02 Periodic Table

Introduction:

The arrangement of all the known elements according to their properties in such a way that the elements with similar properties are grouped together in a tabular form is called periodic table.

Development of periodic table:

(A) Lavoisier Classification:

(i) Lavoisier classified the elements simply in metals and non metals.
Metals are the ones which have the tendency of losing electrons.

$Na \rightarrow Na^{\oplus} + e^{\ominus}$ and $K \rightarrow K^{\oplus} + e^{\ominus}$

Non-metals are the ones which have the tendency of gaining electrons.

$F + e^{\ominus} \rightarrow F^{\ominus}$ and $Cl + e^{\ominus} \rightarrow Cl^{\ominus}$

(ii) Drawback or Limitation :

(a) As the number of elements increases, this classification became insufficient for the study of elements.

(b) There are few elements which have the properties of both metals as well as non-metals and they are called metalloids. Lavoisier could not decide where to place the metalloids.

(B) Dobereiner's triad rule [1817]

(i) He made groups of three elements having similar chemical properties called TRIAD.

(ii) In Dobereiner triad, atomic weight of middle element is nearly equal to the average atomic weight of first and third element.

	Element	At. Weight	Mean weight of first and last element
(i)	Li	7	
	Na	**23**	$\frac{7+39}{2} = 23$
	K	39	
(ii)	Ca	40	
	Sr	**88**	$\frac{40+137}{2} = 88.5$
	Ba	137	
(iii)	Cl	35.5	
	Br	**80**	$\frac{35.5+127}{2} = 81.25$
	I	127	
(iv)	S	23	
	Se	**79**	$\frac{32+127.6}{2} = 79.8$
	Te	127.6	

(iii) Other examples –(K, Rb, Cs), (P, As, Sb), (H, F, Cl), (Sc, Y, La)

- As only few elements were discovered at the time so only few such triads were available.
- Day by day as new elements are discovered, the rule could no longer be generalized.

Practice Question

Q. The law of triads is applicable to -

(1) C, N, O (2) H, O, N (3) Na, K, Rb (4) Cl, Br, I Ans. (4)

Q. The law of triads is not applicable on -

(1) Cl, Br, I (2) Na, K, Rb (3) S, Se, Te (4) Ca, Sr, Ba Ans. (2)

(C) Newland octave rule [1865]

(i) He arranged the elements in the increasing order of their atomic mass and observed that properties of every 8th element was similar to the 1stone, like in the case of musical vowels notation.

Sa	Re	Ga	Ma	Pa	Dha	Ne	Sa
1	2	3	4	5	6	7	8

(ii) At that time inert gases were not known.

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

(iii) The properties of Li are similar to 8th element i.e. Na, Be are similar to Mg and so on.

Drawback or Limitation :

(a) This rule is valid only upto Ca because after Ca due to presence of d-block elements there is difference of 18 elements instead of 8 elements.

(b) After the discovery of Inert gas which when included in the periodic table becomes the 8th element from Alkali metal so this law had to be dropped.

(c) This law found to be applicable only till calcium. It was not applicable to elements of higher atomic masses.

Practice Question

Q. Which of the following set of elements obeyes Newland's octave rule –

(1) Na, K, Rb (2) F, Cl, Br (3) Be, Mg, Ca (4) B, Al, Ga Ans. (3)

Q. For which of the pair Newland octave rule is not applicable –

(1) Li, Na (2) C, Si (3) Mg, Ca (4) Cl, Br Ans. (4)

(D) Lothar Meyer's curve [1869]

(i) He plotted a curve between atomic wt. and atomic volume of different elements.

(ii) The following observation can be made from the curve –

(a) Most electropositive elements i.e. alkali metals (Li, Na, K, Rb, Cs etc.) occupy the peak positions on the curve.

(b) Less electropositive i.e. alkaline earth metals (Be, Mg, Ca, Sr, Ba) occupy the descending position on the curve.

(c) Metalloids (B, Si, As, Te, At etc.) and transition metals occupy bottom part of the curve.

(d) Most electronegative i.e. halogens (F, Cl, Br, I) occupy the ascending position on the curve.

Note : Elements having similar properties occupy similar position on the curve.

Conclusion :

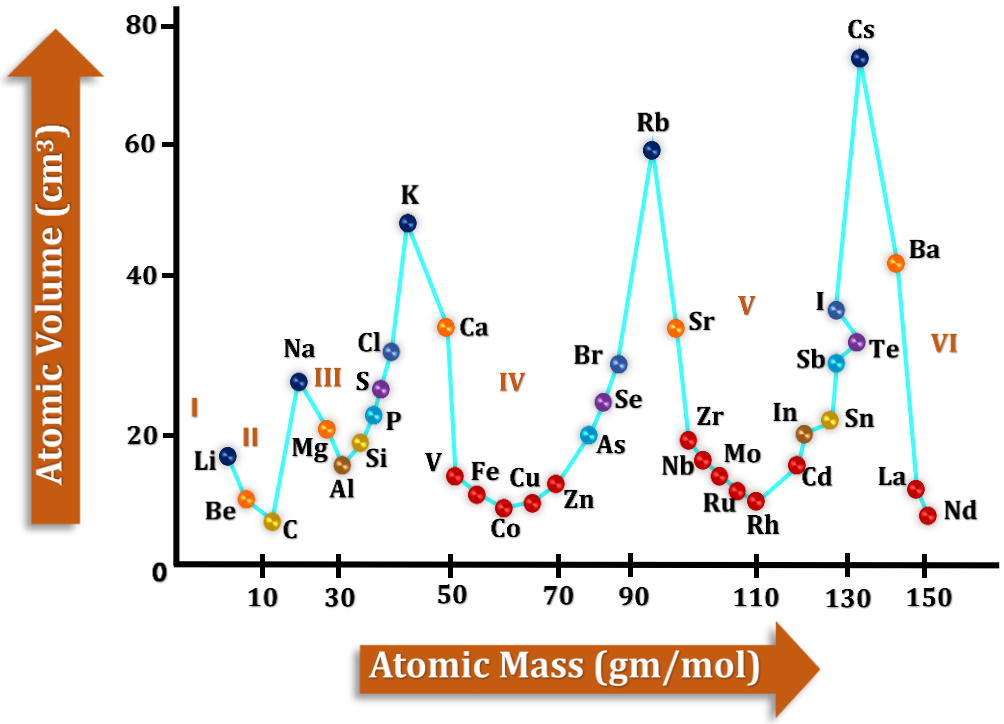
On the basis of this curve Lothar Meyer proposed that the physical properties of the elements are periodic function of their atomic wt. and this become the base of Mendeleev's periodic table.

Illustration 1:

What is periodic function ?

Solution:

When the elements are arranged in the increasing order of their at. wt., elements having similar properties gets repeated after a regular interval.



(E) Mendeleev's Periodic Table [1869]:

(i) Mendeleev's periodic law – The physical and chemical properties of elements are the periodic function of their atomic weight.

(ii) Characteristic of Mendeleev's periodic table –

(a) It is based on atomic weight

(b) 63 elements were known; noble gases were not discovered.

(c) He was the first scientist to classify the elements in a systematic manner i.e. in horizontal rows and in vertical columns.

(d) Horizontal rows are called periods and there were 7 periods in Mendeleev's Periodic table.

(e) Vertical columns are called groups and there were 8 groups in Mendeleev's Periodic table.

(f) Each group upto VIIth is divided into A & B subgroups. 'A' sub group elements are called normal elements and 'B' sub group elements are called transition elements.

(g) The VIIIth group consists of 9 elements in three rows.

(h) The elements belonging to same group exhibit similar properties.

(iii) Merits or advantages of Mendeleev's periodic table –

(a) Study of elements – For the first time, all known elements were classified in groups according to their similar properties. So study of the properties of elements became easier.

(b) Prediction of new elements – It gave encouragement to the discovery of new elements as some gaps were left in it.

Sc (Scandium), Ga (Gallium), Ge (Germanium), Tc (Technetium)
were the elements whose position and properties were well defined by Mendeleev even before their discoveries and he left the blank spaces for them in his table.
e.g. - Blank space at atomic weight 72 in silicon group was called Eka silicon (means properties like silicon) and element (discovered later) was named Germanium .
Similarly other elements discovered after Mendeleev periodic table were :
Eka aluminium – Gallium(Ga)
Eka Boron – Scandium (Sc)
Eka Silicon – Germanium (Ge)
Eka Manganese – Technetium (Tc)

(c) Correction of doubtful atomic weights – Corrections were done in atomic weight of some elements.

Atomic Weight = Valency × Equivalent weight

Initially, it was found that equivalent weight of Be is 4.5 and it is trivalent (V = 3), so the weight of Be was 13.5 and there is no space in Mendeleev's table for this element. So, after correction, it was found that Be is actually divalent (V = 2). So, the weight of Be became 2 × 4.5 = 9 and there was a space between Li and B for this element in Mendeleev's table.
Corrections were done in atomic weight of elements are – U, Be, In, Au, Pt.

(iv) Demerits of Mendeleev's periodic table –

(a) Position of hydrogen – Hydrogen resembles both, the alkali metals (IA) and the halogens (VIIA) in properties so Mendeleev could not decide where to place it.

(b) Position of isotopes – As atomic weight of isotopes differs, they should have been placed in different position in Mendeleev's periodic table. But there was no such place for isotopes in Mendeleev's table.

(c) Anomalous pairs of elements – There were some pair of elements which did not follow the increasing order of atomic weight.
eg : Ar and Co were placed before K and Ni respectively in the periodic table, but having higher atomic weights.

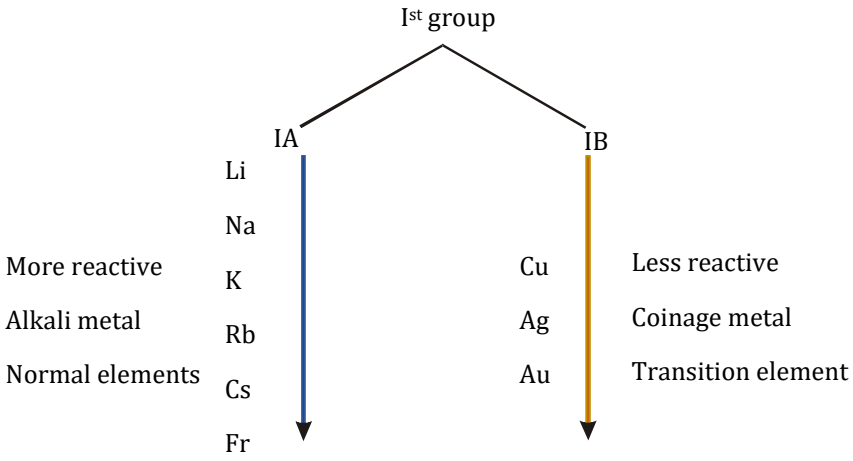
$\begin{pmatrix} Ar & K \\ 39.9 & 39.1 \end{pmatrix}$ $\begin{pmatrix} Te & I \\ 127.5 & 127 \end{pmatrix}$

$\begin{pmatrix} Th & Pa \\ 232 & 231 \end{pmatrix}$ $\begin{pmatrix} Co & Ni \\ 58.9 & 58.6 \end{pmatrix}$

(d) Like elements were placed in different groups.
There were some elements like Platinum (Pt) and Gold (Au) which have similar properties but were placed in different groups in Mendeleev's table.

Pt	Au
VIII	IB

(e) Unlike elements were placed in same group.



Cu, Ag and Au placed in Ist group along with Na, K etc. While they differ in their properties (Only similar in having ns^1 electronic configuration)

(f) It was not clear that 'lanthanides and Actinides' are related with IIIA group or IIIB group.

