PERIODIC TABLE AND ITS PROPERTIES

INTRODUCTION

Periodic table helps us to undergo a systematic study of the various elements found in nature, Without which it would have been impossible for us to study all the elements. By classifying the elements into various groups and periods a comparative study of the elements and their compounds can be done. It also help us to analyze the periodic trend in various properties such as ionization potential, electron affinity, electronegativity etc.

DEVELOPMENT OF PERIODIC TABLE

1. PROUT'S HYPOTHESIS

He simply assumed that all the elements are made up of hydrogen, so we can say that

Atomic weight of element = n × (Atomic weight of one hydrogen atom)

$\boxed{\text{Atomic weight of H} = 1}$ where n = number of hydrogen atom = 1, 2, 3,....

Drawback or Limitation

(a) Every element can not be formed by Hydrogen.

(b) The atomic weights of all elements were not found as whole numbers.

Ex. Chlorine (atomic weight 35.5) and Strontium (atomic weight 87.5)

2. DOBEREINER TRIAD RULE

J.W. Dorbereiner pointed out that within a group of three elements having similar chemical and physical properties, the atomic weight of the middle element is the mean of the other two. Some examples of such triads are given below. He also pointed out the triad - iron, cobalt and nickel in which the atomic weights of the elements are almost the same.

Some representative triads of Dobereiner

Triad Elements	Li	Na	K	Ca	Sr	Ba	S	Se	Te	Cl	Br	I
Atomic weight	7	23	39	40	88	137	32	80	128	35.5	80	127
Mean value		23			88.5			80			81.25	

Other Ex. (K, Rb, Cs), (P, As, Sb) (H, F, Cl) (Sc, Y, La).

Though it was the first successful attempt to rationalise the problem, it could not be generalised or extended.

Drawback or Limitation : All the known elements could not be arranged as triads.

3. NEWLAND'S OCTET LAW

John Alexander Reina newland in England made the first attempt to correlate the chemical properties of the elements with their atomic weight. According to him -

(a) If the elements are arranged in order to their increasing atomic weights, every eighth element had similar properties to first one like the first and eighth note in music. For example

Sa	Re	Ga	Ma	Pa	Dha	Ni	Sa
Li	Be	B	C	N	O	F	Na
Na	Mg	Al	Si	P	S	Cl	K

(b) Inert gases were not discovered till then.

(c) All the elements could not be classified on this basis.

4. LOTHER MEYER'S CURVE

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

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

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

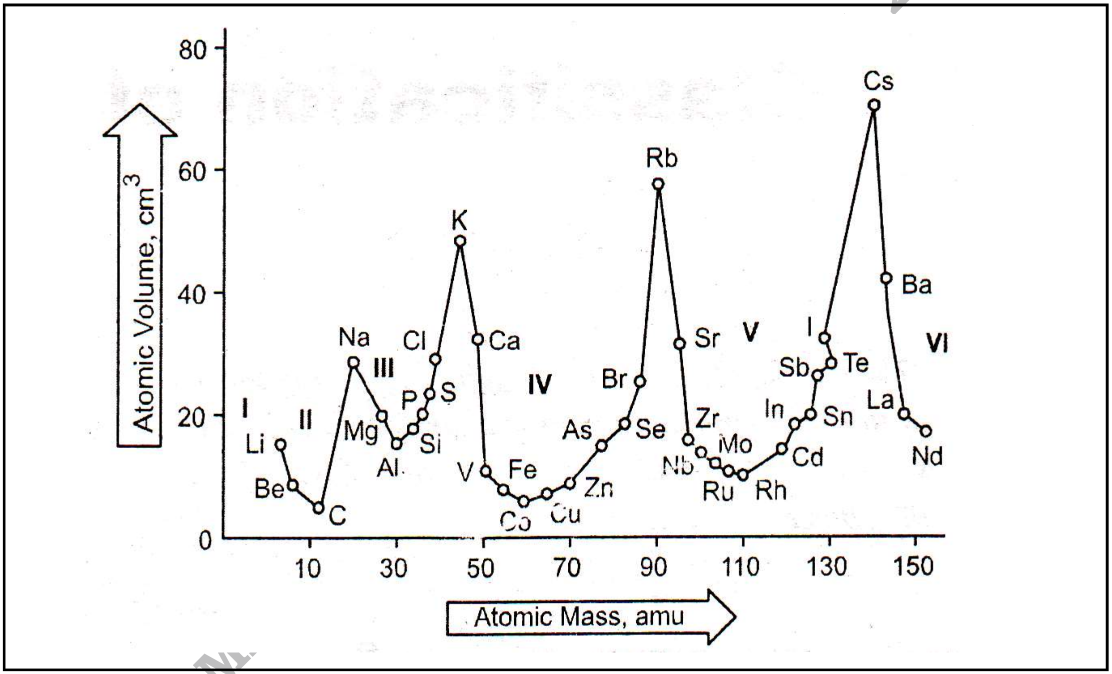
(II) Less electropositive i.e. alkali earth metal (Be, Mg, Ca, Sr, Ba) occupy the descending position on the curve.

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

(IV) 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 Lother Meyer proposed that the physical properties of the elements are periodic function of their atomic wt. and this become the base of Mendeleef's periodic table.



5. TELLURIC HELIX : A three dimensional periodic table given by De-chan-chortois.

6. MENDELEEF'S PERIODIC TABLE

(a) Mendeleef's periodic law : The physical and chemical properties of elements are the periodic function of their atomic weight

(b) Characteristic of Mendeleef's periodic table

(I) It is based on atomic weight

(II) 63 elements were known, noble gases were not discovered.

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

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

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

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

(VII) The VIIIth group was consists of 9 elements in three rows (Transitional metals group).

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

(c) Merits or advantages of Mendeleef's periodic table

(I) **Study of elements :** First time all known elements were classified in groups according to their similar properties. So study of the properties of elements become easier .

(II) **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)

These were the elements for whom position and properties were well defined by Mendeleev even before their discoveries and he left the blank spaces for them in his table.

Ex. 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 mendeleef periodic table were.

Eka aluminium – Galium(Ga)
Eka Silicon – Germanium (Ge)
Eka Boron – Scandium (Sc)
Eka Mangense – Technetium (Tc)

(III) **Correction of doubtful atomic weights :** Correction were done in atomic weight of some elements.

$$\boxed{\text{Atomic weight} = \text{Valency} \times \text{Equivalent weight.}}$$

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.**

(d) Defects of Mendeleef's Periodic Table

(I) Position of hydrogen is uncertain. It has been placed in lA and VII A groups because of its resemblance with both the groups.

(II) No separate positions were given to isotopes.

(III) It is not clear whether the lanthanides and actinides are related to IIA or IIB group.

(IV) Although there is no resemblance except valency of subgroups A and B, they have been put in the same group.

(V) Order of increasing atomic weights is not strictly followed in the arrangement of elements in the periodic table. For e.g. – Co (At. wt. 58.9) is placed before I (127) and Ar (39.9) before K (39).

7. MODERN PERIODIC TABLE (MODIFIED MENDELEEF PERIODIC TABLE)

(a) It was proposed by Moseley.

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

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

He found out that $\sqrt{\nu} \propto Z$ where ν = frequency of X-rays, Z = atomic number.

(d) **Modern periodic law :** The physical & chemical properties of elements are the periodic function of their atomic number.

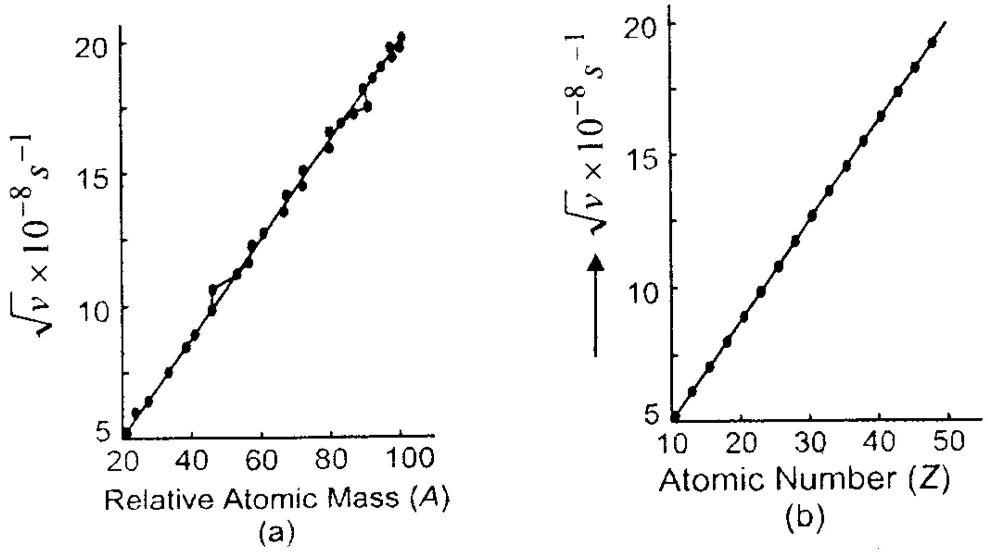
CHARACTERISTICS OF MODERN PERIODIC TABLE

(a) 9 vertical columns called groups.

(b) Ist to VIII group + 0 group of inert gases.

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

(d) 7 horizontal series called periods.



Moseley's Experiment

8. LONG FORM / PRESENT FORM OF MODERN PERIODIC TABLE

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

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

(b) This model is proposed by Rang & Werner

(c) 7 periods and 18 vertical columns (groups)

(d) According to I. U. P. A. C. 18 vertical columns are named as Ist to 18th group.

(e) Elements belonging to same group having same number of electrons in the outermost shell so their properties are similar.

(f) Elements belonging to same group having 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,	26	$_{87}Fr$ – $_{112}Uub$	Incomplete

KEY POINTS

(I) 2nd period elements (Li, Be, B) Shows diagonal relationship with 3rd period elements (Mg, Al, Si) so (Li, Be, B) are called Bridge elements. Because of same ionic potential value they shows similarity in properties.

(Ionic potential = Charge/Radius)

$$\begin{tikzcd} Li \arrow[dr] & Be \arrow[dr] & B \arrow[dr] & \\ Na & Mg & Al & Si \end{tikzcd}$$

(II) 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.

(III) Atomic number of last inert gas element is 86.

(IV) 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.

Number of Solid elements – 95 (if discovered elements are 112)

(V) 2nd period contains maximum number of gaseous elements. They are 4 **(N, O, F, Ne)**

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	IUPAC Official Name	IUPAC 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	Hassnium	Hs
109	Un nil ennium	Une	Meitnerium	Mt

110	Un un nilium	Uun	Darmstadtium	Ds

CLASSIFICATION OF ELEMENTS

s-BLOCK ELEMENTS

(a) The elements of the periodic table in which the last electron enters in s-orbital, are called s-block elements.

(b) s-orbital can accommodate a maximum of two electrons.

(c) Their general formulae are ns^1 and ns^2 respectively, where n = (1 to 7)

(d) IA group elements are known as alkali metals because they react with water to form alkali. II A group elements are known as alkaline earth metals because their oxides react with water to form alkali and these are found in the soil or earth.

(e) Total number of s-block elements are 14.

(f) Fr^{57} and Ra^{88} are radioactive elements while H and He are gaseous elements.

(g) Cs and Fr are liquid elements belonging to s-block.

p-BLOCK ELEMENTS

(a) The elements of the periodic table in which the last electron gets filled up in the p-orbital, called p-block elements.

(b) p-orbital can accommodate a maximum of six electrons. Therefore, p-block elements are divided into six groups which are III A, IV A, V A, VI A, VII A and zero group.

(c) The general formula of p-block elements is $ns^2 p^{1-6}$ (where n = 2 to 6)

(d) The zero group elements having general formula ns^2p^6 are inert, because their energy levels are fully filled.

(e) The total number of p-block elements in the periodic table is 30 (excluding He).

(f) There are nine gaseous elements (Ne, Ar, Kr, Xe, Rn, F_2, Cl_2, O_2 and N_2) belonging to p-block. Gallium (Ga) and bromine (Br) are liquids.

(g) The step-like thick lines drawn in the periodic table in the p-block divides elements into metals, nonmetals and metalloids.

d-BLOCK ELEMENTS

(a) The elements of the periodic table in which the last electron gets filled up in the d-orbital, called d-block elements.

