

Periodic Table

1. PROUT'S HYPOTHESIS

The atomic weights of all elements are simple multiple of atomic weight of hydrogen. Prout gave this hypothesis on the basis of Dalton's atomic theory and the atomic weights of some elements known at that time. But this hypothesis could not last longer, because there are some atomic weights which are fractional and not in whole number.

2. DOBEREINER'S TRIADS

According to Dobereiner when elements of same properties are kept in the increasing order of atomic weights, the atomic weight of middle element is equal to the mean atomic weight of remaining two elements. Such a group of elements is called **Dobereiner's triad**.

Triad of atoms			Mean of first and last element
Li	Na	K	$\frac{7 + 39}{2} = 23$
7	23	39	
Be	Mg	Ca	$\frac{8 + 40}{2} = 24$
8	24	40	

Dobereiner could arrange only a few elements as triads and there are some such elements present in a triad, whose atomic weights are approximately equal, e.g.

Fe	Co	Ni
Ru	Rh	Pd

Therefore, this hypothesis was not acceptable for all elements.

3. NEWLAND'S RULE OF OCTAVE

- As in music, the eighth note is same as the first note. If the elements are arranged in the increasing order of atomic weights, on starting with an element, the first element will exhibit similarities with the eighth element e.g.

	do	re	mi	pha	sol	la	si
	1	2	3	4	5	6	7
Symbole of element	Li	Be	B	C	N	O	F
	7	9	11	12	14	16	19
Symbole of element	Na	Mg	Al	Si	P	S	Cl
	23	24	27	28	31	32	35.5

- It is clear from the above table that sodium is the eighth element from lithium, whose properties resemble that of lithium.
- This type of classification was limited up to only 20 elements.

4. LOTHAR MEYER'S VOLUME CURVES

- The graphs of atomic volumes against atomic weights are known as Lothar Meyer volume curves.
- The alkali metals have highest atomic volumes.
- Alkaline earth metals (Be, Mg, Ca, Sr, Ba, etc.) which are relatively a little less electropositive. Occupy positions on the descending part of the curve.
- Halogens and the noble gases (except helium) occupy positions on the ascending part of the curve.
- Transition elements have very small volumes and therefore these are present at the bottoms of the curve

5. MENDELEEF'S PERIODIC LAW

- According to Mendeleef's periodic law, the physical and chemical properties of elements are periodic function of their atomic weights.

5.1 MENDELEEF'S PERIODIC TABLE

- Periodic table is based on atomic weight.
- In the periodic table, the horizontal lines are called periods and the vertical lines are called groups.
- The periodic table consists of seven periods and nine groups (The earlier periodic table had only 8 groups. The noble gases were added later in the zero group because these were not discovered when Mendeleef put forward his periodic table.
- All the groups (except VIII and Zero groups) are divided into subgroups A and B.
- 2, 8, 18 and 32 are called magic numbers.

5.2 MERITS OF MENDELEEF'S PERIODIC TABLE

- Classification of elements then known, was done for the first time and the elements having similar properties were kept in the same group.
- It encouraged research and led to discovery of newer elements.
- Mendeleef had even predicted the properties of many elements not discovered at that time. This helped in the discovery of these elements. For example. Mendeleef predicted the properties of the following elements.
 - (a) Eka-boron- This was later called scandium (Sc)
 - (b) Eka-aluminium - This was later called gallium (Ga)
 - (c) Eka-silicon - This was later called germanium (Ge)

5.3 DEFECTS OF MENDELEEF'S PERIODIC TABLE

- (1) **Position of Hydrogen** – Hydrogen resembles alkali metals and halogens in its properties. Hence its position was not shoure.
- (2) **Position of Isotopes** – The isotopes have different atomic weights and the periodic table is based on atomic weights. Therefore, isotopes should get different places in the periodic table on the basis of atomic weights.
- (3) The periodic table is not fully based on increasing order of atomic weights.
- (4) It is not proper to place together the elements having differing properties, such as coinage metals (Cu, Ag and Au) with alkali metals; Zn, Cd and Hg with alkaline earth metals and metal like Mn with halogens. Similarly. Pt and Au having similar properties have been placed in different groups.
- (5) There is no indication whether lanthanides and actinides are associated with group IIIA or group IIIB.
- (6) Position of Isobars – These elements have different groups when mass remains same.
- (7) Lot of stress was given to valency of elements.

