactive non metal
e) the metal which can form stable binary halides of the formula MX_2 (X = halogen)

f) The metal which can form a predominantly stable covalent halide of the formula MX (X = halogen).

Ans) a) The least reactive element

Noble gases are least reactive for them $\Delta_{eg}H$ is +ve. more over $\Delta_i H_1$ is high.

Answer is V.

b) The most reactive metal.

Answer is II. It has least $\Delta_i H_1$ (first ionization enthalpy) and low electron gain enthalpy ($\Delta_{eg}H$).

c) The most reactive non-metal.

Answer is III (definite it is F).

It can be understood from $\Delta_{eg}H$.

value [more -ve $\Delta_{eg}H$ & high $\Delta_i H_1$].

d) The least reactive non-metal.

IV (Iodine)

$\Delta_{eg}H$ less than that of F

[$\Delta_i H_1$ not so high comparitively high $\Delta_{eg}H$].

e) The metal which can form a stable binary halide of the formula MX_2 (X = halogen).

VI (Gr II elements)

The difference between $\Delta_i H_1$ & $\Delta_i H_2$ is less or $\Delta_i H_1$ is higher than alkali metals.

f) The metal which can form predominantly stable covalent halide (MX).

Answer is I (Alkali metals - Gr. 1)

low $\Delta_i H_1$

High $\Delta_i H_2$

Li - 2,1

Na - 2,8,1

In such case $\Delta_i H_1$ is very low

But $\Delta_i H_2$ is high due to stable configuration.