PERIODIC PROPERTIES

PERIODIC TABLE — GROUPS, PERIODS

PERIODIC TABLE — Ionisation energy, Electron affinity, Electro negativity, Radii, Lanthanide Contraction, Metallic Character

Old convention	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
New convention	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18

←S– Block→   ← p - Block →

1	H																	He
2	Li	Be				d - Block							B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	La·	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	Ac·															

✸ **Lanthanids ***
(4f-Series)

Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu

✸ **Actinids ** **
(5f-Series)

Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

←———————— f-Block ————————→

☞ Helium belongs to s-block because last entered electron goes in s-block.

☞ Iridium is the most dense element followed by Osmium.

☞ **NOTE :** **(1)** Last entered electron (According to Aufbau's Principle) decides the block of the element.

(2). The valence shell determines the period number.

GENERAL PROPERTIES OF PERIODIC TABLE

(1) There are Seventeen non-metals (including hydrogen) in periodic table.

(2) Five non - metals are solid **C, P, S, Se, I.**

(3) One non - metal is liquid i.e. **Br.**

(4) Eleven non-metals are gaseous.

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(5) Six gases are monoatomic (noble gases) i.e. **He, Ne, Ar, Kr, Xe, Rn.**

(6) Five gases are diatomic, they are **H, F, N, O, Cl**

(7) There are eight metalloids in periodic table like, **B, Si, Ge, As, Sb, Te, Po, At.**

(8) Five elements are liquid at room temperature namely **Cs, Fr, Ga, Hg, and Br,**

(9) s - Block and p - Block together are called Representative elements.

(1 0) Five elements are radioactive amongst representative elements. They are **Po, At, Rn, Fr and Ra**

(1 1) There are seven periods in long - form of periodic table.

1.DEVELOPEMENT OF MODERN PERIODIC TABLE

Dobereiner's Triads :

He arranged similar elements in the groups of three elements called as triads, in which the atomic mass of the central element was merely the arithmatic mean of atomic masses of other two elements or all the three elements possessed nearly the same atomic masses

Li	Na	K	
7	23	39	7 + 39/2 = 23
Fe	Co	Ni	
55.85	58.93	58.71	nearly same atomic masses

It was restricted to few elements, therefore, discarded

NEWLAND'S LAW OF OCTAVES :

He was the first to correlate the chemical properties of the elements with their atomic masses.
According to him if the elements are arranged in the order of their increasing atomic masses the eighth element starting from given one is similar in properties to the first one.
This arrangement of elements is called as Newland's law of Octave.

Li	Be	B	C	N	O	F
Ns	Mg	Al	Si	P	S	Cl
K	Ca					

This classification worked quite well for the ligher elements but it failed in case of heavier elements and, therefore, discarded
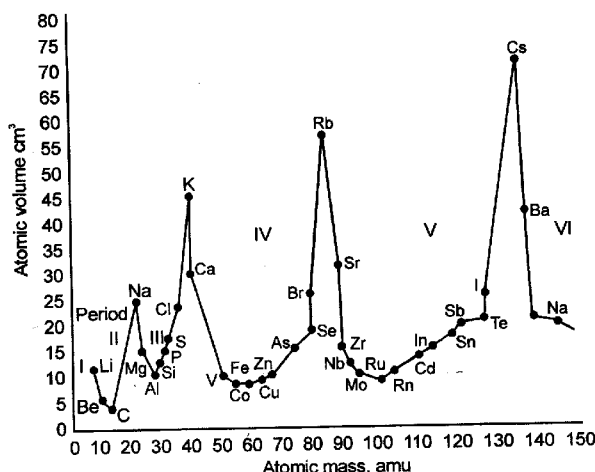
LOTHER MEYER'S CLASSIFICATION

He determined the atomic volumes by dividing atomic mass with its density in solid states.
He ploted a graph between atomic masses against their respective atomic volumes for a number of elements. He found the following observations.
(i) Elements with similar properties occupied similar positions on the curve.
(ii) Alkali metals having larger atomic volumes occupied the crests.
(iii) Transitions elements occupied the troughs.
(iv) The halogens occupied the ascending portions of the curve before the inert gases.
(v) Alkaline earth metals occupied the positions at about the mid points of the descending portions of the curve.
On the basis of these observations he concluded that the atomic volumes (a physical property) of the elements are a periodic function of their atomic masses.
It was discarded as it lacks practical utility.

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(d) Mendeleev's Periodic Table :

Mendeleev's Periodic's Law

According to him the physical and chemical properties ofthe elements are a periodic function oftheir atomic masses. He arranged then known elements in order of their increasing atomic masses considering the facts that elements with similar properties should fall in the same vertical columns and leaving out blank spaces where necessary.

Merits of Mendeleev's Periodic table :

(i) It has simplified and systematised the study of \3lements and their compounds

(ii) It has helped in predicting the discovery of new elements on the basis of the blank spaces given in its periodic table.

Mendeleevs predicted the properties of those missing elements from the known properties of the other elements in the same group. Eka - Aluminium and Eka-silicon names were given for gallium and germanium (not discovered at the time of mendeleevs). Properties predicted by Mendeleevs for these elements and those found experimentally were almost similar.

Proeprty	eka-aluminium (predicted)	gallium (found)	eka-silicon (predicted)	germanium (found)
Atomic Mass	68	70	72	72.6
Density / (g/cm^3)	5.9	5.94	5.5	5.36
Melting point (K)	Low	30.2	High	1231
Formula of oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
Formula of chloride	ECl_3	$GaCl_3$	ECl_4	$GeCl_4$

(iii) Atomic weights of elements were corrected. Atomic weight of Be was calculated to be $3 \times 4.5 = 13.5$ by considering its valency 3, was correctly calculated considering its valency 2 ($2 \times 4.5 = 9$)

Demerits In Mendeleev's Periodic Table:

(i) Position of hydrogen is uncertain .It has been placed in lAand VilA groups because of its resemblance with both the groups.

(ii) No separate positions were given to isotopes.

(iii) Anomalous position of lanthanides and actinides in periodic table.

(iv) Order of increaseing atomic weights is not strictly followed in the arrangment of elements in the periodic table. For e.g.-Ar(At.wt.39.94) is placed before K(39.08) and Te (127.6) is placed before I (126.9)

(v) Similar elements were placed in differents groups(CuIB and Hg IIB) and the elements with different properties were placed in same groups(alkali metals IA and coinage metals IB)

(vi) It didn't explained the cause of periodicity.

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LONG FORM OF THE PERIODIC TABLE OR MOSELEY'S PERIODIC TABLE

He studied (1909) the frequency of the X-ray produecd by the bombardment of a strong beam of electrons on 'metal target. He found that the square root of the frequency of X-rays ($\sqrt{v}$) is directly proportional to number of effective nuclear charge (z) of metal i.e. to atomic number and not to atomic mass of the atom of that metal.(as nuclear charge of metal atom is equal to atomic number)

i.e. ($\sqrt{v}$) = a (z - b)

Where 'a' is the proportionality constant and 'b' is a constant for all the lines in a given series of X-rays. Therefore, he, concluded that atomic number was a better fundamental property of an element than its atomic weight He suggested that the atomic number (z) instead of atomic weight should be basis of the classification of the elements.

Modern Periodic Law (Moseley's Periodic Law)

Physical and chemical properties of elements are the periodic functions of their atomic number.If the elements are arranged in order of their increasing atomic numper, after a regular interval ,element with similar properties are repeated.

Periodicity

The repetition of the properties of elements after regular intervals when the elements are arranged in the order of increasing atomic number is called periodicity.

Cause of Perlodlcty:

The periodic repetition of the properties of the elements is due to the recurrence of similar valence shell electronic configuration after certain regular intervals. For example, alkail metals have same electronic configuration ns^1, therefore, have similar properties.

The long form of periodic table is the contribution of Range, Werner, Bohr and Bury

This table is also referred to as Bohr's table since it follows Bohr's scheme of the arrangements of elements into four types based on electronic configuration of elements

The modern periodic table consits of horizontal rows (periods) and vertical column (groups)

Periods:

There are seven periods numbered as 1, 2,3,4,5,6 and 7.

(i)   Each period consists of a series of elements haVing same valence shell.

(ii)   Each period corresponds to a particular principal quantum number of the valence shell present in it.

(iii) Each period starts with an alkali metal having outermost electronic configuration ns^1.

(iv)   Each period ends with a noble gas with outermost electronic configuration ns^2np^6 except helium having outermost electronic configuration $1s^2$.

(v)   Each period starts with the filling of new energy level.

(vi)   The number of elements in each period is twice the number of atomic orbitals available in energy level that is being filled. To illustrate

☞   1st period shortest period having only two elements. Filling of electron takes place in the first energy shell, for which,

n = 1, ℓ. = 0 (s-subshell) and m = O.

Only one orbital (1s) is available and thus it contains only two elements.

☞   3rd period short period having only eight elements. Filling of electrons takes place in the third energy level. For which,

n = 3, ℓ. = 0, 1,2 and no. of orbitals m = 0,	3,	5
no. of orbitals	1	3	5
	(3s)	(3p)	(3d)
Total no. of orbitals	9		

But the energy of 3d orbitals are higher than 4s orbitals. Therefore, four orbitals (one 3s and three 3p orbitals) corresponding to n = 3 are filled before filling in 4s orbital (next energy elevel). Hence 3rd period contains eight elements not eighteen elements.

Groups:

There are eighteen groups numbered as 1,2,3,4,5, ... 13, 14, 15, 16, 17, 18.