(b) The d-block elements are placed in the groups named III, IV B, V B, VI B, VII B, VIII, I B and II B.

(c) In d-block elements the electron gets filled up in the d-orbital of the penultimate shell.

(d) d-block elements lie between s & p block elements.

(e) The general formula of these elements is $(n–1)s^2, p^6, d^{1-10}\ ns^{1-2}$ where n = 4 to 7.

(f) All of these elements are metals.

(g) Out of all the d-block elements, mercury is the only liquid element.

f-BLOCK ELEMENTS

(a) The element of the periodic table in which the last electron gets filled up in the f-orbital, called f-block elements.

(b) The f-block elements are from atomic number 58 to 71 and from 90 to 103.

(c) The lanthanides occur in nature in low abundance and therefore, these are called rare earth elements.

(d) There are 28 f-block elements in the periodic table.

(e) The elements from atomic number 58 to 71 are called lanthanides because they come after lanthanum (57). The elements from 90 to 103 are called actinides because they come after actinium (89).

(f) All the actinide elements are radioactive.

(g) All the elements after atomic number 92 (i.e. U^{92}) are transuranic elements.

(h) The general formula of these elements is $(n-2)\ s^2\ p^6\ d^{10}\ f^{(1\text{-}14)}\ (n-1)\ s^2\ p^6\ d^{0\text{-}1}\ ns^2$ where n = 6 & 7.

Ex. Elements A, B, C, D and E have the following electronic configurations :

A : $1s^2\ 2s^2\ 2p^1$ B : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
C : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$ D : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
E : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

Which among these will belong to the same group in the periodic table ?

Sol. Out of these, elements A and B will belong to the same group of the periodic table because they have same outer electronic configuration, $ns^2\ np^1$.

TYPE OF ELEMENTS

Using electronic configuration as the criteria, we generally recognize four general type of elements; the inert gas elements, the representative elements, the transition elements, and the inner transition elements. The classification of the elements into these groups is dependent on the extent to which the s, p, d and f orbitals are filled.

Inert Gases

(a) s and p orbitals of the outer most shell of these elements are completely filled. The outermost electronic configuration is ns^2np^6.

(b) Helium is also inert gas but its electronic configuration is $1s^2$

Representative or Normal Elements

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

(b) Inner shell are complete.

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

Transition Elements

(a) Last two shells of these elements namely outermost and penultimate shells are incomplete.

(b) The last shell contains one or two electrons and the penultimate shell may contain more than eight or up to eighteen electrons.

(c) Their outermost electronic configuration is similar to d-block elements i.e. $(n-1)\ d^{1\text{-}10}\ ns^{1\text{-}2}$.

(d) According to definition of transition elements, those elements which have partly filled d-orbitals in neutral state or in any stable oxidation state are called transition elements. According to this definition Zn, Cd and Hg (IIB group) are d-block elements but not transition elements because these elements have d^{10} configuration in neutral as well as in stable +2 oxidation state.

(e) Because of the extra stability which is associated with empty, half-filled, and fully filled subshells, there are some apparent anomalies in electronic arrangements in the transition series. This empirical rule is illustrated by the chromium and copper configuration in the first d series of elements:

h	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
3d	1	2	3	5	5	6	7	8	10	10
4s	2	2	2	1	2	2	2	2	1	2

Inner Transition Elements

(a) In these elements last three shells i.e. last, penultimate and prepenultimate shells are incomplete.

(b) These are related to IIIB i.e. group 3.

(c) The last shell contains two electrons. Penultimate shell may contain eight or nine electrons and pre-penultimate shell contains more than 18, up to 32 electrons.

(d) Their outermost electronic configuration is similar to f-block element i.e. $(n-2)f^{1-14}(n-1)s^2(n-1)p^6(n-1)d^{0-1}ns^2$

PREDICTING ATOMIC NUMBER OF SUCCESSIVE MEMBER IN A GROUP OR FAMILY

Magic Numbers

(a) Knowing the atomic number of the first member of a group, we can write the atomic number of the subsequent elements by adding given magic number

Ex.

Group	1 IA	2 IIA	3 IIIB	4, 5, 6, 7, 8, 9 10, 11, 12,	13, 14, 15, 16, 17 IIIA IV V VI VII	18 'O'gp.
PERIOD 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						

(b) In group IA – Atomic number of H is 1 and atomic number of other element will be as follows –

Magic number $\underbrace{H_1\ 1+2=3}_{2}\ \underbrace{Li\ 3+8=11}_{8}\ \underbrace{Na\ 11+8=19}_{8}\ \underbrace{K\ 19+18=37}_{18}\ \underbrace{Rb\ 37+18=55}_{18}\ Cs$

KEY POINTS

Determination of period, block and group of an element

(a) **Period number:** The period no. of the element can be predicted from the principal quantum no. (n) of the valence shell.

(b) **Block number:** Last electron enter in which orbital is knows as block no. .

(c) **Group number:** It is predicted from the number of electrons in the valence shell and penultimate shell.

Ex.

S.No.	Electronic Configuration	Period number	Block number	Prediction of Group	Group number
1.	[Ar]$4s^2 3d^{10} 4p^6, 5s^1$	5	s	No. of ns e^-	1
2.	[Kr] $5s^2, 4d^{10} 5p^2$	5	p	ns e^- + 10 + np e^-	2+10+2 =14
3.	[Rn] $7s^2, 6d^4 5f^{14}$	7	d	ns e^-+ (n–1)d e^-	2+4 =6
4.	[Xe] $6s^2, 5d^1, 4f^{12}$	6	f	—	3/III B

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 Periodicity

The periodic repetition of the properties of the elements is due to the recurrence of similar valence shell electronic configurations after certain regular intervals. For example, alkali metals have same valence shell 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 configurations of elements.

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

Periods

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

- Each period consists of a series of elements having same valence shell.
- Each period corresponds to a particular principal quantum number of the valence shell present in it.
- Each period starts with an alkali metal having outermost electronic configuration as ns^1.
- Each period ends with a noble gas with outermost electronic configuration ns^2np^6 except helium having outermost electronic configuration as $1s^2$.
- Each period starts with the filling of new energy level.
- The number of elements in each period is twice the number of atomic orbitals available in energy level that is being filled. For illustration.
- I^{st} period shortest period having only two elements. Filling of electrons takes place in the first energy shell, for which,

  $n = 1, \ell = 0$ (s-subshell) and $m = 0$.

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

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

  $n = 3, \quad \ell = 0, 1, 2$ and

number of orbitals	1	3	5
	(3s)	(3p)	(3d)
Total number 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 level). Hence 3^{rd} 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.

Periods	Number of Elements	Called as
$(1)^{st}$ n = 1	2	Very short period
$(2)^{nd}$ n = 2	8	Short period
$(3)^{rd}$ n = 3	8	Short period
$(4)^{th}$ n = 4	18	Long period
$(5)^{th}$ n = 5	18	Long period
$(6)^{th}$ n = 6	32	Very long period
$(7)^{th}$ n = 7	19	Incomplete period

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 : Valency of H = 1

It is defined as the number of hydrogen 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
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^- and from VA to zero group, it is –

[8– (number of valence e^-)].

Valency = No. of valence e^- (IA to IVA)

Valency = (8– no. of valence e^-) (VA to 0)

	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^6
Valence shell e^-	1	2	3	4	5	6	7	8
Valency	1	2	3	4	3	2	1	0
					$(8 - 5) = 3$			$(8 - 8) = 0$

Note : All the elements of a group have same valencies because they have same number of valence shell electrons.

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

(a) Valence shell e^- suffer force of attraction due to nucleus and force of repulsion due to inner shell electrons.

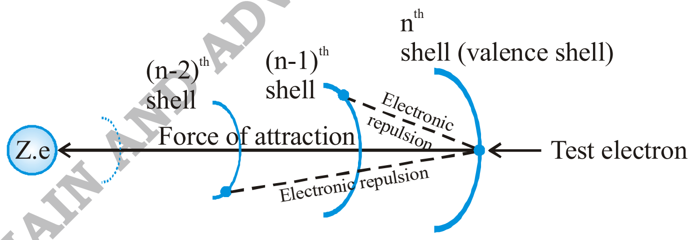
(b) The decrease in force of attraction on valence e^- due to inner shell e^- is called screening effect or shielding effect.(i.e. total repulsive force is called shielding effect.)

(c) Due to screening effect. valence shell e^- experiences less force of attraction exerted by nucleus.(i.e. total attraction force experienced by valence e^- is called Zeff.)

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

(e) If nuclear charge = Z, then effective nuclear charge = Z – σ (Where σ (Sigma)= Screening constant)

So, Zeff = (Z – σ)



Slater's rule to know screening constant (σ)

(a) Screening effect (S.E.) of one e^- of the 1s is 0.30. **Ex.** In He ($1s^2$)

Screening effect of one 1s e^-. where σ = 0.30

∴ Zeff = Z – σ = 2 – 0.30 = 1.7

(b) Screening effect of ns and np (Outermost orbit) electron is 0.35

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

(d) Screening effect of (n – 2) and below all the e^- present in s, p, d, f is 1.0

(Effective Nuclear charge of elements of second period)

Element	Electronic Configuration	Z	σ of ns & np electron	σ (n–1) orbital	Total Screening Constant (a+b)	Effective nuclear charge $Z^*=Z-\sigma$
$_3Li$	$1s^2\ 2s^1$	3	–	0.85×2=1.70	1.70	1.30
$_4Be$	$1s^2, 2s^2$	4	1×0.35=0.35	0.85×2=1.70	2.05	1.95
$_5B$	$1s^2,2s^2,2p^1$	5	2×0.35=0.70	0.85×2=1.70	2.40	2.60
$_6C$	$1s^2,2s^2,2p^1$	6	3×0.35=1.05	0.85×2=1.70	2.75	3.25
$_7N$	$1s^2,2s^2,2p^3$	7	4×0.35=1.40	0.85×2=1.70	3.10	3.90
$_8O$	$1s^2,2s^2,2p^4$	8	5×0.35=1.75	0.85×2=1.70	3.45	4.55
$_9F$	$1s^2,2s^2,2p^5$	9	6×0.35=2.10	0.85×2=1.70	3.80	5.20

Periodic variation

(a) From left to right in a period Zeff increases

(I) That is why in a period Zeff increases by 0.65 and hence atomic size decreases considerably.

(II) In transition series Z increase by + 1 but screening effect increases by 0.85 So Zeff is 0.15

$(1-\ 0.85=0.15)$ [Because e^- enters in $(n-1)$ orbit which has value of $\sigma = 0.85$]

In transition series Zeff increases very less amount, by 0.15 from left to right and hence atomic size remains almost constant.

Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Zeff	3.00	3.15	3.30	3.45	3.60	3.75	3.90	4.05	3.70	4.35

(b) From top to bottom in a group Zeff remain constant

Element	Li	Na	K	Rb	Cs	Fr
Zeff	1.30	2.20	2.20	2.20	2.20	2.20

ATOMIC RADIUS

The average distance of valence shell e^- from nucleus is called atomic radius. It is very difficult to measure the atomic radius because –

(I) The isolation of single atom is very difficult.

(II) There is no well defined boundary for the atom. (The probability of finding the e^- is 0 only at infinity).

So, the more accurate definition of atomic radius is –

— Half the inter-nuclear distance(d) between two atoms in a homoatomic molecule is known as atomic radius.

— This inter-nuclear distance is also known as bond length.Inter-nuclear distance depends upon the type of bond by which two atoms combine.

Based on the chemical bonds, atomic radius is divided into four categories –

1. Covalent radius 2. Ionic radius 3. Metallic radius 4. Vander waal radius

1. Covalent radius

One half of the distance between the nuclei (internuclear distance) of two covalently bonded atoms in homodiatomic molecule is called the covalent radius of that atom. The covalent bond must be single covalent bond. The covalent radius (r_A) of atom A in a molecule A_2 may be given as:

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

i.e. the distance between nuclei of two single covalently bonded atoms in a homodiatomic molecule is equal to the sum of covalent radii of both the atoms

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

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

According to Schomaker and stevenson –

$$D_{A\text{-}B} = r_A + r_B - 0.09\,\Delta_x$$

Where Δ_x is the difference of electronegativities of the atoms A and B.