6. MODERN PERIODIC LAW AND MODERN PERIODIC TABLE

- **Mosley** proved that the square root of frequency (ν) of the rays, which are obtained from a metal on showering high velocity electrons is proportional to the nuclear charge of the atom. This can be represented by the following expression.

$$\sqrt{\nu} = a(Z-b) \text{ where } Z \text{ is nuclear charge on the atom and } a \text{ and } b \text{ are constants.}$$

- The nuclear charge on an atom is equal to the atomic number.
- According to modern periodic law. "The properties of elements are the periodic functions of their atomic numbers"

6.1 MODERN PERIODIC TABLE

- On the basis of the modern periodic law, a scientist named **Bohr** proposed a long form of periodic table that was prepared by Rang and Warner.
- In the periodic table the horizontal lines are periods and the vertical lines are groups.
- The periodic table has a total of seven periods and 18 groups. But according to CAS system, the number of groups is 16, because the eighth group has been divided into three groups.
- There are two elements in the first period eight elements in each of the second and third periods, eighteen elements in each of the fourth and fifth period thirty two elements in the sixth period and only nineteen elements till now in the seventh period. Total 105 elements have been discovered so far.
- The first period is very short period, second and third are short periods fourth and fifth are long periods sixth is very long period, while the seventh is incomplete period.
- The lanthanides (Elements from atomic numbers 58 to 71) and actinides (elements from atomic numbers 90 to 103) are included in the sixth and seventh periods through these have been kept outside the periodic table.
- Period - The details about the seven periods are as follows.

Period	Atomic number		Number of elements
	From	to	
First	H (1)	He (2)	2
Second	Li (3)	Ne (10)	8
Third	Na (11)	Ar (18)	8
Fourth	K (19)	Kr (36)	18
Fifth	Rb (37)	Xe (54)	18
Sixth	Cs (55)	Rn (86)	32 (including lanthanides)
Seventh	Fr (87)	Ha (105)	19 (including actinides)

- Group - The modern periodic table has 18 vertical columns and according to CAS system there are 16 groups having the following number of elements.

Group	Number of Elements
(a) IA group	7 (H, Li, Na, K, Rb, Cs, Fr) Alkali metals
(b) IIA group	6 (Be, Mg, Ca, Sr, Ba, Ra) Alkaline earth metals
(c) IIIA group	5 (B, Al, Ga, In, Tl) Boron family
(d) IV A group	5 (C, Si, Ge, Sn, Pb) Carbon family
(e) V A group	5 (N, P, As, Sb, Bi) Nitrogen family
(f) VI A group	5 (O, S, Se, Te, Po) Oxygen family (chalcogen)
(g) VII A group	5 (F, Cl, Br, I, At) - Halogen family
(h) Zero group	6 (He, Ne, Ar, Kr, Xe, Rn) Inert elements
(i) III B group	32 (Sc, Y, La, Ac and 14 lanthanide elements and 14 actinide elements. These are elements of IIIB group, which could not be accommodated in one column and therefore written separately outside the periodic table.
(j) IV B group	4 (Ti, Zr, Hf, Ru)
(k) V B group	4 (V, Nb, Ta, Hf)
(l) VI B group	3 (Cr, Mo, W)
(m) VII B group	3 (Mn, Tc, Re)
(n) VIII (3) group	9 (Fe, Co, Ni, Ru, Rh, Pd, Os, Ir, Pt)
(o) I B group	3 (Cu, Ag, Au)
(p) II B group	3 (Zn, Cd, Hg)

6.2 MERITS OF LONG FORM OF PERIODIC TABLE OVER MANDELEEF'S PERIODIC TABLE

- Positions of Isotopes and Isobars** - Isotopes have same atomic number and the periodic table is based on atomic numbers. Therefore, various isotopes of the same elements have to be provided the same position in the periodic table. Isobars have same atomic weights but different atomic numbers and therefore they have to be placed at different positions.
- The positions of actinides and lanthanides is more clear now because these have been placed in IIIB groups and due to paucity of space, these are written at the bottom of the periodic table.
- The general electronic configurations of the elements remains same in group.

6.3 DEFECTS OF LONG FORM OF PERIODIC TABLE

- The position of hydrogen is still disputable as it was there in Mendeleev periodic table in group IA as well as IVA & VIIA.
- Helium is an inert gas but its configuration is different from that of the other inert gas elements
- Lanthanide and actinide series could not be adjusted in the main periodic table and therefore they had to be provided with a place separately below the table.