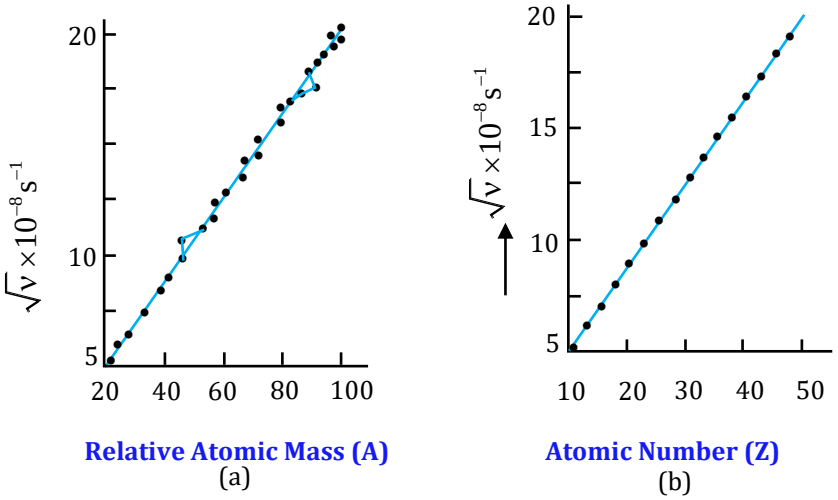
(g) Cause of periodicity : Why physical & chemical properties repeated in a group.

(F) Modern periodic table (Modified Mendeleev's Periodic Table):

(i) It was proposed by Moseley (1913).

(ii) Modern periodic table is based on atomic number.

(iii) Moseley did an experiment in which he bombarded high speed electron on different metal surfaces and obtained characteristic X-rays.



Moseley's Experiment

He found out that $\sqrt{\nu} \propto Z$ (where ν = frequency) of characteristic X-rays from this experiment, Moseley concluded that the physical and chemical properties of the elements are periodic function of their atomic number. It means that when the elements are arranged in the increasing order of their atomic number, elements having similar properties gets repeated after a regular interval. This is also known as 'Modern Periodic Law'.

(iv) Modern periodic law – The physical & chemical properties of elements are a periodic function of their atomic number.

(v) Characteristics of modern periodic table –

(a) 9 vertical columns called groups.

(b) IA to VIIA, IB to VIIB, VIII and 0

(c) Inert gases were introduced in periodic table by Ramsay.

(d) 7 horizontal series called periods.

(G) Long form / Present form of Modern Periodic Table:

(It is also called as 'Bohr-Bury & Rang-Werner Periodic Table.)

(i) It is based on the Bohr-Bury electronic configuration concept and atomic number.

(ii) This model was proposed by Rang & Werner

(iii) It consists of 7 horizontal periods and 18 vertical columns (groups)

(iv) According to I. U. P. A. C. 18 vertical columns are named as 1st to 18th group.

(v) The co-relation between the groups in long form of periodic table and in modern periodic table are given below.

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

(vi) Elements belonging to same group have same no. of electrons in the outermost shell so their properties are similar.

Description of periods:

Period	n	Sub shell	No. of elements	Element	Name of Period
1.	1	1s	2	$_{1}H$ – $_{2}He$	Shortest
2.	2	2s, 2p	8	$_{3}Li$ – $_{10}Ne$	Short
3.	3	3s, 3p	8	$_{11}Na$ – $_{18}Ar$	Short
4.	4	4s, 3d, 4p	18	$_{19}K$ – $_{36}Kr$	Long
5.	5	5s, 4d, 5p	18	$_{37}Rb$ – $_{58}Xe$	Long
6.	6	6s, 4f, 5d, 6p	32	$_{55}Cs$ – $_{86}Rn$	Longest
7.	7	7s, 5f, 6d, 7p	32	$_{87}Fr$ – $_{118}Og$	Incomplete

Conclusion

1. Period number = outermost shell
2. Number of element in a period = Number of electrons in a period subshell

Description of Groups:

1st/IA/Alkali metals

H = $1s^1$

Li = $1s^2$, $2s^1$

Na = $1s^2$, $2s^2$ $2p^6$, $3s^1$

K = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^6$, $4s^1$

General electronic configuration = ns^1

Number of valence shell $e^{\ominus}$ = 1

2nd/IIA/Alkaline earth metals

Be = $1s^2$, $2s^2$

Mg = $1s^2$, $2s^2$ $2p^6$, $3s^2$

Ca = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^6$, $4s^2$

General electronic configuration = ns^2

(n = Number of shell)

Number of valence shell $e^{\ominus}$ = 2

13th/IIIA/Boron Family

B = $1s^2$, $2s^2$ $2p^1$

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

Ga = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^6$ $3d^{10}$, $4s^2$ $4p^1$

General electronic configuration = ns^2 np^1

Number of valence shell $e^{\ominus}$ = 3

14th/IVA/Carbon Family

C = $1s^2$, $2s^2$ $2p^2$

Si = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^2$

Ge = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^6$ $3d^{10}$, $4s^2$ $4p^2$

General electronic configuration = ns^2 np^2

Number of valence $e^{\ominus}$ = 4

15th/VA/Nitrogen family/Pnictogen

(Used in fertilizer as urea)

N = $1s^2$, $2s^2$ $2p^3$

P = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^3$

As = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^6$ $3d^{10}$, $4s^2$ $4p^3$

General electronic configuration = ns^2 np^3

Number of valence shell $e^{\ominus}$ = 5

16th/VIA/Oxygen family/Chalcogen

(Ore forming)

O = $1s^2$, $2s^2$ $2p^4$

S = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^4$

Se = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^6$ $3d^{10}$, $4s^2$ $4p^4$

General electronic configuration = ns^2 np^4

Number of valence shell $e^{\ominus}$ = 6

17th/VIIA/Fluorine family/Halogens
(Salt forming)
F = $1s^2$, $2s^2$ $2p^5$
Cl = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^5$
Br = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^6$ $3d^{10}$, $4s^2$ $4p^5$
General electronic configuration = ns^2 np^5
Number of valence shell $e^{\ominus}$ = 7

18th/Zero group/Inert gases / Noble gases
(Less reactive)
Ne = $1s^2$, $2s^2$ $2p^6$
Kr = $1s^2$, $2s^2$ $2p^6$, $3s^2$ $3p^6$ $3d^{10}$, $4s^2$ $4p^6$
General electronic configuration = ns^2 np^6 (except He)
Number of valence shell $e^{\ominus}$ = 8

Metalss, non-Metals & Metalloids:

Apart from classifying elements into s, p, d and f-blocks, there is yet another broad classification of elements based on their properties. The elements can be broadly classified into

(a) Metals:

Majority of the elements in periodic table are metals and appears on the left side of the periodic table.

Properties:

(i) These are usually solid at room temperature [exception - mercury]
(ii) They have high melting and boiling point [exception Gallium & Caesium have very low melting point (303 K and 302 K respectively)]
(iii) They are good conductor of heat and electricity.
(iv) They are malleable (can be flattened into thin sheets by hammering) and ductile (can be drawn into wires)

(b) Non-Metals:

These are placed at the top right hand side of periodic table. As we move horizontally along a period, the property of elements changes from metallic (on left) to non-metallic (on the right).

Properties :

(i) These are usually solids or gases at room temperature.
(ii) They have low melting point and boiling point (exception : Boron, Carbon).
(iii) Most Non-metallic solids are brittle and are neither malleable nor ductile.

(c) Metalloids (Semi-metals) :

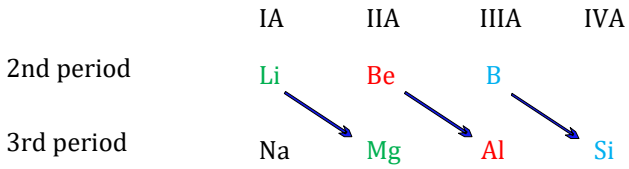
Properties of these elements show the characteristics of both metals and non-metals. Silicon (Si), Germanium(Ge), Arsenic(As), Antimony(Sb) and Tellurium(Te) are metalloids.

Estimating Position of an Element from its Electronic Configuration:

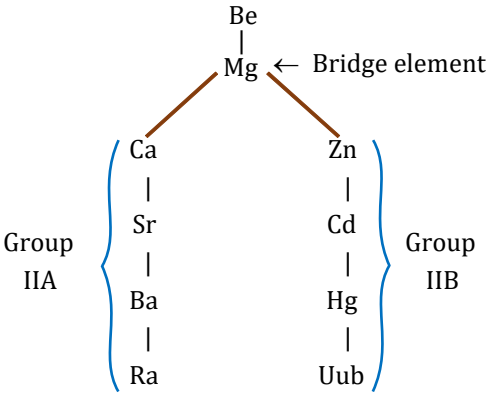
The last electron enters in which subshell gives idea of its block.
[Think :$1s^1$ and $1s^2$ belongs to which block]
Period number = Principal quantum number of valence shell electron in ground state electronic configuration.
Group number for s block = number of valence shell electrons
Group number for p block = 10 + number of valence shell electrons
Group number for d block = number of [ns + (n–1) d] electrons
Group number for f-block = 3
[Use these carefully while locating the position.]

Some Commonly Used Terms:

1. **Noble Gases :** Element of group 18 are called noble gases. These are also called as inert gases because their outermost ns and np orbitals are completely filled (except He has $1s^2$) and these gases are non-reactive in nature under ordinary conditions.
2. **Typical elements :** Elements of third period are known as typical elements.
3. **Diagonal relationship :** Properties of elements of second period resemble with the element of third period. This resemblance between properties of 2nd & 3rd period is called diagonal relationship.



4. **Bridge elements:**
The typical elements of third period are also called bridge elements as the division between two subgroups A and B starts from these elements. In second group Mg acts as a bridge element. The properties of bridge element are some what mixed of the elements of two subgroups as magnesium shows similarities with alkaline earth metals (IIA) on one hand and with zinc metals (IIB) on the other.



Q. Which of the following elements belong to alkali metals ?
(1) $1s^2, 2s^2, 2p^2$ (2) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^1$
(3) $1s^2, 2s^2, 2p^5$ (4) None of these Ans.(2)

Important Points:

(a) 2nd period elements (Li, Be, B) shows diagonal relationship with 3rd period elements (Mg, Al, Si) Because of same ionic potential (Ionic potential = Charge/Radius) value they shows similarity in properties.

(b) 3rd period elements (Na, Mg, Al, Si, P, S, Cl) are called typical elements because they represent the properties of other element of their respective group.
(c) In 6th period all types of elements are included (s, p, d & f)
(d) No inert gas in 7th period.

(e) Normal elements present in all periods.
(f) Atomic No. of last inert gas element is 86.
(g) Long form of modern periodic table can be divided into four portions.
1. Left portion – IA & IIA – s - block.
2. Right portion – IIIA to VII A + 0 group – p - block.
3. Middle portion – IIIB to VIIB + VIII + IB IIB – d - block.
4. Bottom portion – IIIB – f - block elements
(h) The group containing most electro positive elements – IA GROUP .
(i) The group containing most electro negative elements – VIIA GROUP
(j) The group containing maximum number of gaseous elements – ZERO(18^{th}) GROUP
(k) The group in which elements have generally ZERO valency – (18^{th}) GROUP ZERO
(l) In the periodic table –
Number of Gaseous elements – 11 (H, N, O, F, Cl + Noble gases)
Number of Liquid elements – 6 (Cs, Fr, Ga, Hg, Br, Uub)
Bromine is the only non-metal which exists in liquid form(at room temperature).
Number of Solid elements – 95 (if discovered elements are 112)
(m) No p-block elements in 1^{st} and 7^{th} periods.
(n) Ist period has all the elements in gaseous form (H, He)
(o) 0/18 group have all the elements in gaseous form.
(p) 2nd period contains maximum number of gaseous elements. They are 4 : N, O, F, Ne
(q) IIIB/3rd group is called largest group having 32 elements including 14 Lanthanides and 14 Actinides
Sc
Y
La.....Lanthanides (14)
Ac.....Actinides (14)

Q. Properties of Li is similar to :
(1) Mg (2) Na (3) Both (4) None of these Ans.(3)
Q. 6^{th} typical element is .....? Ans. Sulphur
Q. Lanthanides belong to which group of the table ?
(1) IIIA (2) IIIB (3) 3 (4) both 2 and 3 Ans.(4)

- **Merits of long form of periodic table –**
(a) Position of isotopes – Atomic No. of isotopes are similar, so different isotopes can be placed at same place in periodic table.
(b) (Ar – K) (Co – Ni) (Te – I) are now in increasing order of atomic number.
(c) Lanthanides and actinides are in IIIB group or group 3.
(d) In modern periodic table diagonal line separates out metals, metalloids and non metals.
(e) Elements of same group have same general formula of electronic configuration of outer most shell.