According to Pauling – If the electronegativities of the two atoms A and B are x_A and x_B respectively then

$$D_{A\text{-}B} = r_A + r_B - (C_1x_A - C_2x_B)$$

C_1 and C_2 are the Stevenson's coefficients for atoms A and B respectively.

2. Metallic Radius

Metal atoms are assumed to be closely packed spheres in the metallic crystal. These metal atom spheres are considered to touch one another in the crystal. One half of the internuclear distance between the two closest metal atoms in the metallic crystal is called metallic radius.

Metallic > Covalent radius

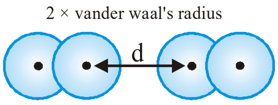
For example – Metallic radius and covalent radius of potassium are 2.3 Å and 2.03 Å respectively.

3. Van Der Wall's Radius or Collision radius

The molecules of non metal atoms are generally gases. On cooling, the gaseous state changes to solid state.

In the solid state, the non metallic elements usually exist as aggregations of molecules are held together by van der wall forces. **One half of the distance between the nuclei of two adjacent atoms belonging to two neighbouring molecules of a compound in the solid state is called van der walls radius.**

It may also be defined as half of the inter nuclear distance of two non bonded neighbouring atoms of two adjacent molecules.



van der Wall's radius = $\frac{1}{2}$ × Internuclear distance between two successive nuclei of two covalent molecules (d)

Van der wall's radius > Metallic radius > Covalent radius

The vander walls radius and covalent radius of chlorine atom are 1.80 Å and 0.99 Å respectively

4. Ionic Radius

A neutral atom changes to a cation by the loss of one or more electrons and to an anion by the gain of one or more electrons. The magnitude of charge on cation and anion is equal to the number of electrons lost or gained respectively. The ionic radii of the ions present in an ionic crystal may be calculated from the inter molecular distance between the two ions.

(a) Radius of Cation

Radius of cation is smaller than that of corresponding atom.

Reasons

(I) During the formation of cation either one shell is removed or

(II) After removing an electron effective nuclear charge increase.

(b) Radius of an Anion

Radius of an anion is invariably bigger than that of the corresponding atom.

Reasons

(I) The effective nuclear charge decrease in the formation of anion. Thus the electrostatic force of attraction between the nucleus and the outer electrons decreases as the size of the anion increases.

(II) Interelectronic repulsion increases.

Factors affecting atomic radius are

(a)	$\text{Atomic radius} \propto \frac{1}{\text{Effective nuclear charge (Zeff)}}$ $Li > Be > B > C > N > O > F$	(b)	Atomic radius $\propto$ number of shells $Li < Na < K < Rb < Cs$
(c)	Atomic radius $\propto$ Screening effect	(d)	Atomic size $\propto$ Magnitude of –ve charge $O < O^{-} < O^{-2}$
(e)	$\text{Atomic radius} \propto \frac{1}{\text{Magnitude of + ve charge}}$ $Mn > Mn^{+2} > Mn^{+3} > Mn^{+4}$	(f)	$\text{Atomic radius} \propto \frac{1}{\text{Bond order}}$ $>N-N< \; > -N=N- > N\equiv N$

Periodic variation of atomic radius

(a) **Across a period :** It decreases from left to right in a period as nuclear charge increases

Ex. $Li > Be > B > C > N > O > F < Ne$

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

Ex. $Li < Na < K < Rb < Cs$

KEY POINTS

Exceptions

(a) Transition elements

$\underbrace{\text{Sc} \quad \text{Ti} \quad \text{V} \quad \text{Cr} \quad \text{Mn}}_{Z_{eff} \; > \; \text{Screening effect}} \qquad \underbrace{\text{Fe} \quad \text{Co} \quad \text{Ni}}_{Z_{eff} \; \approx \; \text{Screening effect}}$

(b) Lanthanide Contraction

(I) Outermost electronic configuration of inner transition elements is

$(n-2)\ f^{1-14},\ (n-1)s^2p^6d^{0-1},\ ns^2\ (n = 6 \text{ or } 7)$

(II) e^- enters in (n – 2) f orbitals

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

(IV) Nuclear charge increases by one (+1) in lanthanides and actinides so atomic size of these elements slightly decreases. It is known as lanthanide contraction. Its effect is also observe in 5d transition series.

Here Nuclear charge > Screening effect.

(V) In Ist, 2nd and 3rd transition series, Radii– $3d < 4d \approx 5d$ **(except IIIrd B)**

	IIIB	IVB	
size	Sc	Ti ↓	size increases
increases	Y	Zr }	Equal due to lanthanide contraction
	La	Hf }	

(c) Transition contraction

IIIA → B<Al ≈ Ga **Note :** While atomic size should increases down the group.

(I) At. size of Ga = At. size of Al, due to transition contraction.

(II) In transition elements nuclear charge increases by 1.

(III) but e^- enters in (n –1)d orbital exerts screening effect.

(IV) Screening effect of (n –1)d e^- balance the nuclear charge by 85%

(V) Z_{eff} on increasing each electron = 1 – 0.85 = 0.15

(VI) Increase in nuclear charge is only 0.15 so atomic size remains almost constant.

Group / Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H • ~0.30																H • ~0.30	He • 1.20*
2	Li ● 1.23	Be • 0.89											B • 0.80	C • 0.77	N • 0.75	O • 0.73	F • 0.72	Ne ● 1.60*
3	Na ● 1.57	Mg • 1.36											Al • 1.25	Si • 1.17	P • 1.10	S • 1.04	Cl • 0.99	Ar ● 1.91*
4	K ● 2.03	Ca ● 1.74	Sc • 1.44	Ti • 1.32	V • 1.22	Cr • 1.17	Mn • 1.17	Fe • 1.17	Co • 1.16	Ni • 1.15	Cu • 1.17	Zn • 1.25	Ga • 1.25	Ge • 1.22	As • 1.21	Se • 1.14	Br • 1.14	Kr ● 2.00*
5	Rb ● 2.16	Sr ● 1.91	Y ● 1.62	Zr • 1.45	Nb • 1.34	Mo • 1.29	Tc • –	Ru • 1.24	Rh • 1.25	Pd • 1.28	Ag • 1.34	Cd • 1.41	In ● 1.50	Sn • 1.40	Sb • 1.41	Te • 1.37	I • 1.33	Xe ● 2.20*
6	Cs ● 2.35	Ba ● 1.98	La ● 1.69	Hf • 1.44	Ta • 1.34	W • 1.30	Re • 1.28	Os • 1.26	Ir • 1.26	Pt • 1.29	Au • 1.34	Hg • 1.44	Tl ● 1.55	Pb • 1.46	Bi ● 1.52	Po	At	Rn
7	Fr	Ra	Ac															

Covalent radius of the elements (In Å)

ISOELECTRONIC SERIES

A series of atoms, ions and molecules in which each species contains same number of electrons but different nuclear charge is called isoelectronic series.

	N^{3-}	O^{2-}	F^{-}	Ne	Na^{+}	Mg^{2+}
Number of e^{-}	10	10	10	10	10	10
Number of p	7	8	9	10	11	12

(a) Number of electrons is same.

(b) Number of protons is increasing.

(c) So the effective nuclear charge is increasing and atomic size is decreasing. In an isoelectronic series atomic size decreases with the increase of charge.

Some of the examples of isoelectronic series are as under.

$S^{2-}, Cl^{-}, K^{+}, Ca^{2+}, Sc^{3+}$ $SO_2, NO_3^{-}, CO_3^{2-}$ N_2, CO, CN^{-} NH_3, H_3O^{+}

Ex. X – X bond length is 1.00 Å and C–C bond length is 1.54 Å. If electronegativities of X and C are 3.0 and 2.0 respectively, then C–X bond length is likely to be ? (using **Stevension & Schomaker** formula).

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

$$= \frac{1.00}{2} + \frac{1.54}{2} - 0.09 \qquad [\Delta\chi = 1] = 1.27 - 0.09\,;$$

Ans. C–X bond length = 1. 18 Å.

Ex. Atomic radius of Li is 1.23 Å and ionic radius of Li^{+} is 0.76Å. Calculate the percentage of volume occupied by single valence electron in Li.

Sol. Volume of Li $= \frac{4}{3} \times 3.14 \times (1.23)^3 = 7.79$ Å $(-\,Li = 1s^2 2s^1)$

Volume of $Li^{+} = \frac{4}{3} \times 3.14 \times (0.76)^3 = 1.84$ Å $(-\,Li^{+} = 1s^2)$

∴ Volume occupied by 2s subshell = 7.79 – 1.84 = 5.95 Å.

∴ % Volume occupied by single valence electron i.e., 2s electron $= \frac{5.95}{7.79} \times 100 = 76.4\%$

Ex. Select from each group the species which has the smallest radius stating appropriate reason.

(a) O, O^{-}, O^{2-} (b) P^{3+}, P^{4+}, P^{5+}

Sol. (a) O is having smallest radius. Anion is larger than its parent atom. Also the anion of the same atom with higher negative charge is bigger in size as compared to anion with smaller negative charge as proton to electron ratio decreases thus attraction between valence shell electrons and nucleus decreases. Hence electron cloud expands.

(b) The ionic radius decreases as more electrons are ionized off that is as the valency increases. So the correct order is $P^{5+} < P^{4+} < P^{3+}$.

Ex. Mg^{2+} is smaller than O^{2-} in size, though both have same electronic configuration. Explain ?

Sol. Mg^{2+} and O^{2-} both are isoelectronic i.e., have same number of electrons. But Mg^{2+} having 12 protons in its nucleus exerts higher effective nuclear charge than O^{2-} having 8 protons and thus valence shell as well as inner shells electrons are more strongly attracted by the nucleus in Mg^{2+} resulting smaller size than O^{2-}.

IONISATION POTENTIAL OR IONISATION ENERGY OR IONISATION ENTHALPY

Minimum energy required to remove most loosely held outer most shell e^- in ground state from an isolated gaseous atom is known as ionisation potential.

(Isolated → Without any bonding with other atom)

1. Successive Ionisation Energy

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

$$M + E_1 \xrightarrow{\qquad} M^+ + e^- \qquad E_1 = I^{st} \text{ Ionisation Potential}$$

$$M^+ + E_2 \xrightarrow{\qquad} M^{+2} + e^- \qquad E_2 = II^{nd} \text{ Ionisation Potential}$$

$$M^{+2} + E_3 \xrightarrow{\qquad} M^{+3} + e^- \qquad E_3 = III^{rd} \text{ Ionisation Potential}$$

I^{st} Ionisation Potential < II^{nd} Ionisation Potential < III^{rd} Ionisation Potential

(b) Electron 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) Ionisation Potential is always an endothermic process ($\Delta H = +ve$)

(d) It is measured in eV/atom (electron volt/atom) or Kcal/mole or KJ/mole

2. Factors affecting ionisation potential

(a) **Atomic size :** Larger the atomic size, smaller is the Ionisation Potential It is due to that the size of atom increases the outermost electrons e^- farther away from the nucleus and nucleus loses the attraction on that electrons and hence can be easily removed.

$$\boxed{\text{Ionisation Potential} \propto \frac{1}{\text{Atomic size}}}$$

(b) **Effective nuclear charge (Zeff) :** Ionisation potential increases with the increase in nuclear charge between outermost electrons and nucleus.

$$\boxed{\text{Ionisation Potential} \propto \text{Effective nuclear charge}}$$

(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 Ionisation Potential

$$\boxed{\text{Ionisation Potential} \propto \frac{1}{\text{Screening effect}}}$$

(d) **Penetration power of sub shells**

(I) Order of attraction of subshells towards nucleus (Penetration power) is -

$s > p > d > f$

(II) As subshell is more closer to nucleus so more energy will be required to remove e^- in comparison to p,d & f. **Ex.** Be B

$1s^2, 2s^2$ $\quad$ $1s^2, 2s^2 2p^1$

Ionisation Potential Be > B

After loosing one e^-, B attains electronic configuration of Be, so II^{nd} ionisation potential of B is more than Be. **II^{nd}onisation Potential of B > Be**

(e) Stability of half filled and fully filled orbitals

(I) 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.

Ex. N O

$1s^2, 2s^2 2p^3$ $1s^2, 2s^2 2p^4$ I^{st} Ionisation Potential order is $O < N$

Because of half filled p-orbitals in N, its ionisation energy (stability) is higher than O.

I^{st} ionisation potential order $Na < Al < Mg$

(II) Because s-orbital in Mg is completely filled and its penetration power is also higher than p-orbital (Al).