Group consists of a series of elements having similar valence shell electronic configuration.

2. CLASSIFICATION OF THE ELEMENTS :

It is based on the type of orbitals which receives the differentiating electron (i.e., last electron).

s-block elements

When shells upto (n − 1) are completely filled and the last electron enters the s-orbital of the outermost (nth) shell, the elements of this class are called s-block elements.

- Group 1 & 2 elements constitute the s-block.
- General electronic configuration is [inert gas] ns^{1-2}
- s-block elements lie on the extreme left of the periodic table.
- This block includes metals.

GROUP – IA OR 1   ALKALI METALS :

(लीना की रब से फरियाद)

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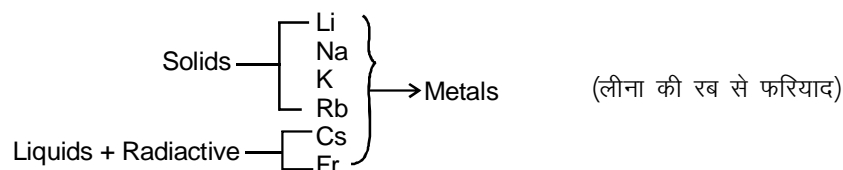
☞ <u>**GROUP – II A OR 2 , ALKALINE EARTH METALS :**</u>

```
          ┌ Be ┐
          │ Mg │
Solids ───┤ Ca ├──→ Metals      (बेटा मांगे कन्या सुन्दर बाप राजी)
          │ Sr │
          │ Ba ┘
          └ Ra ──→ Radioactive
```
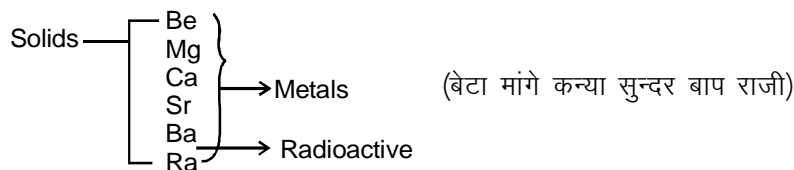
p-block elements

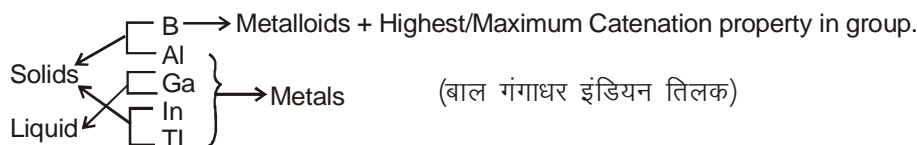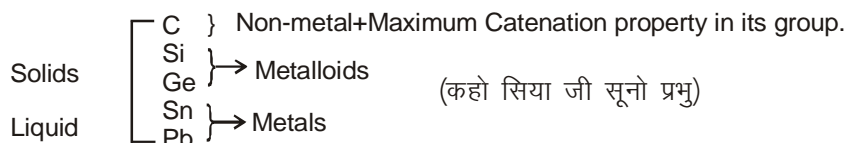When shells upto (n -1) are completely filled and differentiating electron entres the p-orbital of the nth orbit, elements of this class are called p-block elements.

- Group 13 to 18 elements constitute the p-block.
- General electronic configuration is [inert gas] ns^2np^{1-6}
- p-block elements lie on the extreme right of the periodic table.
- This block includes some metals, all non-metals and metalloids.
- s-block and p-block elements are collectively called normal or representative elements.

☞ <u>**GROUP – III A OR 13 , BORON FAMILY :**</u>

```
              ┌ B ──→ Metalloids + Highest/Maximum Catenation property in group.
              │ Al ┐
Solids ←──────┤ Ga ├──→ Metals      (बाल गंगाधर इंडियन तिलक)
              │ In │
Liquid ←──────┘ Tl ┘
```

☞ <u>**GROUP – IV A OR 14 , CARBON FAMILY :**</u>

```
          ┌ C  } Non-metal+Maximum Catenation property in its group.
          │ Si ┐
Solids    │ Ge ├──→ Metalloids      (कहो सिया जी सूनो प्रभु)
          │ Sn ┐
Liquid    └ Pb ┘──→ Metals
```

☞ <u>**GROUP – V A OR 15 , NITROGEN FAMILY :**</u>

```
Gas ←───┌ N  ┐
        │ P  ├──→ Maximum Catenation property in its group
        │ As ┐
Solids ←┤ Sb ├──→ Metalloids      (न्यूजीलैण्ड, पाकिस्तान आस्ट्रेलिया सब भिखारी)
        └ Bi }──→ Metals
```
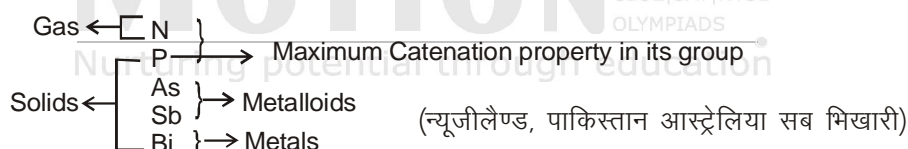
☞ <u>**GROUP – VI A OR 16 , OXYGEN FAMILY :**</u>

```
Gas ←───┌ O  ┐──→ Maximum Catenation property in its group
        │ S  ┤──→ Non-metals
Solids ←┤ Se │
        │ Te ┤──→ Metalloids      (O style से टीपो)
        └ Po ┘──→ Radiactive
```
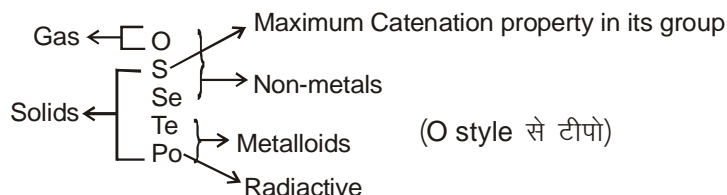
d-Block elements

When outermost (nth) and penultimate shells (n - 1)th shells are incompletely filled and differentiating electon enters the (n -1) d orbitals (i.e., d-orbital of penultimate shell) then elements of this class are called d-block elements.

• Group 3 to 12 elements constitute the d-block.

• General electronic configuration is [inert gas] $(n-1)d^{1-10}\ ns^{0-2}$

• d-block elements are classified into four series

Series	Elements	(n−1)d being filled
3d	$_{21}Sc - _{30}Zn$	3d
4d	$_{39}Y - _{48}Cd$	4d
5d	$_{57}La, _{72}Hf - _{80}Hg$	5d
6d	$_{89}Ac, _{104}Rf - _{112}Uub$	6d (incomplete series)

• Those elements which have paratially filled d-orbits in neutral state or in any stable oxidation state are called stransition elements.

f-Block elements

When n, (n−1) and (n −2) shells are incompletely filled and last electron enters into f-orbital of antepenultimate i.e., (n−2)th shell, elements of this class are called f-block elements., General electronic configuration is $(n-2)f^{1-14}\ (n-1)d^{0-1}\ ns^2$

• All f-block elements belong to 3rd group.

• They are metals

• Within each series, the properties of the elements are quite similar.

• They are also called as inner-transition elements as they contain three outer most shell incomplete and were also referred to as rare earth elements since their oxides were rare in earlier days.

   The elements of f-block have been classfied into two series.

• The actinides and lanthanides have been placed at the bottom of the periodic table to avoid the undue expansion of the periodic table.

1. Ist inner transition or 4 f-series, contains 14 elements $_{58}Ce$ to $_{70}Lu$. Filling of electrons takes place in 4f subshell.

2. IInd inner transition or 5 f-series, contains 14 elements $_{90}Th$ to $_{103}Lr$. Filling of electrons takes placed in 5f subshell.

3.NOMENCLATURE OF THE ELEMENTS WITH ATOMIC NUMBER > 100 (IUPAC)

According to IUPAC, elements with atomic number > 100 are represented by three latter symbols.

• These symbols are based on first letter of numbers from 0 to 9. The names of these number are derived from Greek and latin languages.

• The latin words for various digits of the atomic number are written together in the order of which makes the atomic number and suffix 'ium' is added at the end. In case of bi and tri one 'i' is omitted.

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NOTATION FOR IUPAC NOMENCLATURE OF ELEMENTS

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	p
6	hex	h
7	sept	s
8	Oct	o
9	enn	e

NOMENCLATURE OF ELEMENTS

Atomic Number	Name	Symbol	IUPAC Official Name	IUPAC symbol
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassnium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnilium	Uun	Darmstadtium	Ds
111	Unununium	Uuu	*	*
112	Ununbium	Uub	*	*
113	Ununtrium	Uut	+	
114	Ununquadium	Uuq	*	*
115	Ununpentium	Uup	+	
116	Ununhexium	Uuh	+	
117	Ununseptium	Uus	+	
118	Ununoctium	Uuo	+	

- Official IUPAC name yet to be announced   +   Elements yet to be discovered

•• IUPAC recommended this nomenclature to be followed until their names are officially recognised

4.PREDICTION OF PERIOD, GROUP AND BLOCK

☞ The block of an element corresponds to the type of subshell which receives the last electron

☞ The group is predicted from the number of electrons in the valence shell or/and penultimate shell as follows.

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(a) For s-block elements

   Group number = the no. of valence electrons

(b) For p-block elements

   Group number = 10+ no. of valence electrons

(c) For d-block elements

   Group number = no. of electrons in (n - 1) d sub shell + no. of electrons in valence shell.