- **Demerits of long form of periodic table –**
(a) Position of hydrogen is still controversial.
(b) 'He' is a inert gas but it has different electronic configuration than other inert gas elements.
(c) Lanthanides and actinides are still not placed in main frame.
(d) Isotopes have different physical properties but have same place in periodic table.

Practice Question

Q. Each period in the periodic table starts with a subshell of new shell and ends with –
(1) Small subshell (2) Next higher shell Ans.(3)
(3) p-subshell of the same shell (4) Different subshell of the same subshell

Q. The number of elements in long period is –
(1) 18 (2) 8 (3) 32 (4) 10 Ans.(1)

Q. Long form of the periodic table is based on –
(1) Atomic size (2) Atomic number (3) Atomic mass (4) Atomic volume Ans.(2)

Q. Maximum work on periodic table related with atomic number was given by –
(1) Moseley (2) Mendeleev (3) Aston (4) Hund Ans.(1)

Classification of elements:

(A) Neil Bohr's Classification – Elements can be divided in four parts on the basis of electronic configuration.

(i) Inert gas elements –

(a) The elements in which outermost shell or ultimate shell is completely filled up are called inert gas elements.
(b) General electronic configuration is ns^2p^6 (Except He = $1s^2$)
(c) Because of the most stable configuration, they are very less reactive. Hence known as noble gas or inert gas.
(d) These elements are present in '0' group or 18^{th} group and 1 to 6^{th} period of periodic table.
(e) Number of inert gas elements are 6(One in each period upto 6^{th}) He, Ne, Ar, Kr, Xe, Rn

(ii) Normal or Representative elements –

(a) The elements in which outermost shell or ultimate shell is incomplete, and no. of electrons in the ultimate shell is less than eight while penultimate shell is complete are called normal or representative elements.

Or

The elements of s-block & p-block except inert gases are called normal or representative elements.

(b) Their general electronic configuration is :

IA	IIA	IIIA	IVA	VA	VIA	VIIA
ns^1	ns^2	$ns^2\ np^1$	$ns^2\ np^2$	$ns^2\ np^3$	ns^2np^4	$ns^2\ np^5$
ns^{1-2}		$ns^2\ np^{1-5}$				

(c) These elements lies in group IA to VIIA and period I^{st} to 7^{th}
(d) Elements of 2^{nd} period known as bridge elements.
(e) Elements of 3^{rd} period (Na, Mg, Al, Si, P, S and Cl) are called typical elements These are 7 in numbers.

(iii)Transition elements –

Old concept (Given by Mendeleev's) : They are called transition elements because their properties are intermediate of s and p- block elements they transit the properties of metals to non metals.

Note :

According to this concept, Zn, Cd, Hg and Uub are transition elements.

New concept (Given by Bohr) : The elements in which both ultimate (n) as well as penultimate shells (n – 1) are incomplete either in ground state or in any stable oxidation state are called transition elements.

Note :

According to this concept, Zn, Cd, Hg and Uub are not transition elements because they do not have incomplete penultimate shell either in ground state or in stable +2 oxidation state.

Zn = [Ar] $3d^{10}\ 4s^2$

$Zn^{2\oplus}$ = [Ar] $3d^{10}$

(a) group – IIIB to VIIB + VIII + IB + IIB
Periods – 4^{th} to 7^{th}

(b) Electronic configuration $(n-1)d^{1-10}\, ns^{1 \text{ or } 2}$

(c) Total number of d-block elements = 40
Total number of transition elements = 36 (Except Zn, Cd, Hg and Uub)

Note :

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

(iv) Inner transition elements –

(a) The Elements in which all the three shells that is ultimate (n) penultimate (n–1) and pre or antepenultimate (n – 2) shell are incomplete are called inner transition elements.

$_{58}Ce$ = [Xe] $6s^2$, $5d^1$, $4f^1$

(b) Electronic configuration –

$(n-2)f^{1-14}\,(n-1)\,d^{0 \text{ or } 1}\, ns^2$

(c) The last shell contains 2 electrons, penultimate shell may contain or 8 or 9 electrons and pre or antepenultimate shell contains more than 18 electrons and up to 32 electrons.

(d) These are 28 in number.

(e) Group – IIIB

(f) Period – 6^{th} & 7^{th}

(g) Inner transition elements are divided into two series.

(A) Lanthanide series/Rare earth elements /Lanthanoids (Ce_{58} – Lu_{71} 14 elements)
The first element of this series is Cerium & not Lanthanum.
In these elements, last electron enters into 4f subshell
They are present in IIIB group and 6^{th} period of the periodic table.
Promethium ($_{61}Pm$) is the only lanthanide which is synthetic and radioactive in nature.

(B) Actinide series /Man made elements /Actinoids (Th_{90} – Lw_{103} 14 elements)
The first element of this series is Thorium & not Actinium.
In these elements, last electron enters into 5f subshell.
They are present in IIIB group and 7^{th} period of the periodic table.
All the actinides are radioactive in nature.
First three elements (Th, Pa, U) are found in nature while others are synthetic in nature.
After U_{92} i.e. from $_{93}Np$ onwards elements are called transuranic elements because :
They are heavier than uranium.
They are derived from uranium by nuclear reactions.

Summary of Neil Bohr's Classification

Inert gases	Normal or representative elements	Transition element	Inner transition element representative
outermost shell complete	outermost shell incomplete	n & n–1 shells incomplete either in atomic or ionic form	n, (n–1), (n–2) shells incomplete
6 element	s & p block element except inert gas upto 118 (Atomic no.)	all d block element except = IIB (Zn, Cd, Hg & Uub) 36 element	f-block elements 28 elements

Illustration 2:

Classify the following elements based on Bohr's concept :

(1) [Ar] $4s^2$, $3d^8$ (2) $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$ (3) [Xe] $6s^2$, $5d^1$, $4f^{12}$

(4) [Rn] $7s^2$, $6d^4$, $5f^{14}$ (5) [Kr] $5s^2$, $4d^{10}$, $5p^2$

Solution:

(1) Transition (2) Inert gas (3) Inner transition

(4) Transition (5) Normal element

Illustration 3:

The heaviest naturally occurring element is .....?

Solution:

Uranium.

Classification on the Basis of Sub shell in which last $e^{\ominus}$ enters:

(i) s - block elements –

(a) In these elements last electron enters in s - subshell.

(b) Group – IA(1) and IIA(2) group

(c) Period – 1 to 7^{th}

(d) Electronic configuration – ns^{1-2}

IA group n = 1 to 7

IIA group n = 2 to 7

(e) Total s - block elements are (14) (including H and He)

(f) Total s - block metals are (12) (excluding H)

(g) He is a s-block element as last e^- enters in s-subshell.

General characteristics of s-block elements :

(a) s-block elements are soft metal, that is why these elements have low melting and boiling point.

(b) Their oxides and hydroxides are basic in nature.

(c) They act as reducing agent as they have tendency to donate $e^{\ominus}$.

(d) They have very low ionisation energy and highly electropositive so, they form ionic compound.

(e) Flame of these elements have the property of showing different colour.

(ii) p - block elements –

(a) Last e^- enters in p - sub shell

(b) Group – IIIA (13) to VIIA (17) + 0 group (18) (except He)

(c) He is a s-block element but placed in p-block zero or 18^{th} group as it is a noble gas element.

(d) Period – 2^{nd} to 7^{th}

(e) Electronic configuration – $ns^2\ np^{1-6}$

(f) Total p - block elements – (36)

Characteristic of p- block elements :

(a) p-block possesses all three kind of elements i.e. metals, non-metals and metalloids.

(b) Oxides of non-metals are acidic in nature.

(c) They form covalent compounds mostly.

(d) They are oxidising in nature.

(iii) d - block elements –

(a) Last e^- enters in (n – 1) d subshell
(b) Group – IIIB – VIIB, VIII, IB, IIB or group 3 to 12 (IUPAC)
(c) Period – 4^{th} to 7^{th}
(d) Electronic configuration – (n –1) d^{1-10} $ns^{1 \text{ or } 2}$
(e) Total d - block elements – (40) If 112 elements are included in periodic table
Total transition elements – (36)
(f) IIB elements (Zn, Cd, Hg, Uub) are d-block elements but not transition elements.

General characteristics of d - block elements –

(a) They all are metals, which are very hard having high melting and boiling point.
(b) Elements of this block situated in between 's' and p-block elements. So the character of this block elements lie between s and p-block elements
(c) They show variable oxidation - state. eg. Mn = $Mn^{2\oplus}$, $Mn^{3\oplus}$, $Mn^{4\oplus}$, $Mn^{5\oplus}$, $Mn^{6\oplus}$, $Mn^{7\oplus}$
(d) They form lonic and covalent bond both.
(e) They are good conductor of heat and electricity and form complex compounds.
(f) Metals, which have unpaired electrons show para-magnetism.
(g) They form alloys and most of the elements act as catalyst.

(iv) f - block elements –

(a) Differentiating e^- enters in (n – 2) f subshell.
(b) Group – IIIB
(c) Period – 6^{th} and 7^{th}
(d) From at. no. 58 - 71, 6^{th} period ; Lanthanide series $4f^{1-14}$ $5d^{0 \text{ or } 1}$ $6s^2$
90 – 103 l 7^{th} period ; Actinide series $5f^{1-14}$ $6d^{0-2}$ $7s^2$
(e) Total number of f block elements – (28)
(f) All the actinides are radioactive elements.
(g) Transuranic actinides are man made elements. (Np_{93} – Lw_{103})
(h) Lanthanides are found rarely on earth so these are called rare earth metals.

General characteristics of f-block elements –

(a) All the f-block elements are heavy metals.
(b) They have high melting and boiling point.
(c) The most common oxidation state of these elements is +3.
(d) The elements of 5f series are radioactive in nature.

Determination of period, block and group of an element –

(i) Period No. – The period no. of the element can be predicted from the principal quantum no. (n) of the valence shell.
e.g. electronic configuration of iodine is :
$1s^2 2s^2 2p^6$, $3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^5$. Therefore the period number of iodine is 5 as the valence shell configuration is $5s^2 5p^5$.

(A) When atomic number is given:

Step I : $71 \geq Z \geq 58 \Rightarrow$ Lanthanoids (6^{th} Period)] f-block
$103 \geq Z \geq 90 \Rightarrow$ Actinoids (7^{th} Period)]
Group number = IIIB or Group 3 (largest group of periodic table)

Step II : Z = 104 to 118 (Period number = 7)

Group number = last two digits in atomic number of element

Example: Z = 104

Group no. = 4

Step III: Group number = 18 + given atomic number – atomic number of next noble gas

If the value of this formula is negative then use 32 instead of 18 in formula.

(B) When electronic configuration is given

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

(ii) Block No. – The type of orbital which receives the last electron known as block no.

e.g. An element 'X' has its electronic configuration as $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$. As the last electron enters in the d-orbital, therefore it is a d-block element.

Block identification:

- If np electron present then p - block ($ns^2\ np^{1-6}$)
  group number = 12 + np electrons
- If np electron absent then s/f/d block
  - If $(n-2)f^0\ (n-1)d^0\ ns^{1-2}$ = s block
    group number = ns electrons
  - If $(n-2)f^{1-14}\ (n-1)d^{0-2}\ ns^2$ = f block
    group number = IIIB
- If any other configuration or $(n-1)d^{1-10}\ ns^{0-2}$ (d-block)
  group number = (n–1)d electron + ns electron

(iii) Group No. – It is predicted from the number of electrons in the valence shell and penultimate shell as follows.

(a) For s-block elements, group number is equal to the number of electrons in the valence shell after the noble gas core.
e.g. An element 'Y' having electronic configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ or [Ar] $4s^2$ has two electron in valence shell and it is a s-block element. Therefore it belongs to group 2.