II^{nd} ionisation potential order $Mg^+ < Al^+ < Na^+$

(2,8,1) (2,8,2) (2,8)

Periodic variation of ionisation energy

(a) **Variation in period among the representative elements:** Ionisation energies generally increases along the period because in moving left to right in a period the effective nuclear charge per outermost electron increases while the corresponding principal quantum number remain same.

(b) **Variation in a group among the representative elements :** The ionisation energy generally decreases in moving from top to bottom because the size increases due to the increase of the principal quantum number. On the other hand the effective nuclear charge Zeff for the outermost electron remains almost the same along the group.

KEY POINTS

Exception

(I) Ionisation Potential of Al < Ionisation Potential of Ga (While Ionisation Potential decreases down the group it is due to Transition contraction)

577 kj mol^{-1} 579 kj mol^{-1}

(II) Ionisation Potential of Hf > Ionisation Potential of Zr (While Ionisation Potential should decreases down the group. It is due to lanthanide contraction)

760 kj mol^{-1} 5d 674 kj mol^{-1} 4d

(III) In a period atomic size decreases and zeff increases so removal of electron becomes difficult and ionisation potential increases. But N, Be, P, Mg, show high ionisation energy than corresponding of next element.

Li Be B C N O F Ne

$$\xrightarrow{\hspace{10cm}}$$

atomic size decreases, zeff increases, Ionisation Potential increases.

Order of ionisation potential $Li < B < Be < C < O < N < F < Ne$

Application of ionisation potential

(a) Metallic and non metallic character

Metallic $\longrightarrow$ Ionisation Potential Low (Na, K, Rb etc.)

non metallic $\longrightarrow$ Ionisation Potential High (F, Cl, Br etc.)

$$\boxed{\text{Ionisation Potential} \propto \frac{1}{\text{Metallic property}}}$$

(b) Reducing character

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

(I) IA group has minimum ionisation potential so they are strong reducing agents in gaseous state ($Li < Na < K < Rb < Cs$)

(II) IA group – In Aqueous state

reducing character $Li > K \simeq Rb > Cs > Na$

As the degree of hydration is more in Li due to high charge density.

(III) VIIA group has maximum ionisation potential so they are strong oxidising agents ($F > Cl > Br > I$)

(c) Stability of oxidation states

(I) If the difference between two successive ionisation potential $\geq$ 16eV then lower oxidation state is stable.

Ex. $Na \longrightarrow Na^+$ Ist ionisation potential

$Na^+ \longrightarrow Na^{+2}$ IInd ionisation potential $\}$ 42.7 eV

Difference of ionisation potential > 16 eV So Na^+ is more stable.

(II) If the difference between two successive ionisation potential $\leq$ 11 then higher oxidation state is stable.

Ex. $Mg \longrightarrow Mg^+$ Ist ionisation potential

$Mg^+ \longrightarrow Mg^{+2}$ IInd ionisation potential $\}$ 7.4 eV

Difference of ionisation potential < 11 eV So Mg^{+2} is more stable.

$Al \longrightarrow Al^+$

$\}$ 12.8 eV So Al^+ is more stable

$Al^+ \longrightarrow Al^{+2}$

$Al^{+2} \longrightarrow Al^{+3}$ $\}$ 6.0 eV So Al^{+3} is more stable

$\{$ Al^+ is stable only in gaseous state

Al^{+3} is stable in liquid and solid state.

Ionization energy in KJ mol^{-1}

Group Period	1	2											13	14	15	16	17	18
1	H 1311																H 1311	He
2	2372 Li	Be											B	C	N	O	F	Ne
3	520 Na	899 Mg	3	4	5	6	Group 7	8	9	10	11	12	801 Al	1086 Si	1403 P	1314 S	1681 Cl	2080 Ar
4	496 K	737 Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	577 Ga	786 Ge	1012 As	999 Se	1255 Br	1521 Kr
5	419 Rb	590 Sr	631 Y	656 Zr	650 Nb	652 Mo	717 Tc	762 Ru	758 Rh	736 Pd	745 Ag	906 Cd	579 In	760 Sn	947 Sb	941 Te	1142 I	1351 Xe
6	403 Cs	549 Ba	616 La	674 Hf	664 Ta	685 W	703 Re	711 Os	720 Ir	804 Pt	731 Au	876 Hg	558 Tl	708 Pb	834 Bi	869 Po	1191 At	1170 Rn
7	376 Fr	503 Ra	541 Ac	760	760	770	759	840	900	870	889	1007	589	715	703	813	912	1037

Ex. From each set, choose the atom which has the largest ionization enthalpy and explain your answer with suitable reasons. (a) F, O, N (b) Mg, P, Ar

Sol. (a) Fluorine (F) has the largest ionization enthalpy because in moving from left to right in a period, atomic size decreases and electrons are held more tightly. Since F has the smallest size and maximum nuclear charge. It has the largest ionization enthalpy among these elements.

(b) Argon (Ar) has the largest ionization enthalpy as argon has extra stable fully filled configuration.

Ex. First and second ionisation energies of magnesium are 7.646 eV and 15.035 eV respectively. The amount of energy in kJ needed to convert all the atoms of magnesium into Mg^{2+} ions present in 12 mg of magnesium vapour will be ? [Given 1 eV = 96.5 kJ ml^{-1}].

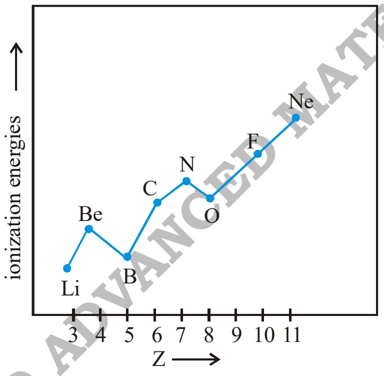
Sol. Total energy needed to convert one Mg atom into Mg^{2+} gas ion,

$= IE_i + IE_{II} = 22.681\ eV = 2188.6\ kJ\ mol^{-1}$.

$\Rightarrow$ 12 mg of Mg $= 0.5 \times 10^{-3}$ mole.

$\therefore$ Total energy $= 0.5 \times 10^{-3} \times 2188.6 = 1.0943$ kJ **Ans.**

Ex. Following graph shows variation of ionization energies with atomic number in second period (Li–Ne). Value of ionization energies of Na(11) will be –



(A) above Ne (B) below Ne but above O. (C) below Li (D) between N and O.

Sol. Na is 3rd period element and is bigger than Li. The distance between the nucleus and outer most electron is more as compared to Li. Thus the outer most electron is loosely bound with nucleus and removal of electron is easier. So option (C) is correct.

Ex. $M(g) \rightarrow M^+(g) + e^-$; $\Delta H = 100$ eV.

$M(g) \rightarrow M^{2+}(g) + 2e^-$; $\Delta H = 250$ eV.

Which is/are correct statement(s) ?

(A) IE_1 of M(g) is 100 eV (B) IE_1 of $M^+(g)$ is 150 eV.

(C) IE_2 of M(g) is 250 eV. (D) IE_2 of M(g) is 150 eV.

Sol. $M(g) \rightarrow M^+(g) + e^-$ IE_1 of M

$M^+ \rightarrow M^{2+} + e^-$ IE_2 of M but IE_1 of M^+

$M \rightarrow M^{2+} + 2e^-$ $(IE_1 + IE_2)$.

Ans. (A,B,D)

ELECTRON AFFINITY

(a) The amount of energy released when an electron is added to the outermost shell of one mole of an isolated gaseous atom in its lower energy state.

(b) The positive value of the electron affinity indicates that the process, i.e. $X_{(g)} + e^- \rightarrow X^-_{(g)}$, is exothermic and the negative value indicates the process to be an endothermic one. Thus the convention accepted in defining the electron affinity apparently contradicts the established convention in the thermodynamics.

Thus, $F_{(g)} + e^- \rightarrow F^-_{(g)}$, $\Delta H = -328$ kJ mol^{-1}, and EA = 328 kJ mol^{-1} and

$N_{(g)} + e^- \rightarrow N^-_{(g)}$, $\Delta H = +31$ kJ mol^{-1}, and EA = –31 kJ mol^{-1}

(c) Electron affinity just defined is actually first electron affinity since it corresponds to the addition of one electron only. In the process of adding further electron, the second electron will be added to gaseous anion against the electrostatic repulsion between the electron being added and the gaseous anion. Energy instead of being released is supplied for the addition of an electron to an anion.

$A_{(g)} + e^- \rightarrow A^-_{(g)} + E_1$; $A^-_{(g)} + e^- + \text{Energy supplied} \rightarrow A^{2-}_{(g)}$

Factors affecting electron affinity

(A)	Atomic size	Electron Affinity $\propto \dfrac{1}{\text{Atomic size}}$
(B)	Screening effect	Electron Affinity $\propto \dfrac{1}{\text{Screening effect}}$
(C)	Effective nuclear charge (Z_{eff})	Electron Affinity $\propto$ Zeff

(d) **Stability of completely filled or half filled orbitals**

Electron affinity of filled or half filled orbital is very less or zero.

Ex. (I) Elements (He, Ne, Ar , ...), Electron Affinity = Zero ($\Delta egH = +$ ve)

(II) Elements (Be, Mg, Ca, ...), Electron Affinity = $\simeq$ Zero ($\Delta egH = +$ ve)

(III) Elements (N, P, As, ...), Electron Affinity = Very less

Periodic variation of electron affinity

(a) **In a period :** The effective nuclear charge increases and the size decreases with the increase of atomic number in a period. This is why, the electron affinity in general increases in a period. In fact, it reaches the climax for the Group VIIA (i.e. halogens) elements.

Electron affinities (kJ mol^{-1})

IA	IIA	IIIA	IVA	VA	VIA	VIIA	Noble gas
Li 59.8	Be –60	B 27	C 122	N –31	O 141	F 328	Ne –90
Na 53	Mg –60	Al 44	Si 134	P 72	S 200	Cl 349	Ar 0
		Ga 29	Ge 120	As 77	Se 195	Br 324	

(b) **In a group :** For the representative elements, in moving down in a group generally the electron affinity falls down with the increases of atomic number because the effective nuclear charge Zeff at periphery per electron remain almost same but the size gradually increases due to addition of new shell.

KEY POINTS

Exceptions

(I) Electron affinity values of nitrogen and phosphorous (VA) are lesser than the electron affinity values of carbon and silicon respectively. It is due to the comparatively stable half filled configuration (np^3) of nitrogen and phosphorus and the tendency to acquire the stable np^3 configuration by the gain of one electron in carbon and silicon (np^2).

(II) The theoretical value of the electron affinity of zero group i.e. inert gas elements is zero due to stable s^2p^6 configuration.

(III) $F < Cl$, $O < S$, $N < P$, $B < Al$. Here it is interesting to note that the electron affinity sequence is in the opposite order as is expected from the size sequence. To explain the observed sequence of electron affinity, we are to consider the other factors. Though the electrostatic attractive pull towards the nucleus favours the 2nd period elements more compared to the corresponding 3rd period elements, the added electron creates an unfavourable effect, i.e. electron-electron repulsion, which is more for the 2nd period elements because of their smaller sizes. This repulsive force is not so large in the 3rd period elements because of their larger size.

Ex. Consider the elements N, P, O and S and arrange them in order of increasing negative electron gain enthalpy.

Sol. Order of increasing negative electron gain enthalpy is $N < P < O < S$. For detail refer text.

Ex. Why do halogens have high electron gain enthalpies (i.e. $-\Delta_{eg}H^{\ominus}$) ?

Sol. The valence shell electronic configuration of halogens is ns^2np^5 and thus they require one electron to acquire the stable noble gas configuration ns^2np^6. Because of this they have strong tendency to accept an additional electron and hence have high electron gain enthalpies.

Ex. Which will have the maximum value of electron affinity O^x, O^y, O^z [x,y and z respectively are 0, – 1 and –2]?

(A) O^x (B) O^y (C) O^z (D) All have equal.

Sol. Being neutral atom oxygen will have higher electron affinity as there is electrostatic repulsion between additional electron and negative ion in case of O^- and O^{2-}. So option (A) is correct.

Ex. The amount of energy when million atoms of iodine are completely converted into I^- ions in the vapour state according to the equation, $I(g) + e^-(g) \rightarrow I^-(g)$ is 5.0×10^{-13} J.