5.METALS AND NON-METALS

• The metals are characterised by their nature of readily giving up the electron and from shinning lusture. Metals comprises more than 75% of all known elements and appear on the left hand side of the periodic table. Metals are usually solids at room temperature (except mercury). They have high melting and boiling points and are good conductors of heat and electricity. Oxides of metals are basic in nature. (Some metals in their higher oxidation state form acid oxides e.g. CrO_3)

• Non-metals do not lose electrons but take up electrons to form corresponding anions. Non-metals are located at the top right hand side of the periodic table. Non-metals are usually solids or gases at room temperature with low melting and boiling points. They are poor conductors of heat and electricity. Oxides of non-metals are acidic in nature.

6. METALLOIDS (SEMI METALS)

• It is very much clear from the periodic table that non-metallic character increases as we move from left to right across a row. It has been found that some elements lying at the border of metallic and non-metallic behaviour, possess the properties that are characteristic of both metals and non-metals. These elements are called semi metals or metalloids.

• The metalloids comprise of the elements B, Si, Ge, As, Sb and Te.

• Oxides of metalloids are generally amphoteric in nature.

7.TYPICAL ELEMENTS :

• Third period elements are called as typical elements. These include Na, Mg, AI, Si, P, S, CI.

• The properties of all the elements belonging to a particular group resemble the properties of the corresponding typical element of that group. For example, the general properties of alkali metals (IA) can be predicted from the properties of Na, not Li, the first member of the group.

• The properties of the elements of second period differ in the many respect belonging to the same group due to the smaller atomic size and absence of vacant d-orbitals.

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8. DIAGONAL RELATIONSHIP :

Some elements of certain groups of 2nd period resemble much in properties with the elements of third period of next group i.e. elements of second and third period are diagonally related in properties. This phenomenon is known as diagonal relationship.

2ⁿᵈ period   Li   Be   B   C

3ʳᵈ period   Na   Mg   Al   Si
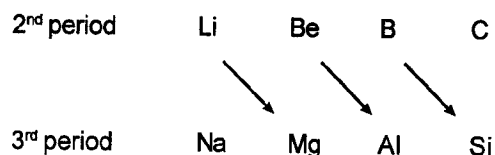
Diagonal relationship arises because of

(i) similar size of atom and ions

(Li = 1.23 Å & Mg = 1.36 Å; Li^+ = 0.60 Å & Mg^{2+} = 0.65 Å

(ii) similar electropositive characters

(iii) similar polarising powers (charge to radius ratio)

(iv) similarity in electronegativity values

    (Li = 1.0 & Mg = 1.2; Be = 1.5 & Al = 1.5)

- Similarities between properties of Li and Mg are as follows.

(a) Li and Mg both reacts directly with nitrogen to form lithium nitride (Li_3N) and magnesium nitride (Mg_3N_2) whereas other alkali metals of IA group does not form nitride.

(b) Fluoride, carbonate and phosphate of Li and Mg are insoluble in water whereas these compounds of other alkali metals are soluble.

(c) Li and Mg both are hard metals, whereas other metals of IA group are soft.

(d) LiOH and $Mg(OH)_2$ both are weak bases, whereas hydroxides of other elements of IA group are strong base.

(f) Metallic bond in Li and Mg both are strong compare to other alkali metals.

(g) Their melting and boiling points are high.

(h) By thermal disintegration of $LiNO_3$ and $Mg(NO_3)_2$; Li_2O and MgO is obtained respectively.

(I) Thermal stability of Li_2CO_3 and $MgCO_3$ is very less compare to other alkali metals and they liberates CO_2 gas easily.

- Similarly Be shows similarity to Al of IIIA group compare to other elements of IIA group which are as follows.

(a) These both elements do not provide colour to Bunsen bumer.

(b) They both are comparatively stable in air.

(c) Both are insoluble in NH_3 therefore do not form blue coloured solution.

(d) There is no tendency of making peroxide and superoxide in them.

(e) Reducing power is very less due to low value of standard electrode potential in the form of oxidation potential.

(f) Be and Al both forms halogen bridge halides.
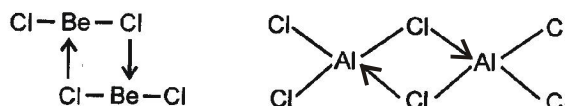
9. THE PERIODICITY OF ATOMIC PROPERTIES:

EFFECTIVE NUCLEAR CHARGE:

Between the outer most valence electrons and the nucleus of an atom, there exists finite number of shells containing electrons. Due to the presence of these intervening electrons, the valence electrons are unable to experience the attractive pull of the actual number of protons in the nucleus. These intervening electrons act as shield between the valence electrons and protons in the nucleus. Thus, the presence of intervening (shielding) electrons reduces the electrostatic attraction between the protons in the nucleus and the valenece electrons because intervening electrons repel the valence electrons. The concept of effective nuclear charge allows us to account for the effects of shielding on periodic properties.

The effective nuclear charge (Z_{eff}) is the charge felt by the valence electron. Z_{eff} is given by $Z_{eff} = Z - \sigma$. Where Z is the actual nuclear charge (atomic number of the element) and σ is the shielding (screening) constant.

ATOMIC RADIUS:

Probability of finding the electron is never zero even at large distance from the nucleus. Based on probability concept, an atom does not have well defined boundary. Hence exact value of the atomic radius can't be evaluated. Atomic radius is taken as the effective size which is the distance of the closet approach of one atom to another atom in a given bonding state.

Atomic radius can be

(A)   Covalent radius:

$$\frac{1}{2} AB = r_{covalent}$$
(of element X)
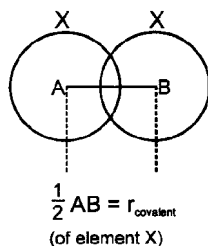
It is one-half of the distance between the centres of two nuclei (of like atoms) bonded by a single covalent bond.

•   Covalent radius is generally used for non-metals.

Single Bond Covalent Radius, SBCR –

(a) For Homoatomic moleucles $d_{A-A} = r_A + r_A$ or $2r_A$

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

•   (b) For hetrodiatomic molecules in which electro negativity remains approximately same.

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

•   For heteronuclear diatomic molecule, A – B, where difference between the electronegativity values of atom A and atom B is relativity larger,

$d_{A-B} = r_A + r_B - 0.09 (X_A - X_B)$

where X_A and X_B electronegativity values of high electronegative element A and less electronegative element B, respectively. This formula is given by **Stevenson & Schomaker.**

Ex.2 *Calculate the bond length of C – X bond, if C – C bond length is 1.54 Å, X – X bond length is 1.00 Å and electronegativity values of C and X are 2.0 and 3.0 respectively*

Sol. (1)   C – C bond length = 1.54 Å

$$r_C = \frac{1.54}{2} = 0.77 \text{ Å}$$

$$r_x = \frac{1.00}{2} = 0.50 \text{ Å}$$

(2)   C – X bond length

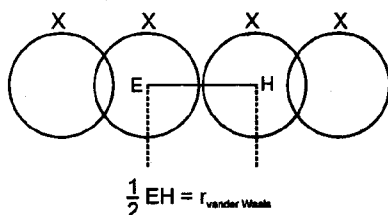$d_{C-x} = r_C + r_x - 0.09 (X_x - X_C)$
$= 0.77 + 0.50 - 0.09 (3 - 2)$
$= 0.77 + 0.50 - 0.09 \times 1$
$= 1.27 - 0.09 = 1.18$ Å

Thus C – X bond length is 1.18 Å

(B)   Van der Waals radius (Collision radius) :

$$\frac{1}{2} EH = r_{vander Waals}$$

It is one - half of the internuclear distance between two adjacent atoms in two nearest neighbouring molecules of the substance in solid state.

•   van der Waa's radius does not apply to metal. its magnitude depends upon the packing of the atoms when the element is in the solid state.

•   **Comparision of convalent radius and van der Waal's radius**

(i)   The van der Waal's force of attractions are weak, therefore, their internuclear distances in case of atoms held by van der Waal's forces are much larger than those of between covalently bonded atoms. Therefore van der Waal's radii are always larger than covalent radii.

(ii)   A covalent bond is formed by the overlaping of two half-filled atomic orbitals, a part of the orbital becomes common. Therefore, covalent radii are always smaller than the van der Waals radii. For example,

| Elements	H	O	F	S	Br
Covalent radius (Å)	0.37	0.66	0.64	1.04	1.11
van der Waal's radius (Å)	1.20	1.40	1.35	1.85	1.95

(C)   Metallic radius (crystal radius) :

$$\frac{1}{2} CD = r_{crystal}$$
(of element X)
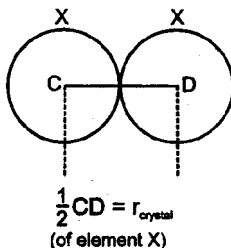
It is one -half of the distance between thenuclei of two adjacent metal atoms in the metallic crystal lattice. Metallic radius of an element is always greater than its covalent radius. It is due to the fact that metallic bond (electrical attraction between positive charge of an atom and mobile electrons) is weaker than covalent bond and hence the hence the internuclear distance between the two adjacent atoms in a metallic crystal is longer than the internuclear distance between the covalently bonded atom.