(b) For p - block element, the group number is equal to 10 + number of electrons in valence shell.
e.g. An element 'Z' with electronic configuration as $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$ has five electrons in its valence shell and belongs to p - block.
Therefore its group number is 10 + 5 = 15. It belongs to group VA of the Mendeleev's periodic table.

(c) For d-block elements group number is equal to the number of electrons in (n–1)d sub-shell and valence shell.
e.g. An elements 'A' having electronic configuration as.
$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$. so its group number will be 10 + 1 = 11.

Predicting atomic no. of the successive member in a group or family –

(i) Magic Numbers –

Knowing the at. no. of the first member of a group, we can write the at. no. of the subsequent elements by adding given magic no. For example

Group	1	2	3	4, 5, 6, 7, 8, 9,	10, 11, 12	13,	14,	15,	16,	17	18
	IA	IIA	IIIB			IIIA	IV	V	VI	VII	'0'gp
I+	2	–									8
II+	8	8						8			8
III+	8	8						18			18
IV+	18	18	18		18			18			18
V+	18	18	18		32			32			32
VI+	32	32	32		32			–			–
VII											

PERIOD

(ii) In group IA –

Atomic no. of H is 1. Atomic no. of other element will be as follows –

	H_1 1 +2 = 3	Li 3 + 8 =11	Na 11 + 8 = 19	K 19 + 18 = 37	Rb 37 + 18 = 55
Magic no.	2	8	8	18	18

Practice Question

Q In the periodic table the elements Z = 30 is expected to be present in –
(1) IA series (2) IB group (3) IIA group (4) IIB group Ans.(4)

Q Elements A, B, C, D and E have the following electronic configurations :-
(A) $1s^2, 2s^22p^1$ (B) $1s^2, 2s^22p^6, 3s^23p^1$ (C) $1s^2, 2s^22p^6, 3s^23p^3$
(D) $1s^2, 2s^22p^6, 3s^23p^5$ (E) $1s^2, 2s^22p^6, 3s^23p^6, 4s^1$
The elements which belong to the same group of the periodic table are :-
(1) (A) and (C) (2) (A) and (B) (3) (A) and (D) (4) (A) and (E) Ans.(2)

Q Uranium is a member of –
(1) Transition series (2) III-period (3) Actinide series (4) VI-period Ans.(3)

Q In modern periodic table 6^{th} period contain elements –
(1) 8 (2) 18 (3) 10 (4) 32 Ans.(4)

Nomenclature of Elements :

(a) IUPAC gave names to elements above atomic number 100 as follows –

0	1	2	3	4	5	6	7	8	9
nil	un	bi	tri	quad	pent	hex	sept	oct	enn

(b) In all the elements suffix is – ium.

Ex. Atomic No.	IUPAC Name	Symbol	Elemental Name	Symbol
101	Un nil unium	Unu	Mendelevium	Md
102	Un nil bium	Unb	Nobelium	No
103	Un nil trium	Unt	Lawrencium	Lr
104	Un nil quadium	Unq	Rutherfordium	Rf
105	Un nil pentium	Unp	Dubnium	Db
106	Un nil hexium	Unh	Seaborgium	Sg
107	Un nil septium	Uns	Bohrium	Bh
108	Un nil octium	Uno	Hassium	Hs
109	Un nil ennium	Une	Meitnerium	Mt
110	Un un nilium	Uun	Darmstadtium	Ds

Screening effect or Shielding effect (σ) and effective nuclear charge (Z_{eff}):

(i) Valence shell $e^{\ominus}$ suffer force of attraction due to nucleus and force of repulsion due to inner shell electrons.

(ii) The decrease in force of attraction on valence $e^{\ominus}$ due to inner shell $e^{\ominus}$ is called screening effect or shielding effect.(i.e. total repulsive force is called shielding effect.)

(iii) Due to screening effect valence shell $e^{\ominus}$ experiences less force of attraction exerted by nucleus.
i.e. total attraction force experienced by valence electrons represented by a number is Z_{eff}.

(iv) There is a reduction in nuclear charge due to screening effect. Reduced nuclear charge is called effective nuclear charge.

(v) If nuclear charge = Z, effective nuclear charge = Z_{eff} , σ (Sigma)= Screening constant or shielding constant. So, $\boxed{Z_{eff} = (Z - \sigma)}$

- **Slater's rule to know screening or shielding constant (σ)**

(a) For single electron species σ = 0

(b) Screening effect (S.E.) for two e^- species

Ex. In He ($1s^2$)

Screening effect of one 1s $e^{\ominus}$. where σ = 0.30

$\therefore$ Zeff = Z – σ = 2 – 0.30 = 1.7

(c) Screening effect of each ns and np (Outermost shell) electrons is 0.35

(d) Screening effect of each (n – 1) penultimate shell s, p, d electrons is 0.85

(e) Screening effect of each (n – 2) and below all the $e^{\ominus}$ present in s, p, d, f is 1.0

From top to bottom in a group Z_{eff} remain constant

(f) Screening effect of each nd & nf electron is 0.35 and for all the $e^{\ominus}$ below nd & nf is 1.

Group	Element	Li	Na	K	Rb	Cs	Fr
	Z_{eff}	1.30	2.20	2.20	2.20	2.20	2.20
Period	Element	Be	B	C	N	O	F
	Z_{eff}	1.95	2.6	3.25	3.90	4.55	5.20

- For same shell shielding effect has the order as s > p > d > f (due to penetration effect)

(i) Z_{eff} for different ions of an element

Zeff ↑ as Positive charge ↑ , and Zeff ↓ as Negative charge ↑

Ex. $N^{\oplus} > N > N^{\ominus} = Z_{eff}$

(ii) Z_{eff} for isoelectronic species (When arranged in increasing order of atomic no.)

Zeff ↑ as Atomic no. ↑

Ex. $H^{\ominus} < Li^{\oplus} < Be^{2\oplus} < B^{3\oplus}$ ($2e^{\ominus}$ species)

$N^{3\ominus} < O^{2\ominus} < F^{\ominus} < Na^{\oplus} < Mg^{2\oplus}$ ($10e^{\ominus}$ species)

Periodic Properties:

Valency : It is defined as the combining capacity of the elements. The word valency is derived from an Italian word "Valentia" which means combining capacity.

Old concept : Given by : Frankland

Valency with respect to Hydrogen and Chlorine : Valency of H=1

It is defined as the number of hydrogen or Chlorine atoms attached with a particular element.

	IA	IIA	IIIA	IVA	VA	VIA	VIIA
	NaH	MgH_2	AlH_3	SiH_4	PH_3	H_2S	H-Cl
	NaCl	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_3	Cl_2O	Cl-Cl
Valency	1	2	3	4	3	2	1

Note :

Valency w.r.t. H across the period increases upto 4 and then again decreases to 1.

Valency with respect to oxygen : **(Valency of O=2)**

It is defined as twice the number of oxygen atoms attached with a particular atom.

	IA	IIA	IIIA	IVA	VA	VIA	VIIA
	Na_2O	MgO	Al_2O_3	SiO_2	P_2O_5	SO_3	Cl_2O_7
Valency	1	2	3	4	5	6	7

Note :

Valency with respect to oxygen increases from 1 to 7 across the period. Valency w.r.t. 'O' is equal to the group number.

New concept : This concept is based on the electronic configuration. According to this concept valency for IA to IVA group elements is equal to number of valence shell $e^{\ominus}$ and from VA to zero group, it is [8– (number of valence $e^{\ominus}$)].

Valency = No. of valence $e^{\ominus}$

Valency = (8 – No. of valence $e^{\ominus}$)

	IA	IIA	IIIA	IVA	VA	VIA	VII	0
	ns^1	ns^2	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^5
Valence shell $e^{\ominus}$	1	2	3	4	5	6	7	8
Valency	1	2	3	4	3	2	1	0
					(8 – 5) = 3			(8 – 8) = 0

Periodicity:

(a) The regular gradation in properties from top to bottom in a group and from left to right in a period is called periodicity in properties.

(b) In a period, the ultimate shell remain same, but the number of electrons gradually increases.

(c) In a group, the number of electrons in the ultimate shell remains same, but the values of n increases.

Cause of Periodicity

(a) The cause of periodicity in properties is due to the same outermost shell electronic configuration repeating over regular intervals.

(b) In the periodic table, elements with similar properties occur at intervals of 2, 8, 8, 18, 18 and 32. These numbers are called magic numbers.

Atomic Radius:

(a) It is distance between outermost $e^{\ominus}$ and nucleus.

(b) Half of the nuclear distance between two atoms is defined as atomic radius.

(c) X- ray diffraction, $e^{\ominus}$ diffraction method and nuclear magnetic resonance (NMR) spectrum methods are used to determine internuclear distance or bond length.

$$\text{Atomic radius} = \frac{\text{Inter nuclear distance}}{2}$$

(d) Atomic radius depends on the type of chemical bond between atoms in a molecule. These are :

(i) Covalent radius (ii) Ionic radius (iii) Metallic radius (iv) Vander waal's radius

(i) Covalent radius : (Single Bonded Covalent Radius)

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

(b) Internuclear distance is minimum in this case.

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

e.g. If internuclear distance of A–A(A_2) molecule is (d_{A-A}) and covalent radius is r_A then

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

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

e.g. – In Cl_2 molecule, internuclear distance is 1.98Å so $r_{Cl} = \frac{1.98Å}{2} = 0.99Å$

(d) SBCR of O, N and C etc. elements can be determined by taking H_2O_2, N_2H_4, C_2H_6 respectively.

Case I –For Heteroatomic molecule with no. E.N. difference.

For A – B molecule Electronegativities of A & B are approximately equal

e.g. C – I, E.N. of C & I are approx. equal (2.5) internuclear distance of C – I is 2.13Å

$d_{C-I} = r_C + r_I$ (r_C is 0.77 Å)

$\therefore \quad r_I = 2.13 - 0.77 = 1.36Å$

Case II – Heteroatomic molecule with significant Δ E.N. difference, using Schomaker and Stevenson law.

Schomaker and Stevenson law:

If in a diatomic molecule electronegativities of A – B have more difference. Then actual bond length will be reduced.

As per Schomaker & Stevenson– The reduction in bond length depends on the difference in electronegativities of atoms by following manner -

$d_{A-B} = r_A + r_B - 0.09 (x_A - x_B)$

Here x_A is E. N. of A & x_B is E.N. of B

e.g. - If bond length of F_2 = 1.44 Å, Bond length of H_2 = 0.74Å.

Find out the bond length of H – F ? (EN of F is 4.0, EN of H is 2.1)

Solution - $d_{H-F} = r_F + r_H - 0.09 (X_F - X_H)$

$\because \quad r_F = 1.44 / 2 = 0.72$ Å, $r_H = 0.72/2 = 0.37$ Å

$\therefore \quad d_{H-F} = 0.72 + 0.37 - 0.09 (4.0 - 2.1)$

$= 1.09 - (0.09 \times 1.9) \quad = 1.09 - 0.171 = 0.919$ Å

(ii) Ionic Radius:

(A) Cationic radius:

(a) When a neutral atom loses $e^{\ominus}$ it converts into cation (+ve charged ion)

Atomic radius > Cationic radius

because after loosing $e^{\ominus}$ no. of $e^{\ominus}$ reduces, but no. of protons remains same, due to this Zeff increases, hence electrons are pulled towards nucleus and atomic radius decreases, moreover after loosing all the electrons from the outer most shell, penultimate shell becomes ultimate shell which is nearer to nucleus so size decreases.