Calculate the electron gain enthalpy of iodine in terms of kJ mol^{-1} and eV per atom.

Sol. The electron gain enthalpy of iodine is equal to the energy released when 1 mole of iodine atoms in vapour state are converted into I^- ions.

$$= -\frac{5.0 \times 10^{-13} \times 6.023 \times 10^{23}}{10^6} = -30.1 \times 10^4 \text{ J} = -301 \text{ kJ}.$$

$$\text{Electron gain enthalpy of iodine in eV per atom} = \frac{-301}{96.5} = -3.12.$$

Ex. Account for the large decrease in electron affinity between Li and Be despite the increase in nuclear charge.

Sol. The electron configurations of Li and Be are $[He]2s^1$ and $[He]2s^2$, respectively. The additional electron enters the 2s orbital of Li but the 2p orbital of Be and hence is much less tightly bound. In fact, the nuclear charge is so well shielded in Be that electron gain is endothermic.

ELECTRONEGATIVITY (EN)

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

(b) Electronegativity and Electron affinity both have tendency to attract electrons but electron affinity is for isolated atoms. Where as electronegativity is for bonded atoms.

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

(I) $A-B \longrightarrow A^{-}: + B^{+}$ (Electronegativity A > Electronegativity B)

or (II) $A-B \longrightarrow A^{+} + :B^{-}$ (Electronegativity A < Electronegativity B)

depending on their tendency to attract bonded electron.

(d) There is no unit of electronegativity because it is a relative value

(e) Electronegativity is property of a bonded atom not an isolated gaseous atom so no energy will be liberated or absorbed.

Difference between electronegativity and Electron Affinity

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 changed 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 $mole^{-1}$

Note : Small atoms are normally having more electronegativity than larger atoms.

Pauling's Scale

Pauling related the resonance energy (Δ_{AB}) of a molecule AB with the electronegativities of the atoms A and B. If x_A and x_B are the electronegativities of atoms A and B respectively then

$$0.208\sqrt{\Delta_{AB}} = x_A - x_B \text{ if } x_A > x_B$$

$$\text{or } \Delta_{AB} = 23.06\,(x_A - x_B)^2$$

$\Delta_{AB} = E_{AB(experimental)} - E_{AB(theoritical)}$ where E_{A-B} is the energy of A–B bond.

$$E_{AB(theoritical)} = \sqrt{(E_{A-A} \times E_{B-B})}$$

In a purely covalent molecule, AB, the experimental and theoritical values of bond energy A–B are equal,

So $\Delta_{AB} = 0$

or $0 = 23.06\,(x_A - x_B)^2$

or $x_A = x_B$

In an ionic molecule AB, $E_{AB(experimental)}$ is more than $E_{AB(theoritical)}$

Pauling assumed the electronegativity value of fluorine 4 and calculated the electronegativity values of other elements from this value.

						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 1.8	Bi 1.9	Po 2.0	
Fr 0.7	Ra 0.9					

In Pauling's scale, elements having almost same electronegativity are-

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

Mulliken's Electronegativity

According to Mulliken, the electronegativity of an element is the average value of its ionisation potential and electron affinity.

$$\text{or} \qquad \text{electro-negativity} = \frac{\text{Electron affinity} + \text{Ionisation potential}}{2}$$

When both are expressed in electron volt

Alred Rochow's Electronegativity : The electronegativity of an element is the electrostatic force of attraction between the electron present on the circumference of outermost shell of this atoms and the atomic nucleus. If the distance between the circumference of outermost shell & the nucleus is r and the effective nuclear charge Z_{eff} then

$$\text{Electron-negativity} = \frac{Z_{eff}\, e^2}{r^2} = \frac{0.359\, Z_{eff}}{r^2} + 0.744 \quad \Rightarrow \quad Z_{eff} = Z - \sigma$$

Z = The actual number of charge present in the nucleus i.e. number of protons and σ = Shielding constant.

Factors Affecting electronegativity

(A)	**Atomic size**	(B)	**Effective nuclear charge (Z_{eff})**
	$\text{Electronegativity} \propto \dfrac{1}{\text{Atomic size}}$		$\text{Electronegativity} \propto \text{Zeff}$
(C)	**Hybridisation state of an atom**	(D)	**Oxidation state**
	Electronegativity $\propto$ % s character in hybridised atom sp > sp^2 > sp^3 s character 50% 33% 25% Electronegativity 3.25 2.75 2.5		Electronegativity $\propto$ oxidation state $Mn^{+2} < Mn^{+4} < Mn^{+7}$ $O^{-2} < O^{-1} < O < O^{+1} < O^{+2}$ $Fe < Fe^{+2} < Fe^{+3}$

KEY POINTS

(I) s-orbital is nearer to nucleus so by increasing s-character in hybridisation state, Electronegativity also increases.

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

Periodic Variation

(a) **In a period :** In moving from left to right in a period, the size gradually contracts and the effective nuclear charge increases. This is why, the electronegativity gradually increases with the increase of atomic number in a period up to the group VIIA elements.

(b) **In a group :** In the representative elements, in moving down in a group, the size increase while Zeff per electron at the periphery remains almost constant. This is why, the electronegativity generally falls in a group with the increase of atomic number.

KEY POINTS

Exceptions

(a) In **'0' group** – Electronegativity is always zero, because inert gas do not form molecule.

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

Note: So effect of nuclear charge balanced the effect of increase in number of shell.

Electronegativity of F > Cl but Electron affinity of Cl > F

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

(d) In IIIA group, value of electronegativity increases down the group, because of transition contraction Electronegativity of Ga > Electronegativity of Al

Application of electronegativity

(a) **Metallic and non metallic nature**

Low electronegativity $\xrightarrow{\quad}$ Metals

High electronegativity $\xrightarrow{\quad}$ Non Metals

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

(b) **Bond length**

$\boxed{\Delta EN \propto \dfrac{1}{\text{Bond length}}}$ Here Δ EN = difference in electronegativities of bonded atoms

Note : HF has minimum bond length as expected because of much difference in the electronegativities of H and F.

(c) **Acidic strength of hydrides**

$NH_3 < H_2O < HF$

In a particular period acidic strength of hydride is depends on electronegativity of M in M – H bond. Higher the electronegativity of M greater the acidic strength of hydride.

(d) **Nature of bonds**

(I) According to 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 $X_A - X_B \geq 2.1$ Ionic % > 50% i.e. Ionic bond

If $X_A - X_B \leq 2.1$ Ionic % < 50% i.e. covalent bond

(II) According to Gallis

$X_A - X_B \geq 1.7$ Ionic

$X_A - X_B \leq 1.7$ Covalent

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

Ex. H—H, F—F

If $X_A > X_B$ and difference of electronegativities is small then

$A^{\delta-}$ —— $B^{\delta+}$ bond will be polar covalent

(e) Nature of hydroxides

(I) As per Gallis, In AOH if electronegativity of A is more than 1.7 (Non metal) then it is acidic in nature.

(II) If electronegativity of 'A' is less than 1.7 (metal) then AOH will be basic in nature

Ex.

	NaOH	ClOH
X_A	0.9	3.0
Nature	Basic	Acidic

(III) If $X_A - X_0 \geq X_0 - X_H$ then AO bond will be more polar and will break up as

$A \text{——} OH \longrightarrow A^+ + OH^-$ It shows basic nature

(IV) If $X_A - X_O \leq X_O - X_H$

$A \text{—} O \text{——} H \longrightarrow H^+ + AO^-$ It shows Acidic nature

Ex. **In NaOH**

$X_O - X_{Na}\,(2.6) > X_O - X_H\,(1.4)$ So hydroxide is basic

In ClOH

$X_O - X_{Cl}\,(0.5) < X_O - X_H\,(1.4)$ So hydroxide is acidic

(f) 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

(I) Along a period acidic nature of oxide increases.

(II) Down the group basic nature of oxide increases

Na_2O MgO Al_2O_3 SiO_2 P_4O_{10} SO_3 Cl_2O_7

$\longleftrightarrow$

$X_A - X_O > 2.3$ Basic

$X_A - X_O = 2.3$ Amphoteric

$X_A - X_O < 2.3$ Acidic

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

$NO_2 > ZnO > K_2O$

$\xrightarrow{\hspace{3cm}}$

acidic character decreases

KEY POINTS

Note: BeO, Al_2O_3, ZnO, SnO, PbO, SnO_2, PbO_2, Sb_2O_3 etc. are amphoteric oxides.

CO, H_2O, NO, N_2O etc. are neutral oxides.

$\boxed{\text{Acidic strength of oxide and oxyacid} \propto \text{Electronegativity}}$

B_2O_3 CO_2 N_2O_5

$\longrightarrow$

EN increase, acidic nature increase.

$SO_3 > SeO_3 > TeO_3$	$HNO_3 > H_2CO_3 > H_3BO_3$	$HClO_4 > HBrO_4 > HIO_4$
$H_2SO_4 > H_2SeO_4 > H_2TeO_4$	$H_3PO_4 > H_3AsO_4 > H_3SbO_4$	$N_2O_5 > P_2O_5 > As_2O_5$
$H_2SO_3 > H_2SeO_3 > H_2TeO_3$	$HOF > HOCl > HOBr > HOI$	$N_2O_3 > P_2O_3 > As_2O_3 > Sb_2O_3$

Acidic properties increases with increasing oxidation state of an element when central atom is same. **Ex.**

(I) $Sb_2O_5 > Sb_2O_3$

(II) $HClO_4 > HClO_3 > HClO_2 > HClO$

(III) $HNO_3 > HNO_2$

(IV) $SO_3 > SO_2$

(V) $N_2O_5 > N_2O_3$

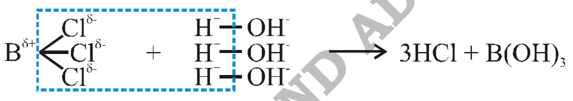
(VI) $H_2SO_4 > H_2SO_3$

(g) Hydrolysis products of AX

Where A = Other element and X = Halogen

(I) If electronegativity of X > Electronegativity of A then on hydrolysis product will be HX.

Ex. (BCl_3), Electronegativity of Cl > Electronegativity of B

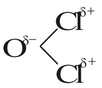


$PCl_3 + 3HOH \longrightarrow 3HCl + H_3PO_3$

$PCl_3.Cl_2 + H_2O \longrightarrow 2HCl + POCl_3$

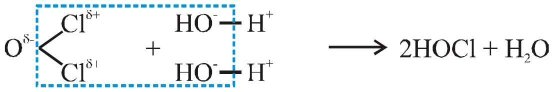
$POCl_3 + 3HOH \longrightarrow 3HCl + H_3PO_4$

(II) If electronegativity of X < electronegativity of A then on hydrolysis product will be HOX (hypohalous acid)

Ex. Cl_2O 

Here electronegativity of O > Cl

So On hydrolysis –



(h) Nomenclature of inorganic compounds

Prefix —— less electronegative element

Suffix —— More electronegative element

Ex. Cl_2O(Right) OCl_2(Wrong)

In Dichloroxide the electronegativity of Cl is less than 'O' i.e. why Cl is in prefix position.

OF_2	Oxygen difluoride
ICl	Iodine chloride
NH_3	Exception (Here H is less electronegative but it is suffix)

Ex. Calculate the electronegativity of carbon from the following data :

$E_{H-H} = 104.2$ kcal mol^{-1} , $E_{C-C} = 83.1$ kcal mol^{-1}

$E_{C-H} = 98.8$ kcal mol^{-1} , $X_H = 2.1$

Sol. Let the electronegativity of carbon be X_C , applying Pauling equation,

$$X_C - X_H = 0.208\,[E_{C-H} - \frac{1}{2}(E_{C-C} + E_{H-H})]^{1/2}$$

$$X_C - 2.1 = 0.208\,[98.8 - \frac{1}{2}(83.1 + 104.2)]^{1/2}$$

$\Rightarrow X_C = 2.5$ **Ans.** 2.5

Ex. If electronegativity of x be 3.2 and that of y be 2.2, the percentage ionic character of xy is –

(A) 19.5 **(B)** 18.5 **(C)** 9.5 **(D)** 29.5.

Sol. $EN_x - EN_y = 3.2 - 2.2 = 1.$

$\Delta = 1$

[Δ = difference of electronegativity values between x and y].

% ionic character $= 16\Delta + 3.5\Delta^2 = 19.5.$ **Ans.** (A)

Ex. Which of the following relation is correct ?