7. s-BLOCK ELEMENTS

- The elements of the periodic table in which the last electron enters in s -orbital, are called s -block elements.
- s -Orbital can accommodate a maximum of two electrons.
- Their general formulae are ns^1 and ns^2 respectively, where $n = (1 \text{ to } 7)$.
- I A 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.
- The total number of s block elements are 14.
- Fr^{87} and Ra^{88} are radioactive elements while H and He are gaseous elements
- Cs and Fr are liquid elements belonging to s -block.

8. p - BLOCK ELEMENTS

- The elements of the periodic table in which the last electron gets filled up in the p -orbital, called p -block elements.
- A 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 groups.
- The general formulae of p block elements are $ns^2 p^{1-6}$ (where $n = 2 \text{ to } 6$)
- The zero group elements having general formula $ns^2 p^6$ are inert, because their energy levels are fully filled.
- The total number of p block elements in the periodic table is 30 (excluding He)
- 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.
- The step-like thick lines drawn in the periodic table in the p -block divides elements into metals nonmetals and metalloids.

9. d-BLOCK ELEMENTS

- The elements of the periodic table in which the last electron gets filled up in the d orbital, called d block elements.
- The d block elements are placed in groups named IIIB, IV B, V B, VI B, VII B, VIII, I B and II B.
- In d block elements the electron gets filled up in the d orbital of the penultimate shell. That is why these elements are known as transition elements.
- Through the total number of d block elements is 33 in the periodic table but there are only 30 transition elements. Because only those elements are transition in which d orbital is partially filled.
- The general formula of these elements is $(n-1)s^2, p^6, d^{1-10} ns^{1-2}$ where $n = 4 \text{ to } 7$.
- All of these elements are metals
- Out of all the d block elements mercury is the only liquid element.

10. f-BLOCK ELEMENTS

- The elements of the periodic table in which the last electron gets filled up in the f orbital, called f block elements.
- The f block elements are from atomic number 58 to 71 and from 90 to 103.
- The lanthanides occur in nature in low abundance and therefore, these are called rare earth elements.
- There are 28 f block elements in the periodic table.
- 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).
- All the actinide elements are radioactive
- All the elements after atomic number 92 (i.e. U^{92}) are transuranic elements.
- The general formula of these elements is $(n-2)s^2 p^6 d^{10} f^{(1-14)} (n-1)s^2 p^6 d^{0-1} ns^2$ where $n = 6 \& 7$.

Ex.1 Which alfabate not use in nomenclature of elements.

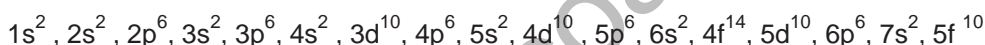
(1) K (2) W (3) V (4) J & Q **Ans. (4)**

Sol. K is potassium (Kalium), W is Tungsten (Wolfram) and V is vanadium. There is no element associated with the letters J and Q.

Ex.2 Which of the following is the period number of the element whose atomic number is 98

(1) 4 (2) 7 (3) 5 (4) 6 **Ans. (2)**

Sol. The electronic configuration of the element with atomic number 98 is as follow



The last electron enters in f orbital, so it belongs to f block in the period.

Ex.3 The nuclei of elements X, Y and Z have same number of protons, but different numbers of neutrons. According to Mendeleef periodic table, the elements X, Y and Z

- (1) belong to same group and same period
- (2) belong to different groups and different periods
- (3) belong to same group and different periods
- (4) are isotopes, which do not have different positions **Ans. (4)**

Sol. Isotopes have same number of protons (i.e. same atomic number). So they occupy same position in the periodic table. However, due to different numbers of neutrons their atomic weights are different.

Ex.4 Which of the following is the artificial element in the periodic table

(1) Tc (2) Te (3) Ru (4) Os **Ans. (1)**

Sol. Tc^{43} is the first artificial element.

Ex.5 Which of the following is not a transition element

(1) Co (2) Ni (3) Mn (4) Zn **Ans. (4)**

Sol. There is only one incomplete orbit in Zn^{+2} and its stable oxidation state is (+2) does not have incomplete d orbital. Therefore it is not a transition element.