(b) Size of cation ↑ when Magnitude of the (+)ve charge ↓ or Zeff ↓

e.g. $Fe > Fe^{2\oplus} > Fe^{3\oplus}$

$Pb^{2\oplus} > Pb^{4\oplus}$

$Mn > Mn^{2\oplus} > Mn^{3\oplus} > Mn^{4\oplus} > Mn^{5\oplus} > Mn^{6\oplus} > Mn^{7\oplus}$

(B) Anionic radius :-

(a) When neutral atom gains $e^{\ominus}$ it converts into an anion.
Anionic radius > Atomic radius

(b) In an anion $e^{\ominus}$ are more than protons so effective nuclear charge reduces, and inter electronic repulsion increases, which also increases screening effect. So distance between $e^{\ominus}$ and nucleus increases and size of anion also increases.
Size of anion ↑ when Magnitude of the (–)ve charge ↑ or Zeff ↓

e.g. Z of fluorine is 9

	F	$F^{\ominus}$
Proton	9	9
$e^{\ominus}$	9	10

so $\frac{Z}{e} = \frac{9}{9} = 1\ \frac{9}{10} = 0.9$ As Zeff of $F^{\ominus}$ is less than F so size of $F^{\ominus}$ > F

(C) Size of iso electronic species:

– Those species having same no. of $e^{\ominus}$ but different nuclear charge forms isoelectronic series.
– for isoelectronic species the atomic radius decreases with increase in nuclear charge

	$S^{2\ominus}$	$Cl^{\ominus}$	$K^{\oplus}$	$Ca^{2\oplus}$
Z	16	17	19	20
e	18	18	18	18

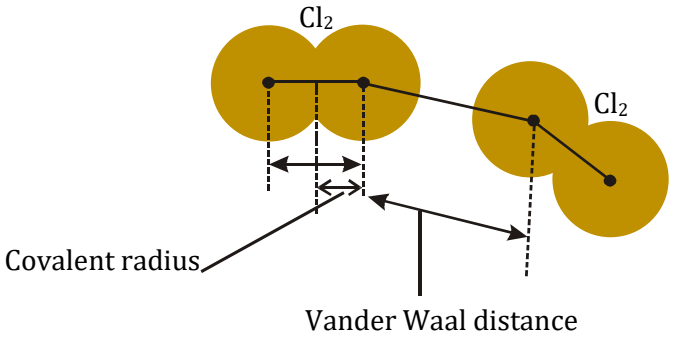
Order of radius : ($S^{2\ominus} > Cl^{\ominus} > K^{\oplus} > Ca^{2\oplus}$), ($N^{3\ominus} > O^{2\ominus} > F^{\ominus} > Na^{\oplus} > Mg^{2\oplus} > Al^{3\oplus}$)

(iii) Metallic radius :

(a) Half of the nuclear distance between two adjacent metallic atoms in crystalline lattice structure.
(b) there is no overlapping of atomic orbitals
(c) so Metallic radius > Covalent radius
(d) Metallic radius ↑ when Metallic bond strength ↓
(e) More metallic radius → loose crystal packing → less bond strength
(f) Less metallic radius → Tight crystal packing FCC → High bond strength

(iv) Vander Waal's radius:

(a) Those atoms (like noble gases) which are not bonded with each other, experiences a weak attractive force to come nearer.
(b) Half of the distance between the nuclei of adjacently placed atoms in solid state of a noble gas is Vander Waal's radius.
(c) Vander Waal radius > Covalent radius.
(d) Inert gas have only Vander Waal radius.
(e) In molecules of nonmetals both covalent and Vander Waal radius exists.
(f) VWR > Metallic R > Covalent R
(g) Vander Waal force of attraction ∝ Molecular wt. or atomic weight (in inert gases)
(h) In a period from left to right VWR decreases.
(i) In a group from top to bottom its values increases.



Cl_2 molecules
Covalent radius = 0.99Å
Vander Waal radius = 1.80Å

Vander Waal radius ≈ 2 × covalent radius

For a same element

Cationic radius < Atomic radius < Anionic radius

Covalent radius < Metallic radius < Vander Waal's radius

(v) Factors affecting atomic size are:

(a) Atomic size ↓ when Effective nuclear charge ↑	Li >Be > B > C > N > O > F
(b) Atomic size ↑ when No. of shells ↑	Li < Na < K < Rb < Cs
(c) Atomic size ↓ when Screening effect ↓	Li >Be > B > C > N > O > F
(d) Atomic size ↓ when Magnitude of +ve charge ↑	$Mn > Mn^{2\oplus} > Mn^{3\oplus} > Mn^{4\oplus}$
Atomic size ↑ when Magnitude of –ve charge ↑	$O < O^{\ominus} < O^{2\ominus}$
(e) Atomic size ↓ when Bond order ↑	N – N > –N = N– > N ≡ N

(vi) Periodic variation of atomic size:

(i) Across a period – It decreases from left to right in a period as nuclear charge increases
e.g. Li > Be > B > C > N > O > F < Ne

(ii) In a group – It increases from top to bottom in a group as number of shell increases
e.g. Li < Na < K < Rb < Cs

Exceptions : Transition elements

(a) From $_{21}Sc$ to $_{30}Zn$ in first transition series Z_{eff} continuously increases, but atomic size does not decreases continuously.

(b) From $_{21}Sc$ to $_{26}Fe$ size decreases
Nuclear charge > Screening effect

(c) From $_{26}Fe$ to $_{28}Ni$ size almost constant
Nuclear charge ~ Screening effect

(d) From $_{28}Ni$ to $_{30}Zn$ size increases. As nuclear charge is less effective than screening effect.

(vii) Lanthanide Contraction :-

(a) Outermost electronic configuration of inner transition elements is
$(n-2)\,f^{1-14}, (n-1)s^2p^6d^{0-1}, ns^2$ (n = 6 or 7)

(b) e^- enters in (n – 2) f orbitals

(c) Mutual screening effect of e^- is very less, because of complicated structure of f-orbital

(d) Nuclear charge increases by one (+1) in lanthanides and actinides so atomic size of these elements slightly decreases. It is known as lanthanide contraction.
Here Nuclear charge > Screening effect.

(e) In transition series Ist 2nd and 3rd

	III_B	IV_B	
size	Sc	Ti	
increases ↓	Y	Zr	Almost equal due to lanthanide contraction
down the groups	La	Hf	

Radii 3d < 4d ≈ 5d except IIIrd B.

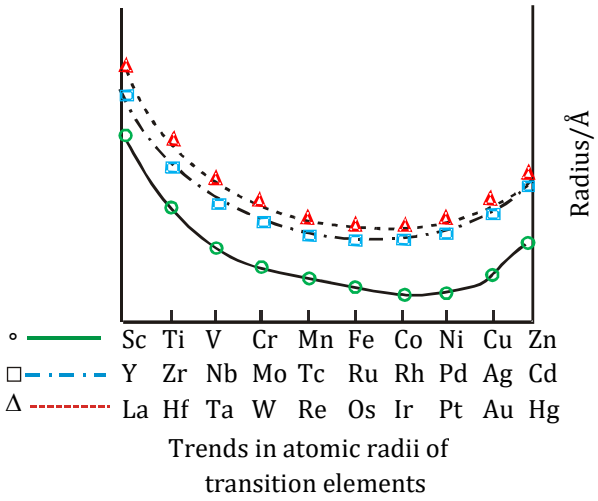
(viii) Orders of atomic and ionic radii –

$H^{\oplus} < H < H^{\ominus}$ $I^{\oplus} < I < I^{\ominus}$

Ar > Na > Cl $Na^{\oplus} < Ne < F^{\ominus}$ Be < Li < Na

Ni < Cu < Zn Sc > Ti > V > Cr < Mn > Fe Sc < Y < La

Cu < Au < Ag Fe ≥ Co ≈ Ni Ti < Zr ≈ Hf



Trends in atomic radii of transition elements

In a group :

(i) The atomic radius of elements increases on moving down the first transition series (3d) to second transition series (4d). This is due to the increases in number of shells with the increase in atomic number.

(ii) The atomic radii of second (4d) and third (5d) transition series in a group is almost same except Y(39) and La(57)

(D) For inner transition elements:

As we move along the lanthanide series, there is a decrease in atomic as well as ionic radius. The decrease in size is regular in ions but not so regular in atoms. This is called lanthanide contraction*.

Exceptions :

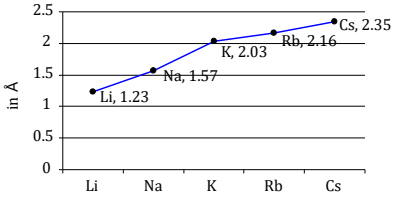
(1) Noble gases have largest atomic sizes [Vander waal's radii]. However, their covalent radii are smaller e.g. Xe.

(2) Size of Al > Ga, [Z_{eff} increasing] (d-contraction or scandide contraction)

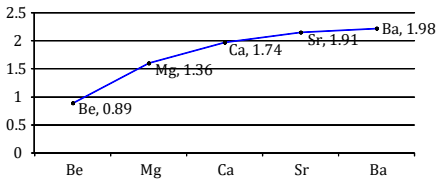
(3) Size of Hf & Zr are same (lanthanide contraction)

Graphical representation of atomic radius:

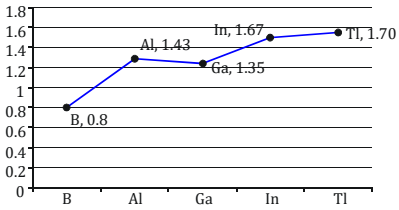
ALKALI METALS



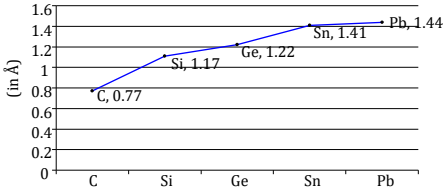
ALKALINE EARTH METALS



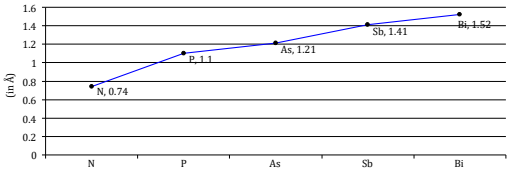
BORON FAMILY



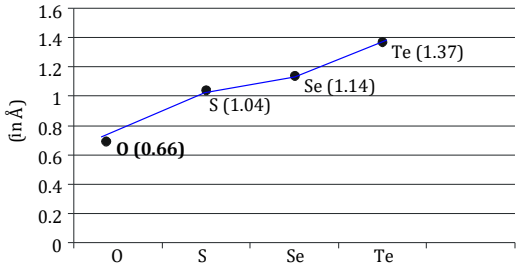
CARBON FAMILY



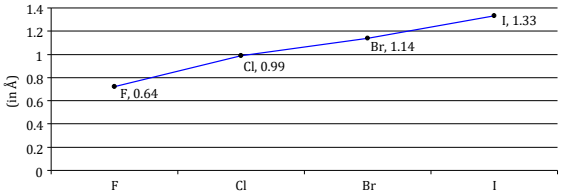
NITROGEN FAMILY (PNICOGENS)



CHALCOGENS



HALOGENS



[what can you predict or say about the increment in size along a group and decrement along a period]

Practice Questions

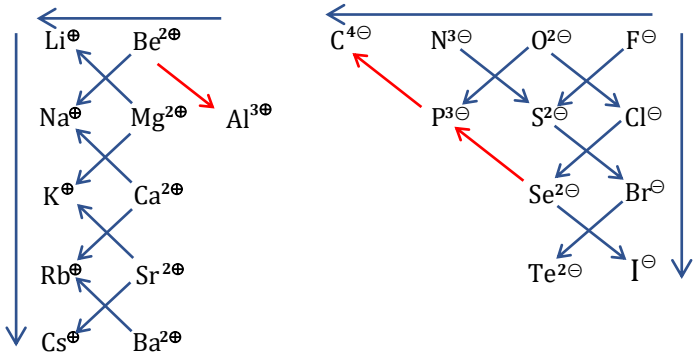
Q. From the given set of species, point out the species from each set having least atomic radius :
(1) $O^{2\ominus}$, $F^{\ominus}$, $Na^{\oplus}$ (2) Ni, Cu, Zn (3) Li, Be, Mg (4) He, $Li^{\oplus}$, $H^{\ominus}$
Correct answer is :-
(1) $O^{2\ominus}$, Cu, Li, $H^{\ominus}$ (2) $Na^{\oplus}$, Ni, Be, $Li^{\oplus}$
(3) $F^{\ominus}$, Zn, Mg, He (4)$Na^{\oplus}$, Cu, Be, He Ans.(2)

Q. Which of the following pairs of elements have almost similar atomic radii :
(1) Zr, Hf (2) Mo, W (3) Co, Ni (4) All Ans.(1)

Q. If the ionic radii of $K^{\oplus}$ and $F^{\ominus}$ are nearly the same (i.e. 1.34 Å) then the atomic radii of K and F respectively are :
(1) 1.34 Å, 1.34 Å (2) 0.72 Å, 1.96 Å
(3) 1.96 Å, 0.72 Å (4) 1.96 Å, 1.34 Å Ans.(2)

Note :

In the direction of arrow (→) ionic size increases.