(A) 2 IP – E.A. – EN = 0 **(B)** 2EN – IP – EA = 0

(C) 2EA – IP – EA = 0 **(D)** EN – IP – EA = 0

Sol. On Mulliken's scale : $E.N. = \frac{I.P. + E.A.}{2}$

$2E.N. - I.P. - E.A. = 0$ **Ans.** (B)

DENSITY

(a) In a group volume of an atom increases along with atomic weight but atomic weight increases more than atomic volume, So density increases in a group.

(b) Factors affecting density are

(I) Inner shell configuration (II) Atomic weight (III) Packing capacity (IV) Chemical bond

Periodic Variation

(a) **In period** density first increases till maximum and then decreases. (s-block to d-block increases, d-block to p-block decreases)

(b) **In group** from top to bottom density increases regularly

Ex. In VIIA group – F and Cl are gases (Low density)

Br is liquid (density 3.19 g/cm^{-3})

I is solid (density 4.94 g/cm^{-3})

From s-block to p-block packing capacity and strength of bond increases.

KEY POINTS

Exceptions

(a) The density of sodium and magnesium is greater than K and Ca respectively. This can be explained by inner shell configuration.

In K, $3d^0$ orbital is vacant, 3^{rd} orbit has the capacity to hold $18e^-$ but it has only $8e^-$. Due to empty orbitals atomic volumes increases and so density decreases

(b) Li is the lightest metal known.

(c) Highest density of solid metal Ir (22.63 gm/cc) and Os (22.6 gm/cc) and liquid metal Hg(13.6 gm/cc)

Order of density of elemetns

Li < K < Na < Rb < Cs	Sc < Y < La	Hg < Au < Os < Ir	Cr < Fe < Cu
Ca < Mg < Be < Sr < Ba	Ti < Zr << Hf	Sc < Zn < Fe < Hg	Fe < Cu < Pb
He < Ar < Kr < Rn	Zn < Cd < Hg		Fe < Hg < Au

BOILING POINTS AND MELTING POINTS

(a) It is a property of aggregate of atoms and not of a single atom.

i.e. why it is a molecular property.

(b) Melting point of a solid depend on –

(I) Structure of solid (II) Chemical bond (III) Bond energy

(c) Structure of solid made up of (I) atomic solid (II) molecular solid (III) metallic solid.

Order of decreasing melting point is : **Atomic solid > Metallic solid > Molecular solid**

(d) Chemical bond among particles in solid. Order of bond strength is –

Covalent bond > Metallic bond > Vander waal bond

(e) Bond energy – Covalent solids like diamond, SiO_2 etc have only covalent bonds between atoms so their bond energy is higher than molecular or metallic solids.

Periodic variation

In period from left to right boiling point and melting point first increases then decreases.

Alkali metals – Crystal structure BCC (low boiling point & melting point)

Transition metals – ...."..... FCC (High boiling point & melting point)

Inert gases .... Lowest boiling point & melting point (Vander waal force)
Transition elements .... Highest melting point

Metals
W(Tungston) maximum melting point (3410° C)
Hg(Mercury) lowest melting point (– 38° C)

Non metals
Carbon (In the form of diamond) Highest melting point (3727° C)
Helium Lowest melting point (- 270° C)

In Group

(a) In **s-block** elements boiling point & melting point decreases down the group.

Li, Na (solid) → Cs, Fr (liquid)

It is due to weak cohesive energy of metallic bond in IA.

(b) In **d-block** elements boiling point & melting point increases down the group (due to lanthanide contraction, zeff increases and hence bond energy increases)

(c) In **p-block** elements

(I) From IIIA – IVA group boiling point & melting point decreases down the group and from VA to '0' group, boiling point & melting point increases down the group. (Atomic or molecular weigh $\propto$ vander waal force)

(II) boiling point and melting point of monoatomic molecules are lesser than diatomic molecules. '0' group < Halogens

(III) Atomic solid non metals like B, C and Si has higher boiling point and melting point due to strong covalent bond.

(IV) Boiling point & melting point of molecular solids are less because of weaker vander waal force among molecules **Ex.** I_2.

Order of melting point

Li > Na > K > Cs	C > Si > Ge	He < Ne < Ar < Rn
Be > Ca > Sr > Ba > Mg	C > B > Si	Zn > Cd > Hg
$F_2 < Cl_2 < Br_2 < I_2$	B > Al > Ga	Cu > Au > Ag

Ex. Arrange the following in decreasing basic nature LiOH, NaOH, KOH, CsOH.

LiOH, NaOH, KOH, CsOH

Sol. The basic nature of hydroxides of elements of group 1st increases on descending the group with increase in size of cation as CsOH > RbOH > KOH > NaOH > LiOH.

Ex. Arrange the following compounds in increasing order of acidic strength Al_2O_3, SiO_2, P_2O_3 and SO_2.

Sol. Along the period acidity of oxides increases as the electronegativity difference between element and oxygen decreases.

Ans. $Al_2O_3 < SiO_2 < P_2O_3 < SO_2$.

BASIC INORGANIC NOMENCLATURE

Oxidation Number

- It is an imaginary or apparent charge developed over atom of an element when it goes from its elemental free state to combined state in molecules.
- It is calculated on basis of an arbitrary set of rules.
- It is a relative charge in a particular bonded state.
- In order to keep track of electron-shifts in chemical reactions involving formation of compounds, a more practical method of using oxidation number has been developed.
- In this method, it is always assumed that there is a complete transfer of electron from a less electronegative atom to a more electronegative atom.

Rules governing oxidation number

The following rules are helpful in calculating oxidation number of the elements in their different compounds. It is to be remembered that the basis of these rule is the electronegativity of the element .

- **Fluorine atom**

  Fluorine is most electronegative atom (known). It always has oxidation number equal to –1 in all its compounds

- **Oxygen atom**

  In general and as well as in its oxides , oxygen atom has oxidation number equal to –2.

In case of (I) peroxide (e.g. H_2O_2 , Na_2O_2) is –1,

(II) super oxide (e.g. KO_2) is –1/2

(III) ozonide (e.g. KO_3) is –1/3

(IV) in OF_2 is + 2 & in O_2F_2 is +1

- **Hydrogen atom**

  In general, H atom has oxidation number equal to +1. But in metallic hydrides (e.g. NaH, KH), it is –1.

- **Halogen atom**

  In general, all halogen atoms (Cl, Br , I) have oxidation number equal to –1.

  But if halogen atom is attached with a more electronegative atom than halogen atom, then it will show positive oxidation numbers.

  Ex. $K\overset{+5}{Cl}O_3$, $H\overset{+5}{I}O_3$, $H\overset{+7}{Cl}O_4$, $K\overset{+5}{Br}O_3$

- **Metals**

  (a) Alkali metal (Li , Na, K, Rb, ...) always have oxidation number +1

  (b) Alkaline earth metal (Be , Mg , Ca ....) always have oxidation number +2.

  (c) Aluminium always has +3 oxidation number

Note : Metal may have positive or zero oxidation number

- Oxidation number of an element in free state or in allotropic forms is always zero

  Ex. $\overset{0}{O_2}$, $\overset{0}{S_8}$, $\overset{0}{P_4}$, $\overset{0}{O_3}$

- Sum of the oxidation numbers of atoms of all elements in a molecule is zero.
- Sum of the oxidation numbers of atoms of all elements in an ion is equal to the charge on the ion .
- If the group number of an element in modern periodic table is ***n,*** then its oxidation number may vary from

  (n – 10) to ***(n – 18)*** (but it is mainly applicable for p-block elements)

  Ex. N- atom belongs to 15th group in the periodic table, therefore as per rule, its oxidation number may vary from

  –3 to +5 ($\overset{-3}{N}H_3$, $\overset{+2}{N}O$, $\overset{+3}{N_2}O_3$, $\overset{+4}{N}O_2$, $\overset{+5}{N_2}O_5$)

- The maximum possible oxidation number of any element in a compound is never more than the number of electrons in valence shell.(but it is mainly applicable for p-block elements)

List of common oxidation sate of an element in Periodic Table

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H **+1** –1																	2 **He**
3 Li +1	3 Be +2											5 **B** **+3** –3	6 **C** **+4** +2 –4 etc.	7 **N** +5 +4 +3 +1 –3 **0** etc.	8 **O** +2 –1/2 –1 **–2**	9 **F** **–1**	10 **Ne**
11 Na +1	12 Mg +2											13 **Al** +3	14 **Si** +4 **–4**	15 **P** **+5** +3 +1 –3	16 **S** **+6** +4 +2 –2	17 **Cl** +5 +7 +3 +1 **–1** etc.	18 **Ar** **0**
19 **K** +1	20 **Ca** +2	21 **Sc** +2 **+3**	22 **Ti** +2 **+3** **+4**	23 **V** +2 **+3** +4 +5	24 **Cr** +2 **+3** +4 +5 +6	25 **Mn** **+2** +3 **+4** +5 +6 +7	26 **Fe** **+2** +3 +4 +5 +6	27 **Co** **+2** +3 +4 +5	28 **Ni** **+2** +3 +4	29 **Cu** **+2** +3	30 **Zn** **+2**	31 **Ga** +3	32 **Ge** +4 **–4**	33 **As** **+5** **+3** –3	34 **Se** **+6** +4 –2	35 **Br** +5 +3 +1 **–1**	36 **Kr** +4 +2 **0**
37 **Rb** +1	38 **Sr** +2											49 **In** **+3** +1	50 **Sn** **+4** +2	51 **Sb** **+5** **+3** –3	52 **Te** **+6** +4 –2	53 **I** +7 +5 +1 **0** **–1**	54 **Xe** +6 +4 +2 **0**
55 **Cs** +1	56 **Ba** +2											81 **Tl** +3 **+1**	82 **Pb** +4 **+2**	83 **Bi** +5 **+3**	84 **Po**	85 **At**	86 **Rn**

Elements

General Rule : The names of metals generally end in-ium or-um (examples are sodium, potassium, aluminum, and magnesium).

The exceptions are metals that were used and named in ancient times, such as iron, copper, and gold.

The names of nonmetals frequently end in-ine,-on, or -gen (such as iodine, argon, and oxygen.)

Given the names of the constituent elements and common ions, most of the common inorganic compounds can be named using the rules presented below.

Acids

Acids are normally classified in two groups, hydracids and oxyacids

Hydracids

Hydracids are acids which contain hydrogen and a non-metal , but no oxygen.

General Rule : The names of hydracids have the prefix hydro-(sometimes shortened to hydr-) and the suffix-ic attached to the stem based on the names of the constituent elements (other than hydrogen.)

For example : HCl (made of hydrogen and chlorine) is hydrochloric acid; HBr (made of hydrogen and bromine) is hydrobromic acid; HI (made of hydrogen and iodine) is hydroiodic acid; HCN (made of hydrogen, carbon , and nitrogen) is hydrocyanic acid; and H_2S (made of hydrogen and sulfur) is hydrosulfuric acid.

Oxoacids or Oxyacids

The acids which contain hydrogen , oxygen and a metal or non-metal.

In this case , more than one possibilities arise because of the presence of different number of oxygen atoms. An example of such an oxoacid series is as follows: HClO, $HClO_2$, $HClO_3$, $HClO_4$. All these contains same three elements but differ in the number of oxygen atoms present.

General Rule - 1

If a class of acids contains only one member, its name is given the suffix-ic.

For example, hydrogen , carbon and oxygen combine to form only one acid i.e. H_2CO_3. It is called carbonic acid (carbonic acid.)

General Rule - 2

If an acid series contains two acids, such as H_2SO_4 and H_2SO_3, the acid containing more oxygen atoms is given the suffix -ic, while the acid with fewer oxygen atoms is given the suffix-ous.

For example : H_2SO_4 is sulfuric acid , and H_2SO_3 is sulfurous acid

Similarly, HNO_3 is nitric acid and HNO_2 is nitrous acid.

In the case of an extensive acid series (such as HClO, $HClO_2$, $HClO_3$, $HClO_4$), the acid with the one oxygen atoms lesser than -ous acid is given the prefix hypo- and the suffix -ous, and the acid with the one oxygen atom more than the -ic acid is given the prefix per and a suffix-ic.

In the above example, HClO is hypochlorous acid $HClO_2$ is chlorous acid, $HClO_3$ is chloric acid, and $HClO_4$ is perchloric acid.