11. EFFECTIVE NUCLEAR CHARGE

- In a polyelectronic atom, the internal electrons repel the electrons of the outermost orbit. This results in decrease the nuclear attraction on the electrons of the outermost orbit.
- Therefore, only a part of the nuclear charge is effective on the electrons of the outermost orbit. Thus, the inner electrons protect or shield the nucleus and thereby decrease the effect of nuclear charge towards the electrons of the outermost orbit.
- Thus the part of the nuclear charge works against outer electrons, is known as **effective nuclear charge**

$$Z^* = Z - \sigma$$

Z^* = effective nuclear charge, σ = shielding constant and Z = nuclear charge

A scientist named **Slater**, determined the value of shielding constant and put forward some rules as following.

- (1) The shielding effect or screening effect of each electron of 1s orbital is 0.30.
- (2) The shielding effect of each electrons of ns and np i.e. electron of the outermost orbit, is 0.35.
- (3) The shielding effect of each electron of s,p or d orbitals of the penultimate orbit ($n - 1$) is 0.85.
- (4) The shielding effect of each electron of s, p, d or f orbital of the prepenultimate orbit ($n - 2$) and below this is 1.0.

12. ATOMIC RADIUS

- The distance of the outermost orbit from the centre of the nucleus of an atom is called atomic radius.

12.1 Single Bond Covalent Radius, SBCR - (a) For Homolatomic molecules

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

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

- (b) For hetrodiatomic molecules while electron negativity remains approx imately same.

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

- For heteonuclear diatomic molecule. A–B, while difference between the electronegativity values of atom A and atom B is relatively larger,

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

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

Note : Covalent radius is slightly smaller than actual radius.

Ex. 6 Calculate the bond length of C–X bond, if C–C bond length is 1.54 Å, X–X bond length is 1.00 Å and electronegativity values of C and X are 2.0 and 3.0 respectively

Sol. (1) C–C bond length = 1.54 Å

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

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

(2) C–X bond length

$$d_{\text{C-X}} = r_{\text{C}} + r_{\text{X}} - 0.09 (X_{\text{X}} - X_{\text{C}})$$

$$= 0.77 + 0.50 - 0.09 (3 - 2)$$

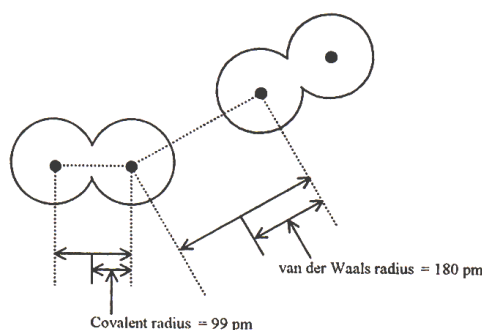
$$= 0.77 + 0.50 - 0.09 \times 1$$

$$= 1.27 - 0.09 = 1.18 \text{ \AA}$$

Thus C–X bond length is 1.18 Å

12.2 VAN DER WAALS RADIUS

- Half of the distance between the nuclei of two non bonded atoms belonging to two different molecules closest to one another is called van der waals radius.
- The values of atomic radii in noble gases are always determined as van der Waals radii. Therefore, the value of van der waals radius of a noble gas is always greater than that of the halogen coming before it in the same period.



van der Waals radius > Metallic radius > Covalent radius

Note : vander Waals radius is slightly larger than the actual radius.

12.3 PERIODICITY IN ATOMIC RADIUS

- The atomic radius depends upon the following two factors.

(a) Effective Nuclear Charge - The effect of increase in the number of protons increases the effective nuclear charge. This results in decrease in the value of atomic radius because protons attract the electronic orbits with greater force.

(b) Number of orbits - The effect of increase in the number of orbit in an atom increases the atomic size.

(c) Shielding effect/screening effect - The electrons of inner shell repel the electrons of valence shell from coming closer to the nucleus. Due to this the atomic radius increasing. This is called as shielding. When an atom has more number of shells, the shielding provided will also be greater.

With in a shell the shielding provided by different types of orbitals follows the order $s > p > d > f$. (As 's' orbitals experience greater penetration towards the nucleus, they provide greater shielding to the electrons of p, d and f orbitals).

In the elements of 'd' and 'f' blocks where number of shells remain the same in a particular series along with shielding provided, the nuclear charge is also compared.

- **In a period** - The number of orbits remains same on going from left to right in a period while there is a unit increase in the atomic number. Thus the electron experiences more force of attraction towards nucleus. Hence atomic radius decreases from left to right in a period.

- **In a group** - The atomic radii increase on going downwards in a group because the number of orbits also increase on going from top to bottom in a group.
- In a period, the size of an alkali metal (Group I A) is second largest because it has minimum number of proton, while the size of the halogen is smallest.
- In the periodic table, Cs is the biggest atom, because Fr is a radioactive element, while H is the smallest atom.