Ionisation Potential Or Ionization Energy Or Ionization Enthalpy:

(i) Minimum energy required to remove most loosely held outer most shell $e^{\ominus}$ in ground state from an isolated gaseous atom is known as Ionisation Potential.

(Isolated → Without any bonding with other atom)

(ii) Successive I.E.

(a) For an atom M, successive ionisation energies are as follows -

$M_{(g)} + E_1 \longrightarrow M^{\oplus}_{(g)} + e^{\ominus}$ $\quad E_1 = IP_1$

$M^{\oplus}_{(g)} + E_2 \longrightarrow M^{2\oplus}_{(g)} + e^{\ominus}$ $\quad E_2 = IP_2$

$M^{2\oplus}_{(g)} + E_3 \longrightarrow M^{3\oplus}_{(g)} + e^{\ominus}$ $\quad E_3 = IP_3$

$IP_3 > IP_2 > IP_1$

(b) $e^{\ominus}$ can not be removed from solid state of an atom, it has to convert in gaseous form, Energy required for conversion from solid state to gaseous state is called Sublimation energy.

(c) IP is always an endothermic process (ΔH = +ve)

(d) It is measured in eV/atom (electron volt/atom) or K.cal/mole or KJ/mole

(iii) Factors affecting ionisation potential:

(A) Atomic size : Larger the atomic size, smaller is the I.P. It is due to that the size of atom increases the outermost electrons $e^{\ominus}$ farther away from the nucleus and nucleus loses the attraction on that electrons and hence can be easily removed. (Generally, down the group I.P. decreases)

$$IP \propto \frac{1}{\text{Atomic size}}$$

(B) Effective nuclear charge (Z_{eff}) : I.P. increases with the increase in nuclear charge between outermost electrons and nucleus. (Generally, Along the period I.P. increases)

I.P. ∝ Effective nuclear charge

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

$$\text{I.P.} \propto \frac{1}{\text{Screening effect}}$$

(D) Penetration power of sub shells : (Subshells of the same shell)

(a) Order of attraction of subshells towards nucleus. (Penetration power) is - s > p > d > f

(b) As s-subshell is more closer to nucleus so more energy will be required to remove $e^{\ominus}$ in comparison to p, d & f.

(c) Order of decreasing IP is s > p > d > f

Example IP_1 of Be > IP_1 of B
$(1s^2 2s^2)$ $(1s^2 2s^2 2p^1)$

After loosing one $e^{\ominus}$, B attains electronic configuration of Be, so 2nd IP of B is more than Be.

IP_2 of B > Be

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

(a) Half filled p^3,d^5, f^7 or fully filled s^2, p^6, d^{10}, f^{14} are more stable than others so it requires more energy.
e.g. In C, N, O, F
IP_1 order is C < O < N < F

(b) Because of half filled p-orbitals in N, its ionisation energy (stability) is higher than O.
– IP_1 order Na < Al < Mg

(c) Because s-orbital in Mg is completely filled and its penetration power is also higher than p-orbital (Al).
IP_2 order $Mg^{\oplus}$ < $Al^{\oplus}$ < $Na^{\oplus}$
(2,8,1) (2,8,2) (2,8)

(F) Oxidation state :

I.P. $\propto$ oxidation state of an atom. Ion with high oxidation state will have high ionisation potential.
e.g. $Fe^{3\oplus}$ > $Fe^{2\oplus}$ > Fe

(iv) IP & Periodic table :

In a group : Size increase so IE decrease.
Li > Na > K > Rb > Cs
Sc > Y > La

Size increases
IE decreases

Exception :

– IP of Al < I.P. of Ga (While I.P. decreases down the group it is due to Transition contraction)
– IP of Zr < IP of Hf (While I.P. should decrease down the group, It is due to lanthanide contraction)
4d 5d
– 3d > 4d < 5d for group 7, 8, 9, 11 and 12
3d < 4d < 5d for group 4, 5, 6 and 10

In a period :- In a period atomic size decreases and Z_{eff} increases so removal of electron becomes difficult and IE increases.

Li Be B C N O F Ne

Z_{eff} increases,
atomic size decreases,
IE increases.

IE = Ne > F > N > O > C > Be > B > Li

I.E. Order in s-block :

IE = H > Be > Mg > Ca > Sr > Li > Ba > Na > K > Rb > Cs

Practice Questions

Q. In which of the following have higher difference in the value of IInd and IIIrd I.P. –
(1) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 2p^1$ (2) $1s^2\ 2s^2\ 2p^6\ 3s^2$
(3) $1s^2\ 2s^2\ 2p^6\ 3s^1$ (4) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$ Ans.(2)

Q. What is the correct order of I.P. –
(1) 5d > 4d (2) 3d < 4d
(3) 4d ~ 5d (4) All Ans.(1)

Q. Amongst the following elements (whose electronic configurations are given below) the one having highest ionisation energy is :-

(1) $[Ne]3s^2 3p^1$ (2) $[Ne]3s^2 3p^3$ (3) $[Ne]3s^2 3p^2$ (4) $3d^{10}, 4s^2 4p^3$ Ans.(2)

Q. In C, N, O, F which of the following order is correct for I.P. –

(1) F > O > C > N (2) O > F > N > C (3) F > N > O > C (4) N > F > O > C Ans.(3)

Q. Which of the following elements has the lowest ionisation potential ?

(1) Na (2) K (3) Mg (4) Al Ans.(2)

(v) Application of ionisation potential:

(A) Metallic and non metallic character:

Metallic $\longrightarrow$ I.P. Low (Na, K, Rb etc.)

non metallic $\longrightarrow$ I.P. High (F, Cl, Br etc.)

$$\text{I.P.} \propto \frac{1}{\text{Metallic Property}}$$

(B) Reactivity :-

$$\text{Reactivity} \propto \frac{1}{\text{Ionisation Potential}}$$

Reactivity increases down the group as I.P. decreases.

(C) Reducing character :

$$\text{Reducing character} \propto \frac{1}{\text{Ionisation character}}$$

(a) IA group has minimum I.P. so they are strong reducing agents in gaseous state (Li < Na < K < Rb < Cs)

(b) VIIA group has maximum I.P. so they are strong oxidising agents (F > Cl > Br > I)

(D) Stability of oxidation states :-

(a) If the difference between two successive I.P. > 16eV then lower oxidation state is stable.

e.g. $Na_{(g)} \longrightarrow Na^{\oplus}_{(g)}$ $IP_1 = 5.13$ eV

$Na^{\oplus}_{(g)} \longrightarrow Na^{2\oplus}_{(g)}$ $IP_2 = 47.28$ eV

$IP_2 - IP_1 = 42.7$ eV

So $Na^{\oplus}$ is stable.

(b) If the difference between two successive I.P. < 11 higher oxidation state is stable.

e.g. $Mg_{(g)} \longrightarrow Mg^{\oplus}_{(g)}$ $IP_1\}$ 7.6 eV

$Mg^{\oplus}_{(g)} \longrightarrow Mg^{2\oplus}_{(g)}$ IP_2 15.1 eV

$IP_2 - IP_1 = 7.5$ eV

So $Mg^{2\oplus}$ is stable.

(E) Basic nature – It is property of elements with loosely held electrons

$$\text{I.P.} \propto \frac{1}{\text{Basic Property}}$$

$M_2O > MO$, $Na_2O > MgO$, $NaOH > Mg(OH)_2$, $Cs_2O > Rb_2O > K_2O > Na_2O > Li_2O$

(F) To determine the number of valence electron :

eg. Li $\longrightarrow$ The value of $IP_1 = 5.4$ eV }

$IP_2 = 75.4$ eV

High energy difference in two successive I.P. shows change in energy levels i.e., valence shell has only one $e^{\ominus}$

(G) Ionic nature : I.P. $\propto \frac{1}{\text{Ionic nature of compound}}$

– Ionic nature of compound decreases from left to right in a period.

(H) Nature of hydroxide :-

In a hydroxide AOH – Nature of hydroxide depends on ionisation potential of A.

A — O — H . If I.P. of A is lower than oxygen, it will loose electron and will be basic

$A - OH \longrightarrow A^{\oplus} + OH^{\ominus}$

If A has high I.P. than oxygen –

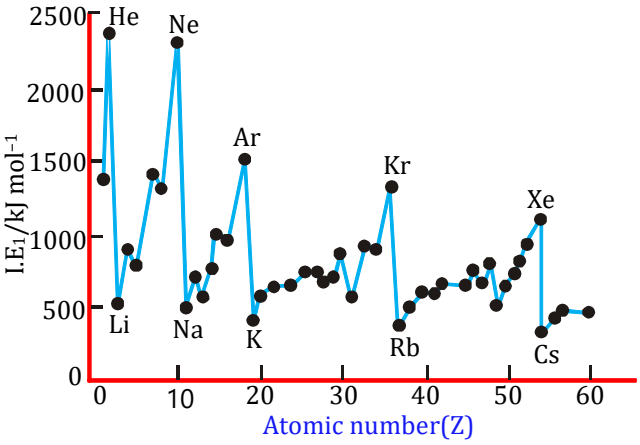
$A - O - H \longrightarrow AO^{\ominus} + H^{\oplus}$

then the hydroxide will be acidic in nature.

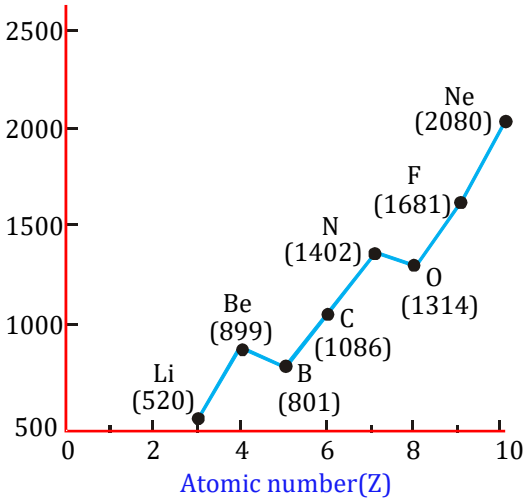
$NaOH, Mg(OH)_2, Al(OH)_3, Si(OH)_4, P(OH)_3, SO_2(OH)_2, ClOH$

$\longrightarrow$

- Along the period
  - I.P. increases
  - Acidic nature increases
  - Basic nature decreases
- In a group
  - I.P. decrease
  - basic nature increases
  - Acidic nature decreases



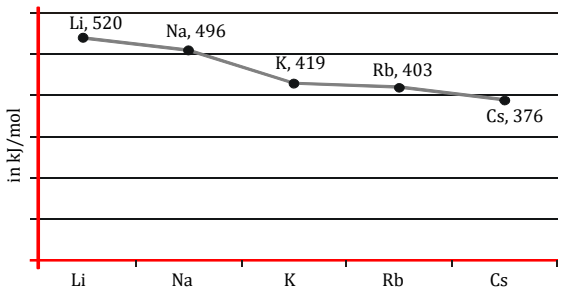
Variation of first ionization energy $(I.E)_1$ with atomic number for elements with Z = 1 to 60



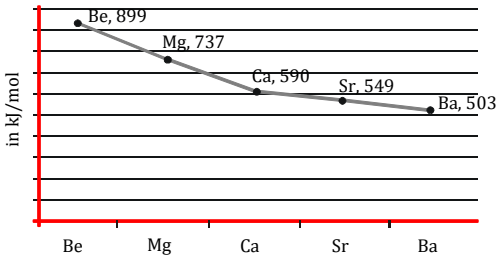
First ionization energy $(I.E)_1$ of elements of the second period as a function of atomic number(Z)