A very valid doubt may arise, “How to decide which one is -ic acid to begin with ? This is somewhat arbitrary and has no unique answer. It depends upon the name given to it by old chemists, when no systematic nomenclature existed. Now these names have becomes so popular that it is difficult to change them. Do not worry, as you go through these notes, you will get a sufficient idea regarding these names.

Cations (Positive ions)

Metal atoms with single positive charge

Rule : Names of positive ions end in-ium if the ion has only one oxidation state (Only one level of net charge). For example , the positive ion of sodium is Na^+ (sodium ion), and the positive ion of aluminium is Al^{3+} (aluminium ion).

Metal atoms with more than one possible charges

Rule : If the cation has variable valency (charge), charge is specified in roman numerals in round brackets immediately after the name of metal atom. For example , Sn^{2+} is written as tin (II) ion.

Alternately, the less positive ion ends in -ous, and the more positive ion ends in-ic . For instance, the two positive ions of copper are Cu^{+} (cuprous) and Cu^{2+} (cupric). The oxidation state of a positive ion can also be designated by placing a Roman numeral after the name of the elements. These positive ions of copper can also be written as copper (I) and copper (II), respectively.

Ions	Name
Cu^{+}	cuprous ion
Sn^{2+}	Stannous ion
Sn^{4+}	Stannic ion
Fe^{3+}	Ferric ion
Fe^{2+}	Ferrous ion

General Rule-3

Suffix-nium is often used with cations containing non metals.

For example , the positive ion of ammonia is NH_4^{+} (ammonium) and the positive ion of water (H_2O) is H_3O^{+} or H^{+} (hydronium).

Remember these names !

NO_2^{+} : nitronium **NO^{+} : nitrosonium** **H_3O^{+} : hydronium**

From NH_3 ammonia is derived NH_4^{+} : **ammonium.** Similarly.

N_2H_4 : hydrazine $\longrightarrow$ N_2H_5 : **hydrazinium**

$C_6H_5NH_2$: aniline $\longrightarrow$ $C_6H_5NH_3^{+}$: **anilinium**

C_5H_5N : pyridine $\longrightarrow$ $C_5H_5NH^{+}$: **pyridinium**

Anions (Negative Ions)

Anions can always be looked upon as ions derived from acids by removal of one or more protons. Accordingly, anions can be classified as follows.

Anions derived from hydracids

Rule : Names of negative ions from hydracids end in –ide.

For example , Cl^{-} (chloride) from HCl, and CN^{-} (cyanide) from HCN. Following examples will give you a better insight in this nomenclature. It is also useful to remember them.

Remember these names

F^{-}	fluoride	from HF	
Cl^{-}	chloride	from HCl	
Br^{-}	bromide	from HBr	
I^{-}	iodide	from HI	
O^{2-}	oxide	S^{2-}	sulphide
N^{3-}	nitride	P^{3-}	phosphide
As^{3-}	arsenide	Cu^{-}	cupride
H^{-}	hydride	Au^{-}	auride

Anions derived from oxyacids

Anion derived from an oxyacid by removal of one or more H^+ ions is termed as oxyanion.

Rule : If the oxyacid is – ic acid, suffix - ate is used with oxy-anion.

For example

CO_3^{2-}	carbonate (from H_2CO_3)
ZnO_2^{2-}	zincate
SiO_3^{2-}	silicate

Rule : If the oxyacid is - ous acid, suffix -ite is used with oxy-anion.

For example

NO_2^- (nitrite) is derived from HNO_2 (nitrous acid), and SO_3^{2-} (sulfite) is derived from H_2SO_3 (sulfurous acid)

Rule : If the oxyacid has prefixes per-or hypo–, the oxyanion will have same prefixes.

For example,

ClO_4^- percholrate ion from $HClO_4$, perchloric acid

ClO^- hypochlorite ion from HClO, hypochlorous acid

Remember these names !

SO_4^{2-}	sulphate SO_3^{2-}	Sulphite NO_3^-	nitrate,	NO_2^-	nitrite
SnO_3^{2-}	stannate SnO_2^{2-}	stannite, PbO_3^{2-}	plumbate,	PbO_2^{2-}	plumbite

Anions containing replacable hydrogen ions

Polyprotic acid. Any acid containing more than one replacable hydrogens is said to be a polyprotic acid.

Replacable hydrogens. H atoms which can be lost as H^+ in reactions with a base.

H atoms connected to O atoms in oxyacids are all replacable. If all the replacable hydrogens are removed, we obtain the anions discussed in the sections above.

However, in all the polyprotic acids it is always possible to remove less than the maximum number of replacable hydrogens . e.g. H_3PO_4 is triprotic. We can remove one, two or three H^+ ions from it to generate $H_2PO_4^-$, HPO_4^{2-} and PO_4^{3-}. You are already familiar with phosphate ion, PO_3^-. The other two anions, $H_2PO_4^-$ and HPO_4^{2-} still contain H atoms that are replacable. We consider their nomenclature in this section.

Rule-1 : A prefix bi- (old notation) or hydrogen – (IUPAC notation) is attached to the name of anion.

Rule -2 : For triprotic or higher acids, numerical prefixes (e.g. mono, bi, tri) are also used to indicate the number of replacable H atoms left in the sample).

Ex. HCO_3^- is bicarbonate or hydrogen carbonate

HSO_3^- bisulphite or hydrogen sulphite

HS^- bisulphide or hydrogen sulphide etc.

when anion has – 3 charge , e.g. PO_4^{3-} then following possibilities arise

HPO_4^{2-} monohydrogen phosphate, $H_2PO_4^-$ dihydrogen phosphate

KEY POINTS

Miscellaneous Anions (To be committed to memory)

NH_2^- amide, NH^{2-} imide, N^{3-} nitride, N_3^- azide, O_2^{2-} peroxide, O_2^- superoxide, O_3^- ozonide

CrO_4^{2-} chromate, $Cr_2O_7^{2-}$ dichromate, MnO_4^{2-} manganate, MnO_4^- permanganate , CN^- cyanide

OCN^- cyanate, SCN^- thiocyanate, $S_2O_3^{2-}$ thiosulphate, CH_3COO^- acetate, $C_2O_4^{2-}$ oxalate

$[Fe(CN)_6]^{3-}$ ferricyanide ,$[Fe(CN)_6]^{4-}$ ferrocyanide

Difference between Atoms and ions

We have been talking for quite sometime. The distinction between atoms and ions, though simple, must be very clear in your mind. The following are the points summarise this difference.

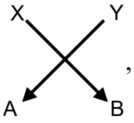
Method of writing formula of an ionic compound

In order to write the formula of an ionic compound which is made up of two ions (simple or polyatomic) having net charges x and y respectively, follow the following procedure.

(I) Write the symbols of the ions side by side in such a way that positive ion is at the left and negative ion at the right as AB.

(II) write their charges on the top of each symbol as A^xB^y.

(III) Now apply crisis- cross rule as

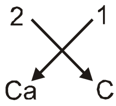


, i.e. formula A_yB_x.

(IV) Cancel out any common factor (or HCF).

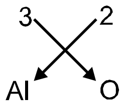
Ex. :

1. Calcium chloride



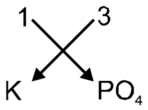
$= CaCl_2$

2. Aluminium oxide



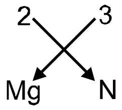
$= Al_2O_3$

3. Potassium phosphate



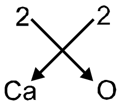
$= K_3PO_4$

4. Magnesium nitride



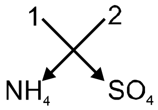
$= Mg_3N_2$

5. Calcium oxide



$= Ca_2O_2$

6. Ammonium sulphate



$= (NH_4)_2SO_4$

Cancelling the common factor, answer is CaO

	Atoms		Ions
1.	Atoms are perfectly neutral	1.	Ions are charged particles containing one or more atoms.
2.	In atoms, the number of portons is equal to the number of electrons. Na (protons 11, electrons 11) Cl (protons 17, electrons 17)	2.	In cations (positively charged ions), number of protons is more than the number of electrons. In anions (negatively charged ions) the no. of protons is less than the number of electrons. e.g. Na^+ (protons 11, electrons 10). Cl^- (protons 17, electrons 18)
3.	Except noble gases, atoms have less than 8 electrons in the outermost orbit e.g. Na 2, 8, 1; Ca 2, 8,8, 2 Cl 2, 8, 7 ; S 2, 8, 6	3.	Ions have generally 8 electrons in the outermost orbit , i.e., ns^3np^6 configuration. Na^+ , 2, 8; Cl^- 2, 8, 8 ; Ca^{2+} 2, 8, 8
4.	Chemical activity is due to loss or gain or sharing of electrons as to acquire noble gas configuration	4.	The chemical activity is due to the charge on the ion. Oppositely charged ions are held together by electrostatic forces.

Extended or Long Form of the Periodic

☐ Metals

Nonmetals

■ Metalloids

s-Block Elements · d-Block Elements · p-Block Elements

Group → Period ↓	IA (1)	IIA (2)	IIIB (3)	VIB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	VIII (9)	VIII (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
1	1 **H** 1.0079 Hydrogen																	2 **He** 4.0026 Helium
2	3 **Li** 6.940 Lithium	4 **Be** 9.0122 Beryllium											5 **B** 10.811 Boron	6 **C** 12.011 Carbon	7 **N** 14.007 Nitrogen	8 **O** 15.999 Oxygen	9 **F** 18.998 Fluonne	10 **Ne** 20.180 Neon
3	11 **Na** 22.990 Sodium	12 **Mg** 24.305 Magne-											13 **Al** 26.982 Alu-	14 **Si** 28.086 Silicon	15 **P** 30.974 Phosporus	16 **S** 32.066 Sulphur	17 **Cl** 35.453 Chlorine	18 **Ar** 39.948 Argon
4	19 **K** 39.098 Potas-	20 **Ca** 40.078 Calcium	21 **Sc** 44.956 Scandium	22 **Ti** 47.867 Titanium	23 **V** 50.941 Vanadium	24 **Cr** 51.996 Chrominum	25 **Mn** 54.938 Manga-	26 **Fe** 55.847 Iron	27 **Co** 58.933 Cobolt	28 **Ni** 58.693 Nickel	29 **Cu** 63.546 Copper	30 **Zn** 65.39 Zinc	31 **Ga** 62.723 Gallium	32 **Ge** 72.61 Germa-	33 **As** 74.922 Arsenic	34 **Se** 78.96 Selenium	35 **Br** 79.904 Bromine	36 **Kr** 83.80 Krypton
5	37 **Rb** 85.468 Rubidium	38 **Sr** 87.62 Strontium	39 **Y** 88.906 Yttrium	40 **Zr** 91.224 Zirconium	41 **Nb** 92.906 Niobium	42 **Mo** 95.94 Molybdenum	43 **Tc** 98 Techne-	44 **Ru** 101.07 Ruthe-	45 **Rh** 102.91 Rhodium	46 **Pd** 106.42 Palladium	47 **Ag** 107.87 Silver	48 **Cd** 112.41 Cadmium	49 **In** 114.82 Indium	50 **Sn** 118.71 Tin	51 **Sb** 121.76 Antimony	52 **Te** 127.60 Tellurium	53 **I** 126.90 Iodine	54 **Xe** 131.29 Xenon
6	55 **Cs** 132.91 Cesium	56 **Ba** 137.33 Barium	57 **La*** 138.91 Lantha-	72 **Hf** 178.49 Hafnium	73 **Ta** 180.95 Tantalum	74 **W** 183.84 Tungsten	75 **Re** 186.21 Rhenium	76 **Os** 190.23 Osmium	77 **Ir** 192.22 Iridium	78 **Pt** 195.08 Platinum	79 **Au** 196.97 Gold	80 **Hg** 200.59 Mercury	81 **Tl** 204.38 Thallium	82 **Pb** 207.2 Lead	83 **Bi** 208.98 Bismuth	84 **Po** 210 Polonium	85 **At** 210 Astatine	86 **Rn** 222 Radon
7	87 **Fr** 223 Francium	88 **Ra** 226 Radium	89 **Ac**** 227 Actinium	104 **Unq** 261 Unnilqua	105 **Unp** 262 Unnilqua	106 **Unh** 266 Unnilqua	107 **Uns** 264 Unnilqua	108 **Uno** 269 Unnilqua	109 **Une** 268 Unnilqua	110 **Uun** 269 Ununnilium	111 **Uuu** 272 Unnilqua	112 **Uub** 277 Ununbium						