Ex.7 The increasing order of atomic size of Li, Be, B and Ne is

Sol. $B < Be < Li < Ne$

Inert gas is biggest in a period

Ex.8 The increasing order of atomic size As, Bi, Sn, Pb and Sb is

Sol. $As < Sb < Sn < Bi < Pb$

Ex.9 Which of the following should be the longest bond ?

(1) S-H

(2) O-H

(3) N-H

(4) P-H

Ans. (4)

Sol. The atomic radius of P is largest out of O, S, N, and P. Therefore, P-H bond will be the longest one .

12.4 IONIC RADIUS

- When a neutral atom loses one or more electrons it forms a **cation** having one or more number of positive charge. Similarly when a neutral atom acquires one or more electrons it forms an **anion** having one or more number of negative charge.
- **Ionic radius** is the distance between the nucleus and the limit of the electron cloud scattered around the nucleus in an ion.

12.4.1 Cationic Radius

- An atom forms a cation on loss of electron/s. The cationic radius can be defined as the distance between the nucleus and the limit of the electron cloud scattered around the nucleus.
- The size of a cation is smaller in comparison to the size of its corresponding atom. Usually a cation has on shell less than the neutral atom hence it has smaller size than the atom. This is because of the fact that an atom on losing electrons/s form a cation, which has lesser number of electrons/s than the number of proton/s. This results in increase in the effective nuclear charge.

Examples - (1) $Mn > Mn^{+2} > Mn^{+3} > Mn^{+4} > Mn^{+6} > Mn^{+7}$

(2) $Pb^{+2} > Pb^{+4}$

12.4.2 Anionic Radius

- When a neutral atom gains electron/s it becomes a negatively charged ion called an anion. The distance between the nucleus of an anion and the limit of the electron cloud scattered around the nucleus, is called its **anionic radius**.
- The size of an anion is greater than the size of its corresponding atom, because the number of electrons present in the anion gets larger than the number of protons due to gain of electron/s. This results in decrease in the effective nuclear charge.

$O^0 < O^{-1} < O^{-2}$

12.5 SIZE OF ISOELECTRONIC SERIES

- The species, which have same number of electrons but different nuclear charges, constitute an isoelectronic series. In the isoelectronic species with the increase in effective nuclear charge, the size of ion goes on decreasing.

Ex. 10 What should be the order of size of H^{-1} , H^{+1} and H ?

H^{-1}	H^{+1}	H
1p	1p	1p
2e	0e	1e

Sol. $H^{+1} < H < H^{-1}$

Ex.11 Which of the following has the largest size

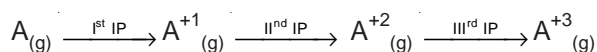
(1) N^{-3} (2) O^{-2} (3) K^{+1} (4) Ca^{+2} **Ans. (1)**

Sol.

N^{-3}	O^{-2}	K^{+1}	Ca^{+2}
7p	8p	19p	20p
10e	10e	18e	18e

13. IONISATION POTENTIAL

- The energy required to remove the most loosely bound electron from the outermost orbit of one mole of isolated gaseous atoms of an element, is called **ionisation potential (IP)**. This ionisation is an endergonic or energy-absorbing process.
- An electron cannot be removed directly from an atom in solid state. For this purpose, the solid state is converted to gaseous state and the energy required for this is called **sublimation energy**.



- The energy required to remove one electron from a neutral gaseous atom to convert it to monovalent cation, is called first ionisation potential ($I^{\text{st}} \text{ IP}$). The energy required to convert a monovalent cation to a divalent cation is called second ionisation potential ($II^{\text{nd}} \text{ IP}$).
- $I^{\text{st}} \text{ IP} < II^{\text{nd}} \text{ IP} < III^{\text{rd}} \text{ IP}$ because as the electrons go out of the atom, the effective nuclear charge increases & the ionic size goes on decreasing. Thus the forces of electrons on valence shell electrons increases and hence the order.