For example :

	Metallic radius	Covalent radius
K	231 pm	203 pm
Na	186 pm	154 pm

- $r_{covalent} < r_{crystal} < r_{vander\ Walls}$
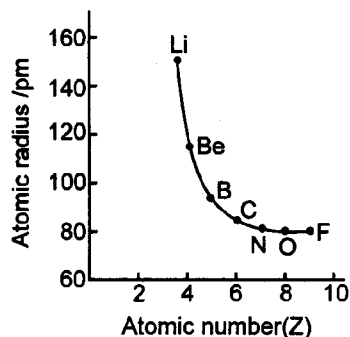
Variation of atomic radius with atomic number across the second period
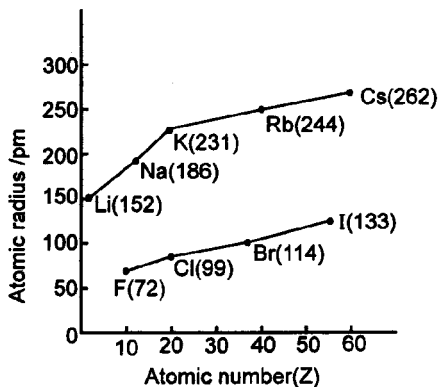
Variation of atomic radius with atomic number for alkali metals and halogens

Variation in a Period	Variation in a Group
In a period left to right :	In a group top to bottom :
Z increases by one unit	Z increases by more than one unit
Z_{eff} also increases	Z_{eff} almost remains constant (due to increased screening effect of inner shells electrons)
n remains constant (no. of orbits)	n increases (no. of orbits)
As a result of these electrons are pulled close to the nucleus by the increased Z_{eff}. $r_n \propto \dfrac{1}{z^*}$ Thus atomic radii decreases with increase in atomic number in a period from left to right	The effect of increased number of atomic shells overweigh the effect of increased screening effect. As a result of this the size of atom increases from top to bottom in a given group.

☞ The atomic radius of inert gas (zero group) is given largest in a period because it is represented as van der Waals's radius which is generally larger than the covalent radius. The van der Waal's radius of inert gases also increases from top to bottom in a group.

☞ In the transition series (e.g. in first transition series), the covalent radii of the elements decrease from left to right across a row until near the end when the size increases slightly. On moving from left to right, extra protons are placed in the nucleus and the extra electron are added. The orbital electrons shield the nucleus charge incompletely. Thus the nuclear charge attracts all the electrons more strongly, hence a contraction in size occurs. The radii of the elements from Cr to Cu, are very close to one another because the successive addition of d-electrons screen the outer electrons (4s) from the inward pull of the nucleus. As a result of this, the size of the atom does not change much in moving from Cr to Cu.

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Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Atomic radius (A)	1.44	1.32	1.22	1.18	1.17	1.17	1.16	1.15	1.17	1.25

☞ The lanthanide contraction counter balances almost exactly the normal size increase on descending a group of transition elements. Thus covalent and ionic radii of Nb (5th peroid) and Ta (6th period) are almost same due to poor shielding of f-orbitals electrons.

IONIC RADIUS:

The effective distance from the centre of nucleus of the ion up to which it has an influence in the ionic bond is called ionic radius.

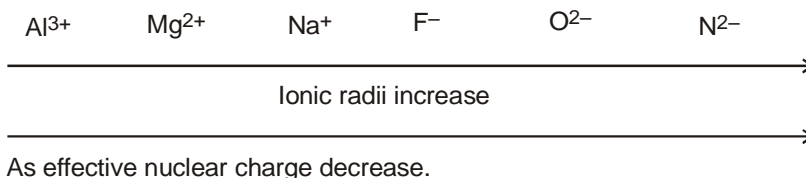| Cation	Anion
It is formed by the loss of one or more electrons from the valence shell of an atom of an element. Cations are smaller than the parent atoms because (i) The whole of the outer shell of electrons is usually removed. (ii) In a Cation, the number	It is formed by the gain of one or more electrons in the valence shell of an atom of an element. Anions are larger than the parent atoms because (i) Anion is formed by gain of one or more electrons in the neutral atom and thus number of electron

		For example :	Na	Na^+

For example :   Na Na^+
Number of Electrons   11   10
Number of Protons   11   11
Electronic Configuration $1s^2\ 2s^2\ 2p^6\ 3s^1$ $1s^2\ 2s^2\ 2p^6$
                             Cl Cl^-
Number of Electrons   17   18
Number of Proton   17   17

- The sizes of ions increases as we go down a group (cosidering the ions of same charge)
  For example
  $Li^+ < Na^+ < K^+ < Rb^+$
  $Be^{2+} < Mg^{2+} < Ca^{2+} < Sr^{2+}$
  $F^- < Cl^- < Br^- < I^-$
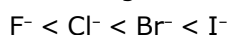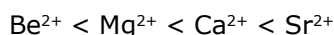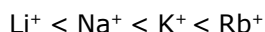
- The d and f orbitals do not shield the nuclear charge very effectively. Therefore there is significant reduction in the size of the ions, just after d or f orbitals have been filled completely. This is called a lanthanide contraction. Atomic radii of Zr and Hf are almost identical due to lanthanide contraction.

- The species containing the same number of electrons but differ in the magnitude of their nuclear charges are called as isoelectronic species. For example, N^{3-}, O^{2-}, F^-, Ne, Na^+, Mg^{2+} and Al^{3+} are all isoelectronic species with same number of electrons (i.e, 10) but different nuclear charges of +7, +8, +9, +10, +11, +12 and +13 respectively.
 Within a series of isoelectronic species as the nuclear charge increases, the force of attraction by the nucleus on the electrons also increases. As a result, the ionic radii of isoelectronic species decrease with increases in the magnitude of nuclear charges. For example,

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

  ――――――――――――――――――――――――――――――→
  Ionic radii increase

  ――――――――――――――――――――――――――――――→
  As effective nuclear charge decrease.

- Pauling's empirical formula : Ionic radius $\propto \dfrac{1}{nuclear\ charge}$

- Following are the examples of isoelectronic series
  (i) $S^{2-}\ Cl^-\ K^+\ Ca^{+2}\ Sc^{+3}$ (ii) $SO_2,\ NO_3^-,\ CO_3^{2-}$ (iii) $N_2,\ CO,\ CN^-$ (iv) $NH_3,\ H_3O^-$

IONISATION ENTHALPY

Ionisation enthalpy/energy (IE) , sometimes also called ionisation potential (IP) , of an element is defined as the amount of energy required to remove an electron from an isolated gaseous atom of that element resulting in the formation of positive ion.

$$M(g) \xrightarrow{(E_1)} M^+(g) + e^-$$

$$M^+ + IE_2 \longrightarrow M^{2+} + e^-$$

$$M^{2+} + IE_3 \longrightarrow M^{+3} + e^-$$
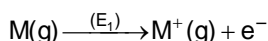
IE_2 & IE_3 are the IInd & IIIrd ionization energies to remove electron from monovalent and divalent cations respectively.

In general: $(IE)_1 < (IE)_2 < (IE)_3 < $ .... because, as the number of electrons decreases, the attraction between the nucleus'and the remaining electrons increases considerably and hence sUbsequent 1.E.(s) increase.

- Units of ionisation energy: KJ mol^{-1} ,K Cal mol^{-1}, eV (electron volt)
- Factors Influencing Ionisation enthalpy

    (IE) variation in a period and group may or may not be regular and can be influenced by:

(A)   Size of the Atom :

Ionisation energy decreases with increase in atomic size. As the distance between the outermost electrons and the nucleus increases, the force of attraction between the valence shell electrons and the nucleus decreases. As a result, outer most electrons are held less firmly and lesser amount of energy is required to knock them out.
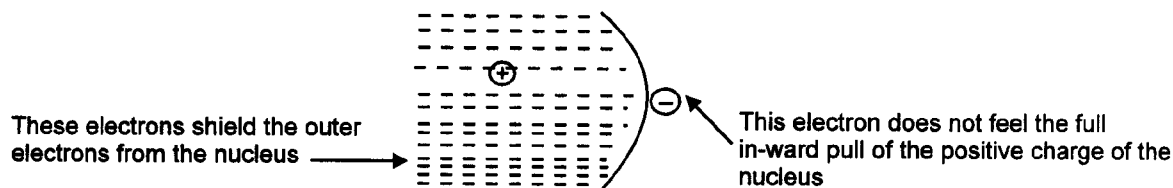
For example, ionisation energy decreases in a group from top to bottom with increase in atomic size.

(B)   Nuclear Charge:

The ionisation energy increases with increase in the nuclear charge. This is due to the fact that with increase in the nuclear charge, the electrons of the outer most shell are more firmly held by the nucleus and thus greater amount of energy is required to pull out an electron from the atom. For example, ionisation energy increases as we move from left to right along a period due to increase in nuclear charge.

(C)   Shielding effect:

The electrons in the inner shells act as a screen or shield between the nucleus and the electrons in the outermost shell. This is called shielding effect. The larger the number of electrons in the inner shells, greater is the screening effect and smaller the force of attraction and thus (IE) decreases.

These electrons shield the outer electrons from the nucleus ⟶

This electron does not feel the full in-ward pull of the positive charge of the nucleus

(D)   Penetration Effect of the Electron :

The ionisation energy increases as the penentration effect of the electrons increases.

It is a well known fact that the electrons of the s-orbital has the maximum probability of being found near the nucleus and this probability goes on decreasing in case of p, d and f orbitals of the same energy level. Within the same energy level, the penetration effect decreases in the order
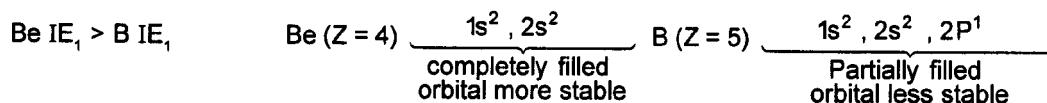
    s > p > d > f

Greater the penetration effect of electron more firmly the electron will be held by the nucleus and thus higher will be the ionisation energy of the atom.