Graphical representation of ionisation energy:

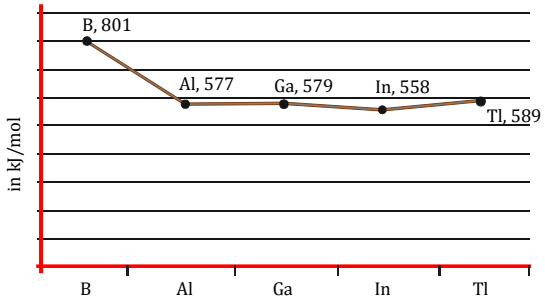
ALKALI METALS



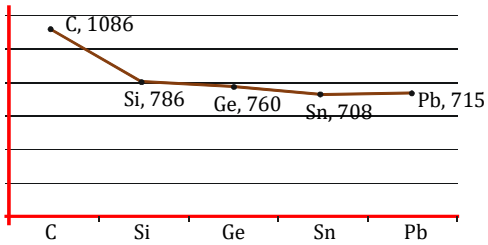
ALKALINE EARTH METALS



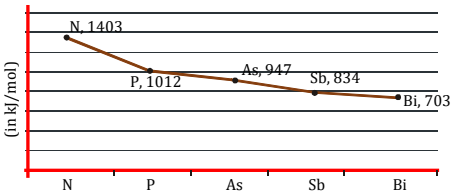
BORON FAMILY



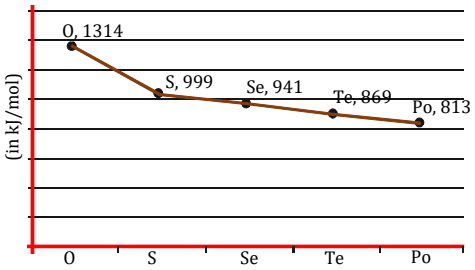
CARBON FAMILY



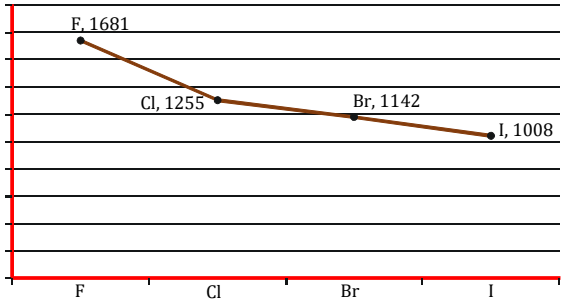
NITROGEN FAMILY (PNICOGENS)



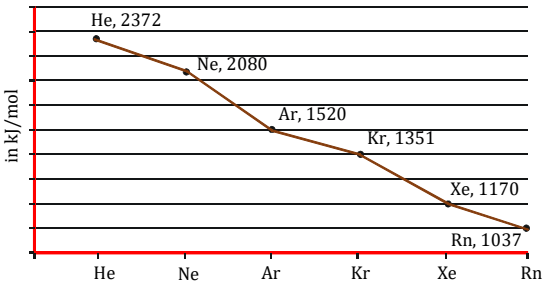
CHALCOGENS



HALOGENS



NOBLE GASES



Ionisation energy of d-block elements:

1st, 2nd, 3rd IE's are increasing from left to right for 1st Transition series, but not regularly.
For 2nd IE Cr > Fe > Mn and Cu > Zn
For 3rd IE Mn > Cr > Fe and Zn has highest.

Element	Sc	Ti	v	Cr	Mn	Fe	Co	Ni	Cu	Zn
Atomic number Electronic configuration	21	22	23	24	25	26	27	28	29	30
M	$3d^1 4s^2$	$3d^2 4s^2$	$3d^3 4s^2$	$3d^5 4s^1$	$3d^5 4s^2$	$3d^6 4s^2$	$3d^7 4s^2$	$3d^{10} 4s^1$	$3d^{10} 4s^1$	$3d^{10} 4s^2$
M^{2+}	$3d^1$	$3d^2$	$3d^3$	$3d^4$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$3d^9$	$3d^{10}$
M^{3+}	[Ar]	$3d^1$	$3d^2$	$3d^3$	$3d^4$	$3d^5$	$3d^6$	$3d^7$	-	-
Enthalpy of atomisation Δ_0H / kJ mol^{-1}										
	326	473	515	397	281	416	425	430	339	126
Ionisation Enthalpy, Δ_0H / kJ mol^{-1}										
Δ_0H I	631	656	650	653	717	762	758	736	745	906
II	1235	1309	1414	1592	1509	161	1644	1752	1958	1734
III	2393	2657	2833	2990	3260	2962	3243	3402	3556	3829

Practice Questions

Q. In which of the following the energy change corresponds to first ionisation potential only :-

(1) $X_{(g)} \rightarrow X^{\oplus}_{(g)} + e$ (2) $X_{2\,(g)} \rightarrow X^{\oplus}_{(g)} + e$

(3) $X_{(s)} \rightarrow X^{\oplus}_{(g)} + e$ (4) $X_{(aq)} \rightarrow X^{\oplus}_{(aq)} +$ Ans.(1)

Q. The correct order of decreasing second ionization energy of Li, Be, Ne, C, B

(1) Ne>B>Li>C>Be (2) Li>Ne>C>B>Be

(3) Ne>C>B>Be>Li (4) Li>Ne>B>C>Be Ans.(4)

Q. Which of the following statement defines the correct representation of second ionisation enthalpy –

(1) The energy required to remove the electron of an unipositive atom

(2) The energy required to remove the any second electron of an atom

(3) The energy required to remove the second most loosely bound electron

(4) The energy required to remove two electron of an atom. Ans.(1)

Q. Match the column.

Column-I	Column-II
Valence electronic configuration	Successive ionisation energies
(a) ns^1	(p) 19, 27, 36, 48, 270
(b) ns^2	(q) 16, 28, 34, 260
(c) $ns^2\ np^1$	(r) 18, 26, 230, 250
(d) $ns^2\ np^2$	(s) 14, 200, 220, 240

Ans. (a-s ; b-r ; c-q ; d-p)

Electron Affinity/Electron gain enthalpy:

(i) The amount of energy released or absorbed when electron is added to the valence shell of an isolated gaseous atom.

$X_{(g)} + e^{\ominus} \longrightarrow X^{\ominus}_{(g)} + E.A.$

known as Electron affinity.

(ii) Mostly energy is released in the process of first EA.

$X_{(g)} + e^{\ominus} \longrightarrow X^{\ominus}_{(g)} + EA_1$

(iii) If EA_1 is exothermic then ($\Delta_{eg}H$ (electron gain enthalpy) = – ve)

(iv) EA of neutral atom is equal to I.P. of its anion.

{Energy evolved in E.A. to convert neutral atom to anion} = {Energy absorbed to remove e^- from anion)

i.e. $X_{(g)} + e^{\ominus} \underset{IP}{\overset{EA}{\rightleftharpoons}} X^{\ominus}_{(g)} + \text{energy}$

On adding 2nd electron in anion $X^{\ominus}$

$X^{\ominus}_{(g)} + e^{\ominus} \longrightarrow X^{2\ominus}_{(g)} - EA_2$

or $X^{\ominus}_{(g)} + e^{\ominus} + EA_2 \longrightarrow X^{2\ominus}_{(g)}$ $\Delta_{eg}H_2$ (2^{nd} electron gain enthalpy) = + ve

(v) EA_2 is endothermic process, i.e. energy has to be given to introduce 2^{nd} electron in an anion, because of repulsion between negative charge on anion ($X^{\ominus}$) and 2^{nd} electron.

(vi) Factors affecting electron affinity :

(a) Atomic size –

$$EA \propto \frac{1}{\text{Atomic size}}$$

(b) Screening effect –

$$EA \propto \frac{1}{\text{Screening effect}}$$

(c) Effective nuclear charge (Z_{eff}) –

$EA \propto Z_{eff}$

(d) Stability of completely filled or half-filled orbitals –

Electron affinity of elements having fully filled or half-filled configuration is very less or nearly zero and in some elements energy is given to add any electron in their valence shell. It is because of stability of fully filled or half-filled configuration.

(vii) In period

(a) Electron affinity increases along the period due to increase in Zeff and decrease in atomic size.

(b) Halogens possess maximum electron affinity due to small size and maximum Zeff and after gaining one electron. They attain a stable inert gas configuration.

(c) Electron affinity of IIA and noble gases is zero.

(d) In IInd period value of electron affinity of C is higher than N, because of half-filled orbitals present in valence shell of N, which is more stable.

(e) In IIIrd period value of electron affinity of Si is higher than P, because of half-filled orbitals present in P.

(f) The value of electron affinity of inert gases is zero due to fully filled p-orbitals.

Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
	Zero			Less			zero

Ne < Be < N < B < Li < C < O < F

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

(viii) In Group –

(a) Electron affinity decreases down the group. As the size increases and attraction on additional electron is less.

(b) Electron affinity of 3rd period element is greater than electron affinity of 2nd period elements of the respective group.

F $2s^2 2p^5$

Cl $3s^2 3p^5$

Due to small size of fluorine, electron density around the nucleus increases. The incoming electron suffers more repulsion. In case of chlorine electron density decreases due to large size, decreasing order of electron affinity of halogen. Cl > F > Br > I

S > Se > O

S > O > P > N

Si > C > P > N

N & P have low electron affinity due to stable half filled configuration.

(ix) Application of Electron affinity :-

(a) Electron affinity ∝ Oxidising nature

But F has more oxidising power than Cl because F has more standard reduction potential.

(b) Electron affinity ∝ Reactivity

(c) Electron affinity ∝ electronegativity

(d) Elements of high electron affinity form oxide and hydroxides, which are acidic in nature

(x) Difference between EN and EA

Electronegativity	**Electron Affinity**
– Tendency of an atom in a molecule to attract the bonded electrons	– Energy released when an electron is added to neutral isolated gaseous atom
– Relative value of an atom	– Absolute value of an atom
– It regularly changes in a period or group	– It does not changes regularly
– It has no unit	– It is measured in eV/atom or KJ mol^{-1} or K.cal $mole^{-1}$

Electron affinity (Data Table) (eV/atom):

1							18
H							He
+0.754		13	14	15	16	17	–0.5
Li		B	C	N	O	F	Ne
+0.618		≤ +0.277	+1.263	–0.073	+1.461	+3.399	–1.2
Na		Al	Si	P	S	Cl	Ar
+0.548		+0.441	+1.385	+0.747	+2.077	+3.617	–1.0
K		Ga	Ge	As	Se	Br	Kr
+0.502		+0.03	+1.2	+0.81	+2.021	+3.365	–1.0
Rb		In	Sn	Sb	Te	I	Xe
+0.486		+0.3	+1.2	+1.07	+1.971	+3.059	–0.8

Practice Questions

Q. Which of the following element has the lowest value of electron affinity -

(1) Carbon (2) Oxygen (3) Fluorine (4) Neon Ans.(4)

Q. Which of the following element has the highest value of electron affinity -

(1) Fluorine (2) Chlorine (3) Sulphur (4) Oxygen Ans.(2)

Electronegativity (EN):

(i) The tendency of an atom to attract shared pair of electrons towards itself is called electronegativity.

(ii) EN and EA both have tendency to attract electrons but electron affinity is for isolated atoms. Where as electronegativity is for bonded atoms.

(iii) A polar covalent or ionic bond of A – B may be broken

as (a) $A - B \longrightarrow A^{\ominus} : + B^{\oplus}$ (EN A > EN B)

or (b) $A - B \longrightarrow A^{\oplus} + :B^{\ominus}$ (EN A < EN B)

depending on their tendency to attract bonded electron.

(iv) There is no unit of electronegativity as EN is tendency to attract electrons of a bonded atom not an energy

(v) Pauling explained it first time.