The symbols for elements 104-109 used in this table are those proposed by the American Chemical Society and 110-112 proposed by IUPAC

f-Block Elements

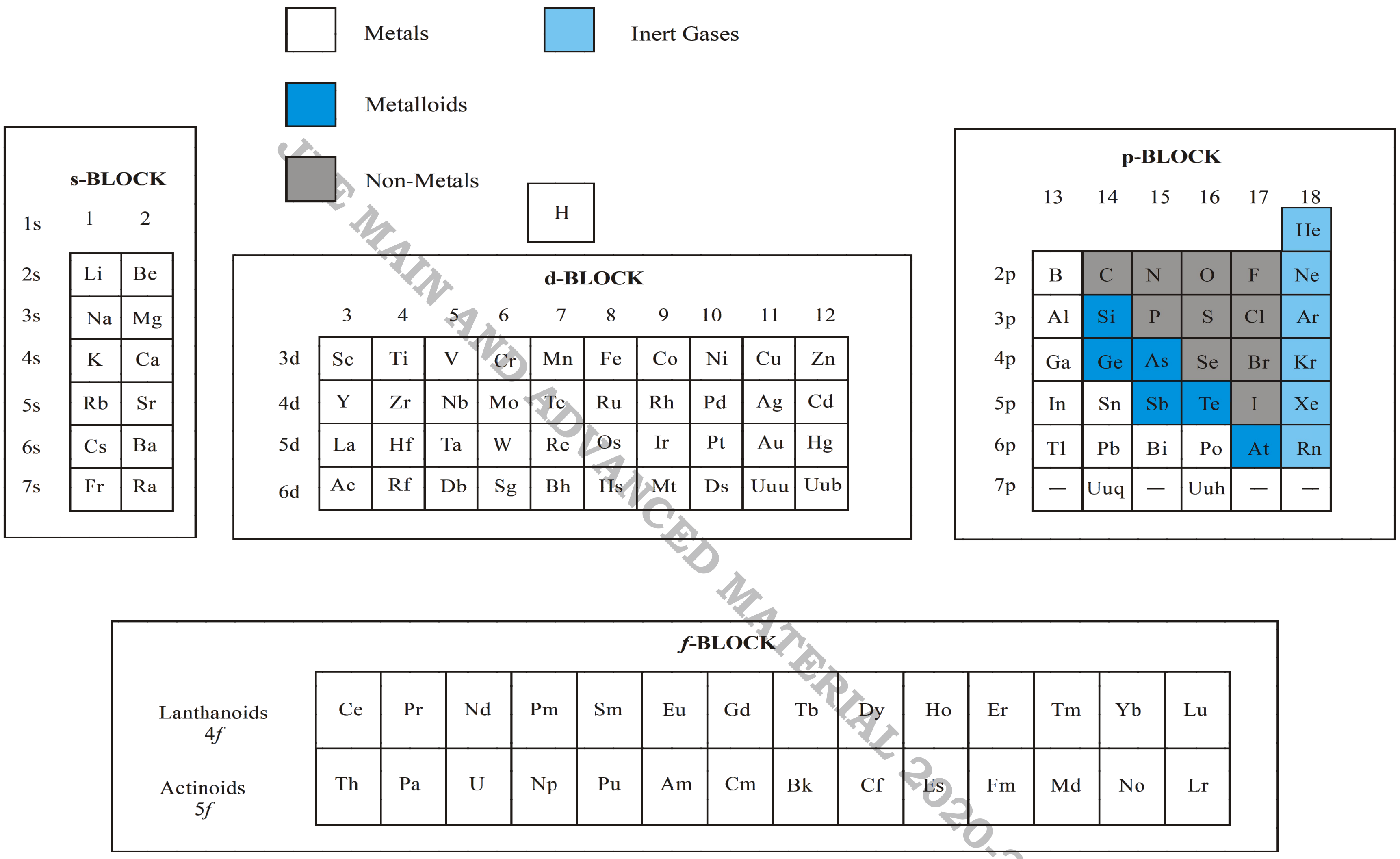
6	***Lanthanide Series**	58 **Ce** 140.12 Cerium	59 **Pr** 140.91 Praseodymium	60 **Nd** 144.24 Neodymium	61 **Pm** 145 Promethium	62 **Sm** 150.36 Samarium	63 **Eu** 151.96 Europium	64 **Gd** 157.25 Gado-	65 **Tb** 158.93 Terbium	66 **Dy** 162.50 Dyspro-	67 **Ho** 164.93 Holmium	68 **Er** 167.26 Erbium	69 **Tm** 168.93 Thulium	70 **Yb** 173.07 Ytterbium	71 **Lu** 174.97 Lutetium
7	****Actinide Series**	90 **Th** 232.04 Thorium	91 **Pa** 231.04 Protactinium	92 **U** 238.03 Uranium	93 **Np** 237 Nep-	94 **Pu** 244 Plutonium	95 **Am** 243 Ameri-	96 **Cm** 247 Curium	97 **Bk** 247 Berkelium	98 **Cf** 251 Californium	99 **Es** 252 Einsteinium	100 **Fm** 257 Fermium	101 **Md** 258 Men-	102 **No** 259 Nobelium	103 **Lr** 262 Lawrencium

IUPAC designations of groups of elements are given in

GENERAL TREND OF DIFFERENT PROPERTIES IN THE PERIOD AND GROUPS

	PERIODS ⟶	GROUPS ⟶
Atomic radius	Decreases	Increases
Ionisation potential	Increases	Decreases
Electronegativity	Increases	Decreases
Electron affinity	Increases	Decreases
Covalent character of halides	Increases	Decreases
Metallic character	Decreases	Increases
Oxidising nature	Increases	Decreases
Reducing nature	Decreases	Increases
Screening effect	Increases	Increases
Effective nuclear charge (Z_{eff})	Increases	Constant
Valency w.r.t Oxygen	Increases	Constant
Basic character of hydroxides	Decreases	(Metals)Increases
Basic character of hydrides	Decreases	(Non metal) Decreases (Metals) Increases
Basic character of Oxides	Decreases	Increases
Strength of oxy acids	Increases	Decreases
Thermal stability of sulphate	Decreases	Increases
Thermal stability of carbonates	Decreases	Increases
Thermal stability of nitrates	Decreases	Increases
Therma stability of hydroxide	Decreases	Increases
Density	First increases then decreases	Increases
Electro positivity	Decreases	Increases

APPENDIX
B



s-BLOCK

	1	2
1s		
2s	Li	Be
3s	Na	Mg
4s	K	Ca
5s	Rb	Sr
6s	Cs	Ba
7s	Fr	Ra

d-BLOCK

	3	4	5	6	7	8	9	10	11	12
3d	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
4d	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
5d	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg
6d	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Uuu	Uub

p-BLOCK

	13	14	15	16	17	18
						He
2p	B	C	N	O	F	Ne
3p	Al	Si	P	S	Cl	Ar
4p	Ga	Ge	As	Se	Br	Kr
5p	In	Sn	Sb	Te	I	Xe
6p	Tl	Pb	Bi	Po	At	Rn
7p	—	Uuq	—	Uuh	—	—

***f*-BLOCK**

Lanthanoids *4f*	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Actinoids *5f*	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

The Types of elements in the Periodic Table based on the orbitals that are being filled.

IMPORTANT FACTS TO REMEMBER

1. Lowest electronegativity : Cs
2. Highest electronegativity : F
3. Highest ionisation potential : He
4. Lowest ionisation potential : Cs
5. Lowest electron affinity : Noble gases
6. Highest electron affinity : Chlorine
7. Least electropositive element : F
8. Lowest m. pt. metal : Hg
9. Highest m. pt. and b. pt. metal : W (Tungsten)
10. Lowest m. pt. and b. pt. non metal : He
11. Notorious element : Hydrogen
12. Lightest element : Hydrogen
13. Smallest atomic size : H
14. Largest atomic size : Cs
15. Largest anionic size : I^-
16. Smallest cation : H^+
17. Most electropositive element : Cs
18. Element with electronegativity next to Fluorine : Oxygen
19. Group containing maximum no. of gaseous elements in periodic table : Zero group
20. Total number of gaseous elements : in periodic table : 11 (H, N, O, F, Cl, He, Ne, Ar, Kr, Xe, Rn)
21. Total number of liquid elements in periodic table : 6 (Ga, Br, Cs, Hg, Fr, Uub)
22. Liquid element of radioactive nature : Fr
23. Total number of radioactive elements in periodic table : 25

24. Volatile d block elements : Zn, Cd, Hg, Uub

25. Element containing no neutron : H

26. Most abundant element on earth : Oxygen

27. Most abundant metal on earth : Al

28. Element having maximum tendency for catenation : Carbon

29. Non metal having highest m. pt., b.pt. : Carbon (dimond)

30. Metals showing highest oxidation number : Os (+8), Ru

31. Most electrovalent compound : CsF

32. Most stable carbonate : Cs_2CO_3

33. Strongest alkali : CsOH

34. Strongest basic oxide : Cs_2O

35. Best electricity conductor among metals : Ag

36. Best electricity conductor among non metals : graphite

37. Most poisonous element : Pu (Plutonium)

38. Liquid non metals : Br

39. Element kept in water : Phosphorous

40. Elements kept in kerosene : IA group element (except Li)

41. Elements sublime on heating : I_2

42. Noble metals : Au, Pt etc.

43. Amphoteric metal : Be, Zn, Al, Sn, Pb, Ga

44. Metalloids elements : Si, As, Te, At, Ge, Sb

45. Non metals having metallic lusture : Graphite, Iodine

46. Heaviest naturally occurring element : Uranium

47. Poorest conductor of electricity : Dimond

48. Hardest naturally ocurring element : Dimond

49. Lightest solid metal : Li

50. 90% of Sun mass : Hydrogen

51. Amphoteric oxides : BeO, Al_2O_3, ZnO, PbO, PbO_2,
SnO, SnO_2,
Sb_2O_3, As_2O_3, Cr_2O_3 etc.

52. Neutral oxides of non metals : NO, CO, H_2O, N_2O

53. Dry bleacher : H_2O_2

54. Dry ice : Solid CO_2

55. Artificial explosive : TNT, RDX (Research Developed Explosive etc.)

56. First noble prize of chemistry was given to : Vantt Haff

57. Some isomorphous substances : $FeSO_4.7H_2O$, $MgSO_4.7H_2O$, $ZnSO_4.7H_2O$

58. Some effloroscent substances : $Na_2CO_3.10H_2O$, $MgSO_4.7H_2O$ etc.

59. First man made element : Tc_{43} (Technicium)

60. Smallest period : Ist (2 elements)

61. Largest period in periodic table : 6th (32 element)

62. Largest group in periodic table : IIIB (32 element)

63. Most abundant d-block metal : Fe

64. Most abundant s-block metal : Ca

65. Most stable element : Te

66. Highest density (Metals) : Os, Ir

67. Highest density (Non Metals) : Boron

IMPORTANT SCIENTISTS & THEIR CONTRIBUTIONS

1.	Dobereiner	:	Law of Triads
2.	Fajans	:	Factors of polarisation.
3.	Lavoisier	:	Father of chemistry, classified element into metals & non metals.
4.	Glesspie - Nyholm	:	VSEPR Theory
5.	Mendeleef	:	Periodic classification of elements.
6.	Mosley	:	Concept of atomic number, modern periodic law, modern periodic table
7.	Newland	:	Law of octaves.
8.	Slater	:	Screening constant, Geometry of molecules.
9.	Schoomaker & Stevenson	:	Bond length depends on difference in electronegativity
10.	Lewis and Kossel	:	Electronic theory - octet rule
11.	Rang & Warner	:	Long form of modern periodic table
12.	Ramsay	:	Introduced noble gases in periodic table, Argon was discovered.
13.	Seaborg	:	Post-Uranic elements
14.	Heitler and London	:	Valence bond theory.
15.	Hund's rule	:	Distribution of electrons in atomic orbitals in a sub - shell.
16.	Pauling	:	Concept of dipole moment, valence bond theory Hybridization, Electronegativity scale.
17.	Sidgwick and powell	:	Hybridisation.
18.	Hanny - Smith	:	Nature of bonds.