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For example, ionisation energy of aluminium is comparatively less than magnesium as outer most electron is to be removed from p-orbital (having less penetration effect) in aluminium where as in magnesium it will be removed from s-orbital (having large penetration effect) of same energy level.

(E)   Electronic Configuration :
If an atom has exactly half-filled or completely filled orbitals, then such an arrangement has extrastability. The removal of an electron from such an atom requires more energy than expected. For example,

$$\text{Be IE}_1 > \text{B IE}_1 \qquad \text{Be (Z = 4) } \underbrace{1s^2 , 2s^2}_{\substack{\text{completely filled} \\ \text{orbital more stable}}} \quad \text{B (Z = 5) } \underbrace{1s^2 , 2s^2 , 2P^1}_{\substack{\text{Partially filled} \\ \text{orbital less stable}}}$$
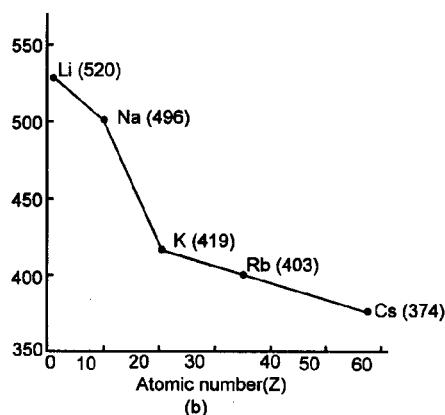
As noble gases have completely filled electronic configuration, they have highest ionisation energies in their respective periods.
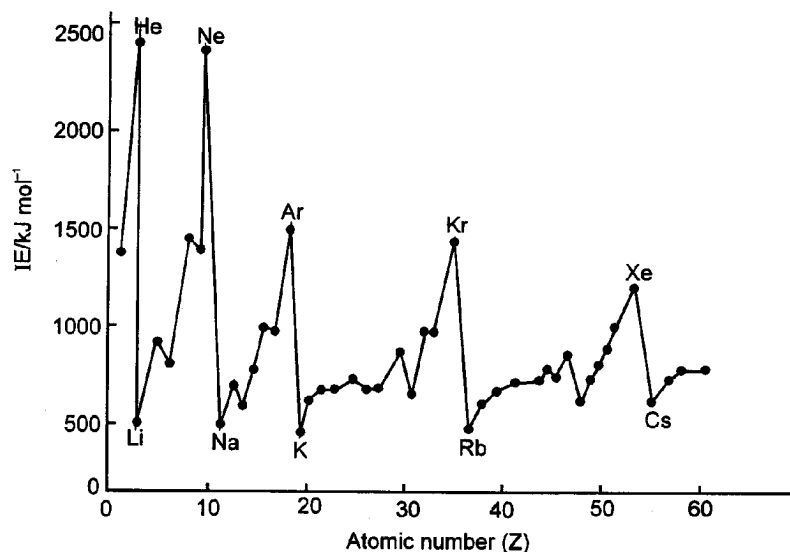
Variation of first ionization energy with atomic number of elements with Z = 1 to 60
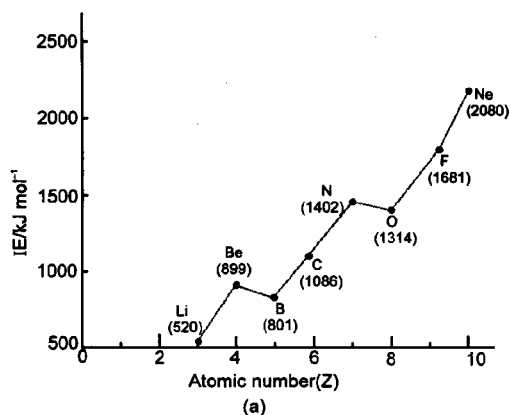
(a) First ionization energy of elements of the second period as a function of atomic number (b) First ionization energy of alkali metal as a function of atomic number.

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- Metallic or electropositive character of elements increases as the value of ionisation energy decreases.
- The relative reactivity of the metals increases with the decrease in ionisation energy.
- The reducing power of elements increases as the value of ionisation energy decreases (Li is exception in Alkali metals group which has highest reducing power)

Ex.3 *First and second ionisation energies of Mg(g) are 740 and 1450 kJ mor⁻¹. Calculate percentage of Mg⁺(g) and Mg²⁺(g), if 1 g of Mg(g) absorbs 50 kJ of energy.*

Sol. Number of moles of 1g of Mg $= \dfrac{1}{24} = 0.0417$

Energy required to convert Mg(g) to Mg⁺(g)
$= 0.0417 \times 740 = 30.83$ kJ
Remaining energy $= 50 - 30.83 = 19.17$ kJ

Number of moles of Mg²⁺ formed $= \dfrac{19.17}{1450} = 0.0132$

Thus, remaining Mg⁺ will be $= 0.0417 - 0.0132 = 0.0285$

% Mg⁺ $= \dfrac{0.0285}{0.0417} \times 100 = 68.35\%$

% Mg⁺ $= 100 - 68.35 = 31.65\%$
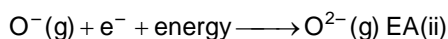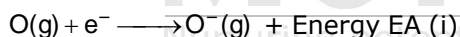
ELECTRON GAIN ENTHALPY (ELECTRON AFFINITY) :

When an electron is added to a neutral gaseous atom (X) to convert it into a negative ion, the enthalpy change accompanying the process is defined as the electron gain enthalpy.
Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion.

$$X(g) + e^- \longrightarrow X^-(g); \ \Delta H = \Delta H_{eq} \qquad \dots(i)$$
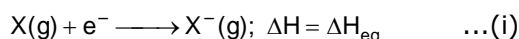
The negative of the enthalpy change for the process depicted in equation (i) is defined as the electron affinity (EA) of the atom undergoing the change for the formation of.an anion.
Depending on the elements, the process of adding an electron to the atom can be either endothermic or exothermic. When an electron is added to the atom and the energy is released, the electron gain enthalpy is negative and when energy is needed to add an electron to the atom, the electron gain enthalpy is positive. The addition of second electron to an anion is opposed by electrostatic repulsion and hence the energy has to be supplied for the addition of second electron.

$$O(g) + e^- \longrightarrow O^-(g) + \text{Energy} \quad \text{EA (i)}$$

$$O^-(g) + e^- + \text{energy} \longrightarrow O^{2-}(g) \quad \text{EA(ii)}$$

EA (i) is exothermic whereas EA(ii) is endothermic.

- Group 17 elements (halogens) have very high negative electron gain enthalpies because they can attain stable noble gas electronic configuration by picking up an electron.
- Noble gases have large positive electron gain enthalpies because the electron has to enter the next higher energy level leading to a very unstable electronic configuration.
- Electron gain enthalpy of O or F is less than S or CI. This is due to the fact that when an electron is added to O or F, the added electron goes to the smaller n = 2 energy level and experiences significant repulsion from the other electrons present in this level. In S or CI, the electron goes to the larger n = 3 energy level and consequently occupies a larger region of space leading to much less electron-electron repulsion.

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☛ Electron gain enthalpies of alkaline earth metals are very less or positive because the extra electron is to be added to completely filled s-orbitals in their valence shells.

☛ Across a period, with increase in atomic number, electron gain enthalpy becomes more negative because left to right across a period effective nuclear charge increases and consequently it will be easier to add an electron to a small atom.

☛ As we move in a group from top to bottom, electron gain entalpy becomes less negative because the size of the atom increases and the added electron would be at larger distance from the nucleus.

☛ (i) Electron affinity $\propto \dfrac{1}{\text{Atomic size}}$ (ii) Electron affinity $\propto$ Effective nuclear charge (z_{eff})

(iii) Electron affinity $\propto \dfrac{1}{\text{screening effect}}$ (iv) Stability of half filled and completely filled orbitals of a subshell is comparatively more and the addition of an extra electron to such an system is difficult and hence the electron affinity value decreases.

Ex.4 *How many CI atoms can you ionise in the process* $Cl \longrightarrow Cl^+ + e$ *the energy liberated for the process* $Cl + e \longrightarrow Cl^-$ *for one Avogadro number of atoms. Given IP = 13.0 eV and EA = 3.60 eV*

Sol. Let n atoms be ionised.

$6.02 \times 10^{23} \times EA = n \times IP$

$n = \dfrac{6.02 \times 10^{23} \times 3.60}{13} = 1.667 \times 10^{23}$

Ex.5 *The first ionisation potential of Li is 5.4 eV and the electron affinity of CI is 3.6 eV Calculate ΔH in kcal mol⁻¹ for the reaction.*

$Li(g) + Cl(g) \longrightarrow Li^+ + Cl^-$
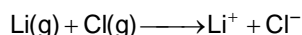
Carried out at such low pressures that resulting ions do not combine with each other.