(vi) In Pauling's scale, elements having almost same electronegativity are –

N = Cl = 3.0

C = S = I = 2.5

P = H = 2.1

Cs = Fr = 0.7

Be = Al = 1.5

(vii) EN of some other elements are as follows –

Electronegativity of Some Elements (on Pauling's Scale)

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	**Cl** 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9					

In **Pauling's Scale** elements having almost same electronegativity are-

C $\simeq$ S $\simeq$ I $\simeq$ 2.5
N = Cl = 3.0
P = H = 2.1
Cs = Fr = 0.7
Be = Al = 1.5

Note :

- Small atoms are normally having more EN than larger atoms
- Electronegativity of F > Cl but Electron affinity of Cl > F

(viii) Factors Affecting electronegativity:

(A) Atomic size :

$$\text{Electronegativity} \propto \frac{1}{\text{Atomic size}}$$

(B) Effective nuclear charge (Z_{eff}) :–

Electronegativity $\propto$ Zeff

(C) Oxidation state :

Electronegativity $\propto$ oxidation state

$Mn^{2\oplus} < Mn^{4\oplus} < Mn^{7\oplus}$

$O^{\ominus} < O < O^{\oplus}$

$Fe < Fe^{2\oplus} < Fe^{3\oplus}$

- As atomic radius decreases by increasing oxidation state of cation species, EN increases.
- In anionic species, the order of electronegativity is $O^{2\ominus} < O^{\ominus} < O$

(D) Electronegativity does not depends on filled or half filled orbitals, because it is a tendency to attract bonded electron, not to gain electron from out side.

(ix) Periodic table & Electronegativity:

(a) Electronegativity decreases down the group.

(b) In period on moving from left to right EN increases.

Exceptions:

– '0' group – Electronegativity of '0' group is always zero, because inert gas do not form molecule.

(c) Electronegativity of Cs and Fr in 6th and 7th period (IA) are equal, it is because from $_{55}Cs$ to $_{87}Fr$ only one shell increases but nuclear charge (No. of proton) increases by +32.

(d) So effect of nuclear charge balanced the effect of increase in number of shell.
Electronegativity of F > Cl but Electron affinity of Cl > F
So Fluorine is called Black sheep element.

(e) In group of IIB elements (Zn, Cd, Hg) value of electronegativity increases down the group, because of lanthanide contraction

(f) In IIIA group, value of electronegativity increases down the group, because of transition contraction (+18 charge)

EN of Ga > EN of Al

(x) Application of electronegativity :

(A) Metallic and non metallic nature :-

Low electronegativity $\longrightarrow$ Metals

High electronegativity $\longrightarrow$ Non metals

Metallic character increases down the group but decreases along a period.

(B) Bond length :-

$$\Delta EN \propto \frac{1}{\text{Bond length}}$$

Here Δ EN = difference in electronegativities of bonded atoms

HF < HCl < HBr < HI

– HF has minimum bond length because of much difference in the electronegativities of H and F.

(C) Bond energy :- By increasing Δ EN bond length decreases and hence bond energy increases

Bond energy $\propto$ Electronegativity difference

HF > HCl > HBr > HI

(D) Acidic strength of hydrides :

Bond energy (Strength) $\propto$ stability of molecule.

– Order of stability of hydrohalides is

HF > HCl > HBr > HI and so order of their acidic strength will be –

HF < HCl < HBr < HI

– In VA group –

NH_3 ↓ * Thermal stability decreases

PH_3 * Basic character decreases

AsH_3 * Acidic character increases

In PH_3 and AsH_3 there is less difference in the electronegativities of X_A and X_B, so their bond energy decreases and hence acidic character (loosing H^+ ion) increases.

(E) Reactivity :-

$$\text{Bond energy} \propto \text{Stability} \propto \frac{1}{\text{Reactivity}}$$

As bond energy difference of electronegativities

so, Δ Electronegativity $\propto$ Stability $\propto \frac{1}{\text{Reactivity}}$

HF < HCl < HBr < HI

$\longrightarrow$

Reactivity

– HI is most reactive hydrohalides or strongest acid among all hydrohalides.

(F) Nature of bonds –

It can be determined by Hanny & Smith formula -

Ionic %= $16 (X_A - X_B) + 3.5 (X_A - X_B)^2$

Here X_A= Electronegativity of A

X_B = Electronegativity of B

If $X_A - X_B > 2.1$ Ionic % > 50% i.e. Ionic bond

IF $X_A - X_B < 2.1$ Ionic % < 50% i.e. covalent bond

– If $X_A = X_B$; then A – B will be non polar,

e.g. H—H, F—F

– If $X_A > X_B$ and difference of EN is small then $A^{\delta\ominus}$ —— $B^{\delta\oplus}$ bond will be polar covalent

e.g. H_2O ($H^{\delta\oplus}$ —— $O^{\delta\ominus}$ —— $H^{\delta\oplus}$)

– If $X_A >> X_B$ and $X_A - X_B$ ΔEN is high then $A^{\ominus}$ —— $B^{\oplus}$ bond will be polar ionic

e.g. $Na^{\oplus}Cl^{\ominus}$

(G) Nature of hydroxides :-

(a) If $X_A - X_O > X_O - X_H$

then AO bond will be more polar and will break up as

(b) If $X_A - X_O < X_O - X_H$ A—OH → $A^{\oplus}$ + $OH^{\ominus}$

It shows basic nature

e.g. In NaOH A—O—H → $H^{\oplus}$ + $AO^{\ominus}$

It shows acidic nature

$X_O - X_{Na}$ (2.6) > $X_O - X_H$ (1.4)

So hydroxide is basic

In ClOH –

$X_O - X_{Cl}$ (0.5) < $X_O - X_H$ (1.4)

So hydroxide is acidic

(H) Nature of oxides – Consider an oxide AO

If $X_A - X_O > 2.3$ Basic oxide

If $X_A - X_O = 2.3$ Amphoteric oxide

If $X_A - X_O < 2.3$ Acidic oxide

(a) Along a period acidic nature increases.

(b) Down the group basic nature increases

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl

$X_A - X_O > 2.3$ Basic

$X_A - X_O = 2.3$ Amphoteric

$X_A - X_O < 2.3$ Acidic

i.e. when in periodic table the distance between the element and oxygen increases, basic character increases.

$NO_2 > ZnO > K_2O$

⟶

acidic character decreases

– ZnO, Al_2O_3, Ga_2O_3, BeO, PbO, PbO_2, SnO, SnO_2, Sb_2O_3, As_2O_3, V_2O_5, MnO_2, PtO_2 etc. are amphoteric oxides.

– CO, H_2O, NO, N_2O are neutral oxides.
– Acidic strength of oxide and oxyacid

(a) Along the period from left to right

(i) When central atom is diff.

EN of C.A. ↑, Acidic nature of oxide and oxyacid ↑

B_2O_3 < CO_2 < N_2O_5

$N_2O_5 > P_2O_5 > As_2O_5$

$N_2O_3 > P_2O_3 > As_2O_3 > Sb_2O_3$

$H_3BO_3 < H_2CO_3 < HNO_3$

$H_3PO_4 > H_3AsO_4 > H_3SbO_4$

$H_2SO_3 > H_2SeO_3 > H_2TeO_3$

HOF > HOCl > HOBr > HOI

$HClO_4 > HBrO_4 > HIO_4$

(ii) When central atom is same

Oxidation state of C.A. ↑, EN of C.A. ↑, Acidic nature of oxide and oxyacid ↑

$SO_3 > SO_2$

$N_2O_5 > N_2O_3$

$Sb_2O_3 < Sb_2O_5 < N_2O < NO < NO_2 < N_2O_5$

$HClO_4 > HClO_3 > HClO_2 > HClO$

$HNO_3 > HNO_2$

$H_2SO_4 > H_2SO_3$

(I) Nomenclature of inorganic compounds :-

Prefix —— less electronegative element

Suffix —— More electronegative element

e.g. Cl_2O(Right) OCl_2(Wrong)

In Dichloroxide the EN of Cl is less than 'O' That's why chloro for Cl is in prefix position.

OF_2 Oxygen difluoride

ICl Iodine chloride

NH_3 Exception (Here H is less electronegative but it is suffix)

(xii) Electronegativity Scale : -

(A) Pauling Scale (Bond energy scale) :-

$0.208\sqrt{\Delta_{AB}} = X_A - X_B$, When B.E. in Kcal/mol

Here Δ_{AB} is resonance energy of AB molecule ; $\Delta_{AB} = BE_{A\text{-}B} - \sqrt{BE_{A\text{-}A} \times BE_{B\text{-}B}}$

BE → Bond Energy.

This equation gives difference in electronegativity values.

(B) Mulliken scale :- According to Mulliken electronegativity is average value of IP and EA of an element,

$$X_m = \frac{IP + EA}{2}$$

Relation between Pauling scale & Mulliken scale

$X_p = 0.336\ (X_m - 0.615)$

X_p = Electronegativity given by pauling $X_p = \frac{X_m}{2.8}$

X_m = Electronegativity given by mulliken

– If IP and EA are given in eV, then electronegativity by Mulliken on Pauling scale will be

$$\text{EN of an atom} = \frac{IP + EA}{5.6}$$

– If IP and EA are given in kcal/mole then

$$\text{EN of an atom} = \frac{IP + EA}{2 \times 62.5}$$

(C) Allred Rochow's scale :- As per Rochow, electronegativity is the force by which nucleus of an atom attracts electron which are on the covalent radius.

(X_{AR}) Electronegativity = $\frac{0.359\ Z_{eff}}{r^2_{covalent}}$

Relation between X_P & X_{AR} or $X_P = X_{AR} + 0.744$

PRACTICE QUESTIONSS

Q. The electronegativities of F and H are 4.0 and 2.1 respectively. The percent ionic character in H and F bond is

(1) 43 (2) 34 (3) 94 (4) 39 Ans.(1)

Q. Which will be the correct order of electronegativity –

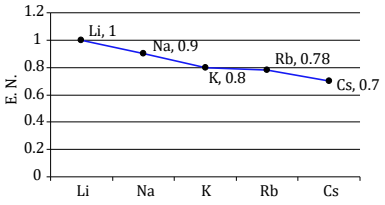
(1) I < Br < F < Cl (2) F > Cl > Br > I (3) I > Br > F > Cl (4) F < Cl < Br < I Ans.(2)

Properties Dependent On Electro Negativity:

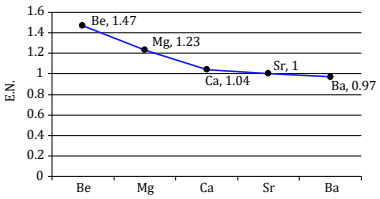
(1) % ionic character (2) Strength of bond
(3) Bond Length (4) Nature of hydrides
(5) Nature of hydroxide.

Graphical representation of Electronegativity:

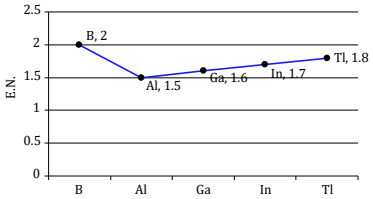
ALKALI METALS



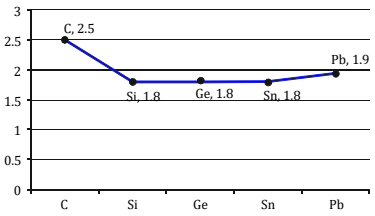
ALKALINE EARTH METALS



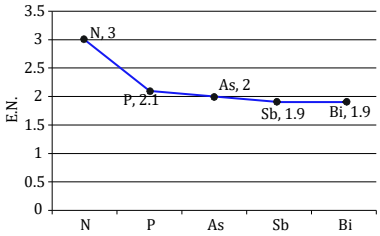
BORON FAMILY



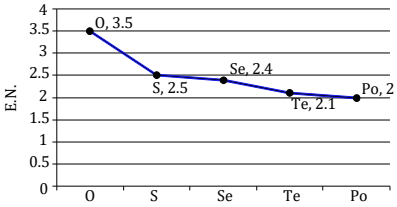
CARBON FAMILY



NITROGEN FAMILY (PNICOGENS)



CHALCOGENS



HALOGENS