Sol. The overall reaction is written into two partial equations

$Li(g) \longrightarrow Li^+ + e$ ΔE_1 = 5.4 eV

$Cl(g) + e \longrightarrow Cl^-$ ΔE_2 = −3.6 eV

$\Delta H = \Delta E_1 - \Delta E_2 = 5.4 - 3.6 = 1.8$ eV
$= 1.8 \times 23.06$ kcal mol⁻¹
$= 41.508$ kcal mol⁻¹

Ex.6 *For the gaseous reaction,*

$K_{(g)} + F_{(g)} \longrightarrow K^+ + F^-, \Delta H$ *was calculated to be 19 kcal under conditions where the cations and anions were prevented by electrosatic separation from combining with each other. The ionisation potential of K is 4.3 eV. What is the electron effinity of F ?*

Sol. $K \longrightarrow K^+ + e$ ΔE_1 = 4.3 eV

$F + e \longrightarrow F^-$ $\Delta E_2 = -E$ eV

$\dfrac{19.0}{23.06} = \Delta E_1 - \Delta E_2 = 4.3 - E$

$0.82 = 4.3 - E \Rightarrow E = 3.48$

Ex.7 *The electron affinity of chlorine is 3.7 eV. How much energy in kcal is released when 2 g of chlorine is completely converted to CI⁻ ion in a gaseous state?*
(1 eV = 23.06 kcal mol⁻¹)

Sol.

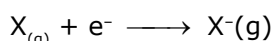$C + e \longrightarrow Cl^- + 3.7eV$

35.5 3.7×23.06 kcal

.'. Energy released for conversion of 2 g gaseous chlorine into CI⁻ ions

$= \dfrac{3.7 \times 23.06}{35.5} \times 2 = 4.8$ kcal

ELECTRO-NEGATIVITY

Electronegativity is a measure of the tendency of an element to attract electrons towards itself in a covalently bonded molecules .

$$X_{(g)} + e^- \longrightarrow X^-(g)$$

The magnitude of electronegativity of an element depends upon its ionisation potential & electron affinity. Higher ionisation potential & electron affinity values indicate higher electronegativity value.

☛ With increase in atomic size the distance between nucleus and valence shell electrons incerases, therefore, the force of attraction between the nucleus and the valence shell electrons decreases and hence the electronegativity values also decrease.

☛ With increase in nuclear charge force of attraction between nucleus and the valence shell electrons increases and, therefore, electronegativity value increases

☛ In higher oxidation state, the element has higher magnitude of positive charge. Thus, due to more positive charge on element, it has higher polarising power. Thus, with increase in the oxidation state of element, its electronegativity also increases.

☛ Charge on cation α electronegativity of the atom

☞ The electronegativity increases as the s-character in the hybrid orbitals increases.

Hybrid orbital	sp³	sp²	sp
s-character	25%	33%	50%

Electronegativity increases →

Variation of (EN) in a group	Variation of (EN) in a period
On moving down the groups Z increases but Z_{eff} almost remains constant, number of shells (n) increases, r_n (atomic radius) increases. Therefore (EN) decreases moving down the groups.	While moving across a period left to right. Z, Z_{eff} increases & r_n decreases. Therefore (EN) increases along a period.

There is no direct method to measure the value of electronegativity, however, there are some scales to measure its value .

(a)   Pauling's Scale:

Linus Pauling developed a method for calcUlating relative electronegativities of most elements. According to Pauling

$$\chi_A - \chi_B = 0.208\sqrt{\Delta} \qquad \text{where } \Delta = \varepsilon_{A-B} - \sqrt{\varepsilon_{A-A}\varepsilon_{B-B}}$$

where ε'^S refer to bond enthalpies (K Cal mol⁻¹). By assigning $\chi_F = 4.0$, χ values of other atoms can be computed.

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(b) **Mulliken's scale**

Electronegativity (EN) can be regarded as the average of the ionisation energy (IE) and the electron affinity (EA) of an atom.

$$(EN) = \frac{IE + EA}{2}$$

If both (EA) and (IE) are determined in eV units then paulings's electronegativity $(EN)_p$ is related to Mulliken's electronegativity. Mulliken's values were about 2.8 times larger than the Pauling's values.

(c) **Allred-Rochow's Electronegatlvlty**

Allred and Rochow defined electronegativity as the force exerted by the nucleus of an atom on its valence electrons:

$$(EN)_{AR} = \frac{0.359\, Z_{effective}}{r^2} + 0.744$$

where $Z_{elleclive}$ is the effective nuclear charge and r the covalent radius (in Å).

☛ The electron negativity of Cs (55) is less than Fr (87). This is due to the increase of + 32 units in nuclear charge of Fr which makes the effective nuclear charge comparatively high.

☛ The electroneativity of inert gas elements of zero group is zero. Inert gases exist as monoatomic molecules and the electronegativity is the property of bonded atoms.

Ex.8 *Ionisation potential and electron affinity of fluorine are 17.42 and 3.45 eV respectively. Calculate the electronegativity of fluorine.*

Sol. According to Mulliken equation

$X = \dfrac{IP + EA}{5.6}$ when both IP and EA are taken in eV.

$$X_F = \frac{17.42 + 3.45}{5.6} = 3.726$$

APPLICATIONS OF ELECTRONEGATIVITY :

(I) **Nomenclature:**

Compounds formed from two nonmetals are called binary compounds. Name of more electronegative element is written at the end and 'ide' is suffixed to it. The name of less electronegative element is written before the name of more electronegative element of the formula.

Ex.9 *Write the correct formula and name of the following*

(a) ICI or CIl   *(b) FCI or CIF*   *(c) BrCI or CIBr*   *(d) BrI or IBr*

(e)OF$_2$ or F$_2$O   *(f)Cl$_2$O or OCI$_2$*

Sol.   **Correct formula**   **Name**

(a) $I^+ Cl^-$ Iodine chloride

(b) $CI^+ F^-$ Chlorine fluoride

(c) $Br^+ CI^-$ Bromine chloride

(d) IBr   Iodine bromide

(e) OF_2 Oxygen difluoride

(f) Cl_2O Dichlorine oxide

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(II)   Nature of Bond:

If difference of electronegativities of the two elements is 1.7 or more, then ionic bond is formed between them whereas if it is less than 1.7, then covalent bond is formed. (HF is exception in which bond is covalent although difference of electronegativity is 1.9)

(iii)   Metallic and Nonmetallic Nature :

Generally values of electronegativity of metallic elements are low, whereas electronegativity values of nonnmetals are high.

(iv)   Partial Ionic Character in Covalent bonds

Partial ionic characters are generated in covalent compounds by the difference of electronegativities.

Hanny and smith calculated percentage of ionic character from the difference of electronegativity.

$$\text{Percentage of ionic character} = 16(X_A - X_B) + 3.5 (X_A - X_B)^2$$
$$= 16\Delta + 3.5\Delta^2$$
$$= (0.16\Delta + 0.035 \Delta^2) \times 100$$

X_A is electronegativity of element A

X_B is electronegativity of element B

$\Delta = X_A - X_B$

(v)   Bond length

When difference of electronegativities of atoms present in a molecule is increased, then bond length decreases. Shoemaker and stephensen determined.

$$\text{Bond length } d_{A-B} = r_A + r_B - 0.09 |X_A - X_B|$$

$$\text{or } d_{A-B} = \frac{1}{2} (D_{A-A} + D_{B-B}) - 0.09 |X_A - X_B|$$

(vi)   Bond Strength & Stability

Bond strength and stability of A – B increases on increase in difference of electronegativities of atoms A and B bonded A – B. Therefore H – F > H – Cl > H – Br > H – I

Ex.10 *Electronegativity of which of the following is high ?*

(1) −CH₃(sp³)   *(2) H₂C = CH₂(sp²)* *(3) CH ≡ CH(sp)*   *(4) Equal in all*

Ans. **(3)**

Ex.11 *CF₃NH₂ is not a base, whereas CH₃NH₂ is a base. What is the reason ?*

Sol.   Due to high electronegativity of F tendency of donating the lone pair of electrons present on N will be less

Ex.12 *OF₂ is called oxygen difluoride, whereas Cl₂O is called dichlorine monoxide. Why ?*

Sol.   Electronegativity of O in OF₂ is less than F. Therefore, there will be positive charge on oxygen and negative charge on fluorine. Whereas in Cl and O, electronegativity of Cl is less than that of O therefore there will be positive charge on Cl and negative charge on O. Positive charge is written first followed by negative charge.

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Ex.13 *Calculate the electronegativity of fluorine from the following data :*

E_{H-H} = 104.2 kcal mol^{-1}, E_{F-F} = 36.6 kcal mol^{-1}

E_{H-F} = 134.6 kcal mol^{-1}, X_H = 2.1

Sol. Let the electronegativity of fluorine be X_F.

Applying Pauling's equation.

$$X_F - X_H = 0.208[E_{H-F} - (E_{F-F} \times E_{H-H})^{1/2}]^{1/2}$$

In this equation, dissociation energies are taken in kcal mol^{-1}.

$X_F - 2.1 = 0.208 [134.6 - (104.2 \times 36.6)^{1/2}]^{1/2}$

$X_F = 3.87$

Ex.14 *The electron affinity of chlorine is 3.7 eV. How much energy in kcal is released when 2 g of chlorine is completely converted to Cl$^-$ ion in a gaseous state ?*

(1 eV = 23.06 kcal mol^{-1})

Sol. Cl + e $\longrightarrow$ Cl$^-$ + 3.7 eV

35.5   3.7 × 23.06 kcal

∴ Energy released for conversion of 2 g gaseous chlorine into Cl$^-$ ions

$= \dfrac{3.7 \times 23.06}{35.5} \times 2 = 4.8$ kcal

Ex.15 *Calculate the electronegativity of fluorine from following data :*

E_{H-H} = 104.2 kcal mol^{-1}

E_{F-F} = 36.6 kcal mol^{-1}

E_{H-F} = 134.6 kcal mol^{-1}

Electronegativity of H is 2.05.

Sol. On Paulling scale :

$x_F - x_H = 0.182\sqrt{\Delta_H - F}$

(using B.E. in kcal mol^{-1})

$\Delta_{H-F} = E_{H-F} - \sqrt{E_{H-H} \times E_{F-F}} = 13.45 - \sqrt{104.2 \times 36.6} = 72.84$ kcal

From (i)

$x_F - x_H = 0.182 \sqrt{72.84} = 1.5534$

$x_F = x_H + 1.4434$ $= 2.05 + 1.5534 = 3.6034$

UNSOLVED EXERCISE :

1. The standard enthalpies of formation of gaseous XeF$_2$, XeF$_4$ & XeF$_5$ are –108, –216 & –294 kJ mol^{-1} respectively and the bond energy in F$_2$ is 159 kJ mol^{-1}. Calculate the average Xe–F bond energy in each of these compounds and use of the value for XeF$_2$ to obtain a value for the electronegativity of xenon on the Pauling scale assuming the electronegativity of fluorine to be 4.

2. Calculate the electronegativity value of chlorine on Mulliken's scale, given that IP = 13.0 eV and EA = 4.0 eV.

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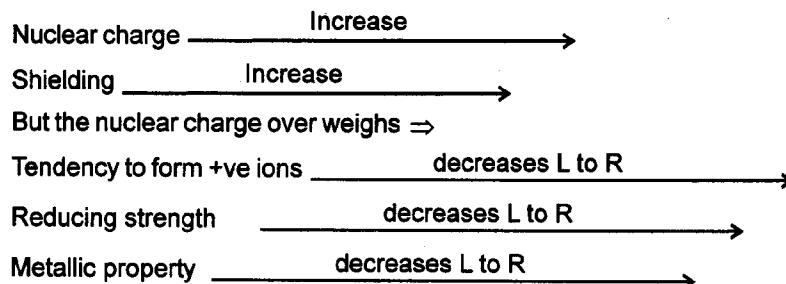
(VII) METALLIC PROPERTY

Metals have the tendency to form cations by loss of electrons and this property makes the elements as **electropositive** elements or metals.

$$M(g) \longrightarrow M^+_{(g)} + e^-$$

For a given group:

Nuclear Charge	Shielding		Tending to make + ve	Reducing Stength	Metallic Property
increases ↓	increases ↓	But the shielding effect over weighs so ⇒	increases ↓	increases ↓	increases ↓

For a given period:

Nuclear charge ————— Increase ————→

Shielding ————— Increase ————→

But the nuclear charge over weighs ⇒

Tendency to form +ve ions ——————— decreases L to R ———————→

Reducing strength ————— decreases L to R —————→

Metallic property ————— decreases L to R —————→
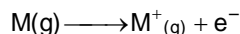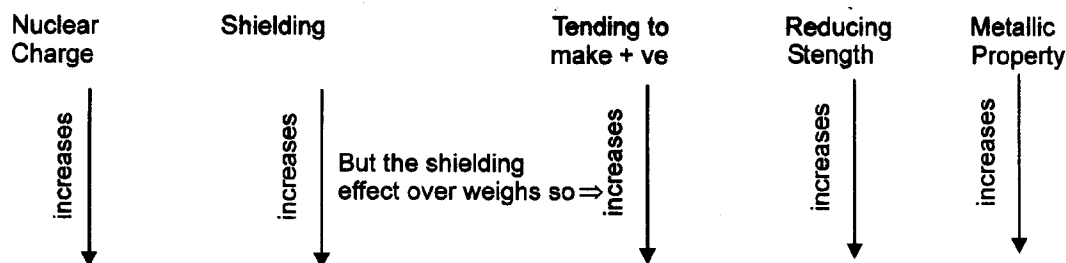
(IX) OXIDES

Oxygen react with all elements except noble gases, Au, Pd and Pt to form oxides. In general, metallic oxides (O^{2-}), peroxides (O_2^{2-}) and super oxides (O_2^{-1}) are ionic solids.

The tendency of group IA metals (alkali metals) to form oxygen rich compounds increases from top to bottom i.e. with increasing cation radii and decreasing charge density on the metal ion.

IIA metals also show the similar trend. Except Be, the IIA metals react with oxygen at normal conditions to form normal ionic oxides and at high pressure of O_2, they form peroxides (CaO_2, SrO_2, BaO_2). Oxides of metals are called as basic anhydries as most of them combine with water forming hydroxides with no change in oxidation state of metals.

Oxides of IA and IIA dissolve in water forming basic solution where as other oxides do not dissolve in water.

$$Na_2O + H_2O \longrightarrow 2NaOH$$
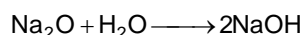
Oxygen combines with many non-metals to form covalent oxides such as CO, CO_2, SO_2, P_4O_{10}, Cl_2O_7 etc. Non-metals with limited supply of oxygen usually form oxides in which non-metals are present in lower oxidation states where as with excess of oxygen, oxides with higher oxidatin state are formed. Oxides of non-metals are called as acid anhydrides as most of them dissolve in water forming acids of oxy-acids.

$$P_4O_{10} + 6H_2O \longrightarrow 4H_3PO_4 \; ; \; SO_3 + H_2O \longrightarrow H_2SO_4 : Cl_2O_7 + H_2O \longrightarrow 2HClO_4$$

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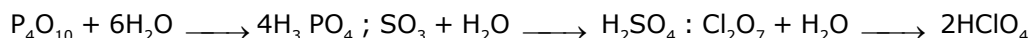
☛ In a group, basic nature of oxides increases or acidic nature decreases. Oxides of the metals are generally basic and oxides of the non-metals are acidic. The oxides of the matalloids are amphotetric in nature. The oxides of Al, Zn, Sn, As and Sb are amphoteric.
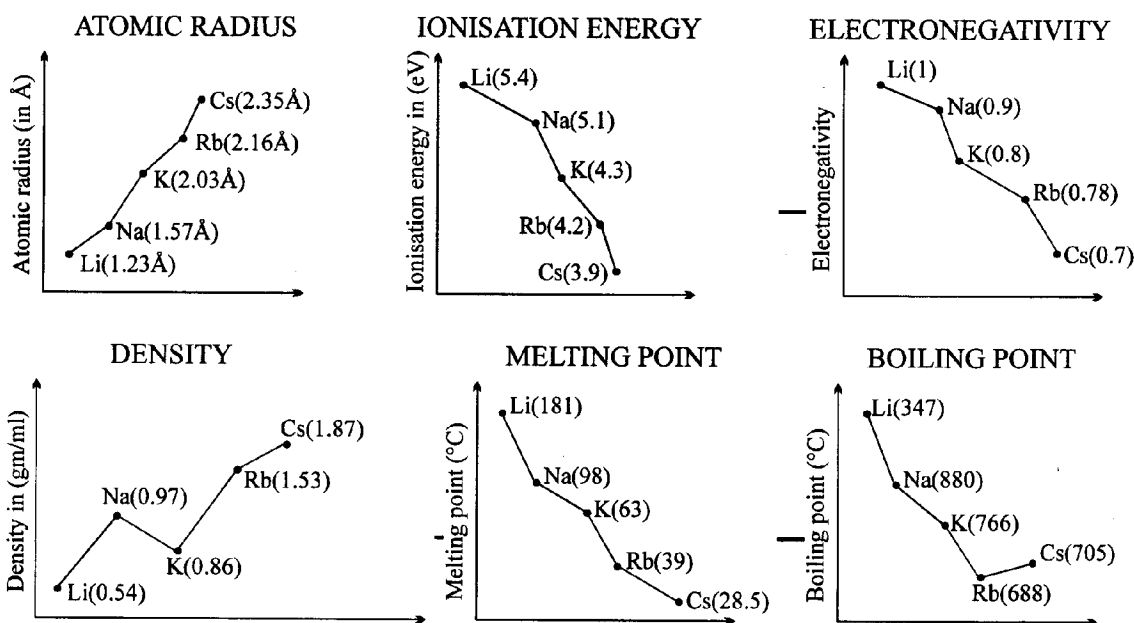
☛ In a period the nature of the oxides varies from basic to acidic.

Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
Strongly basic	Basic	amphoteric	Weakly acidic	Acidic	Acidic	Strongly acidic

☛ CO, N_2O, NO and H_2O are neutral oxides.
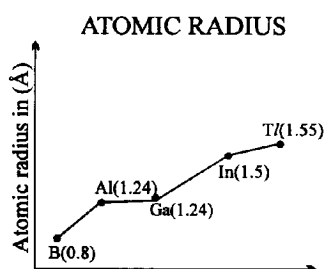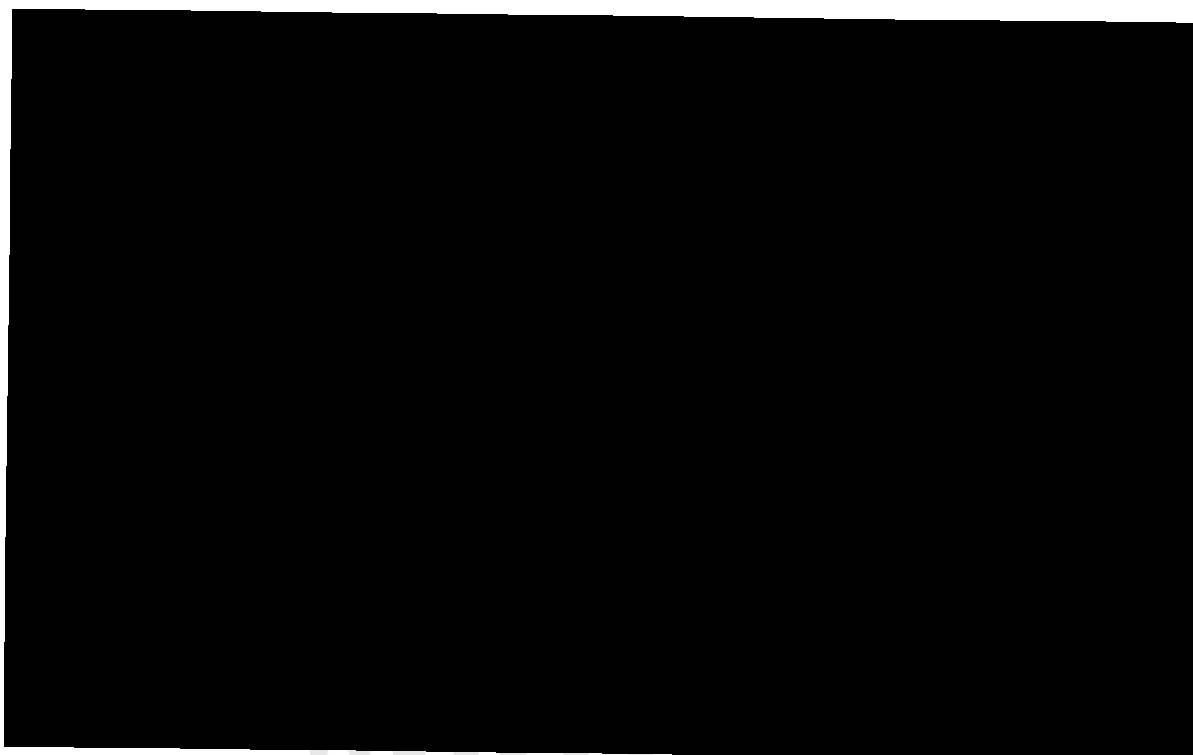
GRAPHS OF PERIODIC PROPERTIES

ALKALI METALS

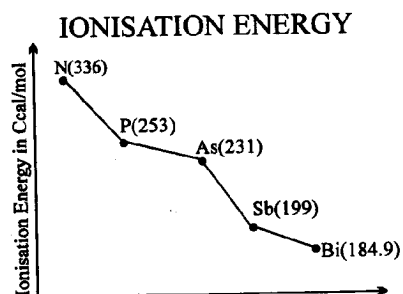## LANTHANIDES

ATOMIC RADIUS (LANTHANIDE CONTRACTION)

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ALKALINE EARTH METALS

ATOMIC RADIUS

Atomid radius (in Å)

Ba(2.22)
Sr(2.15)
Ca(1.97)
Mg(1.6)
Be(1.12)

IONISATION ENERGY

I.E (in eV)

Be(9.32)
Mg(7.64)
Ca(6.11)
Sr(5.7)
Ba(5.2)

ELECTRONEGATIVITY

Be(1.47)
Mg(1.23)
Ca(1.04)
Sr(1)
Ba(0.97)

DENSITY

Density (gm/ml)

Ba(3.62)
Sr(2.63)
Be(1.84)
Mg(1.74)
Ca(1.55)

MELTING POINT

Melting point (°C)

Be(1277)
Ca(838)
Sr(763)
Mg(650)
Ba(714)

BOILING POINT

Boiling point (°C)

Be(2770)
Ca(1440)
Ba(640)
Mg(1107)
Sr(1380)

ELEMENT OF GROUP IIIA

ATOMIC RADIUS

Atomic radius in (Å)

Tl(1.55)
In(1.5)
Al(1.24)
Ga(1.24)
B(0.8)

IONISATION ENERGY

I. E. in KJ/mol

B(801)
Ga(579)
Al(577)
Tl(589)
In(558)

ELETRONEGATIVITY

B(2)
Tl(1.8)
In(1.7)
Ga(1.6)
Al(1.5)

DENSITY

Density (mg/ml)

Tl(11.9)
In(7.3)
Ga(5.9)
Al(5.9)
B(2.4)

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BOILING POINT

B(2550)
Al(2440)
Ga(2240)
In(2050)
T*l*(1470)

Boiling point (°C)

MELTING POINT

B(2300)
Al(660)
T*l*(303)
In(157)
Ga(30)

Melting point (°C)

NITROGEN FAMILY (PNICOGENS)

ATOMIC RADIUS

Atomic radius (in A)

Bi(1.52)
Sb(1.41)
As(1.21)
P(1.1)
N(0.74)

IONISATION ENERGY

Ionisation Energy in Ccal/mol

N(336)
P(253)
As(231)
Sb(199)
Bi(184.9)

ELECTRONEGATIVITY

N(3)
P(2.1)
As(2)
Sb(1.9)
Bi(1.9)

DENSITY

Density (gm/ml)

Bi(9.78)
Sb(6.62)
As(5.73)
P(1.823)
N(0.809)

CHALCOGENS

ATOMIC RADIUS

Atomid radius (in Å)

Te(1.37)
Se(1.14)
S(1.04)
O(0.74)

IONISATION ENERGY

Ionisation potential in eV

O(13.6)
S(10.4)
Se(9.75)
Te(9.01)
Po(8.43)

ELECTRONEGATIVITY

O(3.5)
S(2.5)
Se(2.4)
Te(2.1)
Po(2)

MELTING POINT

Melting point (°C)

Te(450)
Se(217)
S(119)
O(-219)

BOILING POINT

Boiling point (°C)

Te(990)
Se(688)
S(445)
O(-183)

HALOGENS

ATOMIC RADIUS

Atomic radius (Å)

I(1.33)
Br(1.14)
Cl(0.99)
F(0.72)

IONISATION ENERGY

Ionisation energy (in eV)

F(17.4)
Cl(13)
Br(11.8)
I(10.4)

ELECTRONEGATIVITY

F(4)
Cl(3)
Br(2.8)
I(2.5)

ELECTRO AFFINITY

Electron Affinity (eV)

Cl(-3.8)
Br(-3.5)
F(-3.6)
I(-3.2)

BOND ENERGY

Bond dissociation energy in Kcal mol

Cl-Cl (57)
Br-Br(45.5)
F-F(38)
I-I(35.6)

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3-D SERIES

ATOMIC RADIUS

DENSITY

IMPORTANT POINTS TO REMEMBER

1. The basis of Mendeeleev's periodic table was his periodic law. According to it the physical and chemical properties of the elements are a periodic function of their atomic masses.

2. According to Moseley, that a plot of $\sqrt{v}$ (where v is frequency of X-Rays emitted) against atomic number (Z) gave a straight line and not the plot of $\sim$ vs atomic mass. Therefore, he concluded that atomic number (Z) instead of atomic mass was a better fundamental property of an element and atomic number instead of atomic mass should be basis of the classification of the elements.

3. Long form of periodic table contains seven periods (horizontal rows) and eighteen groups (vertical columns).

4. In the modern periodic table, the period indicates the value of principle quantum number.

5. The number of elements in each period is twice the number of atomic orbitals available in the energy level that is being filled.

6. Group consists of a series of elements having similar valence shell electronic configuration.

7. In modern periodic table each block (s, - p, - d - and f -) contains a number of columns equal to the numbers of electrons that can occupy that sub-shell.

8. The 4f - (i.e., actinides) and 5f - (i.e., lanthanides) inner transition series of element are placed separately in the periodic table to maintain its structure and to preserve the principle of classification.

9. Metals comprise more than 78% of all known elements and appear on the left side of the periodic table. Silicon, germanium, arsenic, antimony, selenium and tellurium all are semi - metals or metalloids.

10. The combined effect of attractive force due to nucleus and repulsive force due to intervening electrons, acting on the valence electrons is that the valence - electron experiences less attraction from the nucleus. This is called shielding or screening effect.

11. Covalent radius < metallic or crystal radius < Van der Waal's radius.

12. Species having same number of electrons but different in the magnitude of their nuclear charges are called as isoelectronic species and their size is inversely proportional to their effective nuclear charge.

13.   The smaller the ionisation energy, the easier it is for the neutarl atom to change in to a positive ions in gaseous state. $IE_1 < IE_2 < IE_3$ ....•

14.   The greatest increase in ionization enthalpy is experienced on removal of electrons from core noble gas configuration. End of valence electrons is marked by a big jump in ionization enthalpy.

15.   Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion:

$$x(g) + e^- \longrightarrow X^-(g); \Delta H = \Delta_{eg}H$$
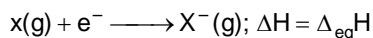
16.   The negative value of electron gain enthalpy of $C\ell > F$ (similarly S > 0) because there is weak electron-electron repulsion in the bigger 3p subshell of $C\ell$ as compared to compact 2p - subshell of F.

17.   Noble gases have larger positive electron gain enthalpies because the electron has to enter the next higher energy level.

18.   Addition of 2nd electron to an anion is opposed due to electrostatic repulsion and thus requires the absorption of energy e.g in case of the formation of S^{2-}, O^{2-} etc.

19.   The relative reactivity of metals increases with the decrease in their ionisation energies. Similarly the relative reactivity of non-metals increases with increases in the negative value of electron gain enthalpy.

20.   According to Pauling, the electronegativity difference between two atoms is equal to $0.208 \sqrt{\Delta}$ where L\ is the extra bond energies in K cal mol^{-1}. The acidic character of oxides increases when electronegativity difference decreases between element and oxygen (E - O).

21.

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