13.1 FACTORS AFFECTING IONISATION POTENTIAL

- Number of shells :** With the increase in number of shells the atomic radius increases i.e. the distance of outer most shell electron from the nucleus increases and hence the ionisation potential decreases.
- Effective Nuclear Charge :** Atomic size decreases with increase in effective nuclear charge because, higher the effective nuclear charge stronger will be the attraction of the nucleus towards the electron of the outermost orbit and higher will be the ionisation potential
- Shielding Effect :** The electrons of internal orbits repel the electron of the outermost orbit due to which the attraction of the nucleus towards the electron of the outermost orbit decreases and thus atomic size increases and the value of ionisation potential decreases.
- Stability of half filled and fully filled orbitals :** The atoms whose orbitals are half-filled (p^3, d^5, f^7) or fully-filled (s^2, p^6, d^{10}, f^{14}) have greater stability than the others. Therefore, they require greater energy for removing an electron. However stability of fully filled orbitals is greater than that of the half filled orbitals
- Penetration power :** In any atom the s orbital is nearer to the nucleus in comparison to p, d and f orbitals. Therefore, greater energy is required to remove an electron from s orbital than from p, d and f orbitals. Thus the decreasing order of ionisation potential of s, p, d and f orbitals is as follows

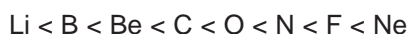
$$s > p > d > f$$

13.2 PERIODIC TRENDS IN IONISATION POTENTIAL

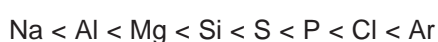
- (a) **In a Period :-** The value of Ionisation potential normally increase on going from left to right in a period, because effective nuclear charge increases and atomic size decreases.

13.2.1 Exceptions

- In second period ionisation potential of Be is greater than that of B, and in the third period ionisation potential of Mg is greater than that of Al due to high stability of fully filled orbitals.
- In second period ionisation potential of N is greater than O and in the third period ionisation potential of P is greater than that of S, due to stability of half filled orbitals.
- The increasing order of the values of ionisation potential of the second period elements is



The increasing order of the values of ionisation potential of the third period elements is



13.2.2 IONISATION POTENTIAL OF TRANSITION ELEMENTS

- In transition elements, the value of ionisation potential has very little increase on going from left to right in a period because the outermost orbit remains the same but electrons get filled up in the $(n-1)d$ orbitals resulting in very little increase in the values of ionisation potential.
- In transition element series the first ionisation potential normally increases with increase in atomic number on going from left to right, but this periodicity is not uniform. The value of ionisation potential of transition elements depends on the following two important factors.

(a) The value of ionisation potential increases with increase in effective nuclear charge.

(b) The value of ionisation potential decreases with increase in shielding effect when the number of electrons increases in $(n-1)d$ orbitals

- In the first transition element series the first ionisation potential normally increases on going from left to right from Sc to Cr because shielding effect is much weaker in comparison to effective nuclear charge. The value of first ionisation potential of Fe, Co and Ni remains constant, because shielding effect and effective nuclear charge balance one another. The value of ionisation potential shows slight increase from Cu to Zn because they have fully filled s and d orbitals. The value of first ionisation potential of Mn is maximum because it has maximum stability due to fully filled s and d orbitals.

13.2.3 INNER TRANSITION ELEMENTS

The size of inner transition elements is greater than that of d block elements. Therefore the value of ionisation potential of f block elements is smaller than that of d block elements and due to almost constant atomic size of f block elements in a period the value of their ionisation potential remains more constant than that of d block elements.

13.2.4 In a Group

- The value of ionisation potential normally decreases on going from top to bottom in a group because both atomic size and shielding effect increase.

Exception :

- The value of ionisation potential remains almost constant from Al to Ga in the III A group. ($B > Al$, $Ga > In$)
- In IV_B group i.e. Ti, Zr and Hf the I.P. of Hf is higher than that of Zr due to Lanthanide contraction. Thus the I.P. of IV_B group varies as $Ti > Zr < Hf$.

- In the periodic table the element having highest value of ionisation potential is He.
- The values of ionisation potential of noble gases are extremely high, because the orbitals of outermost orbit are fully-filled (ns^2 , np^6) and provide great stability.
- In a period, the element having least value of ionisation potential is an alkali metal (group I A) and that having highest value is inert gas (Group 0)

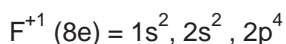
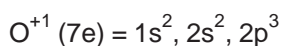
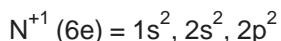
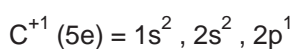
13.3 APPLICATIONS OF IONISATION POTENTIAL

- The elements having high values of ionisation potential have low reactivity, e.g. inert gases.
- The value of ionisation potential decreases more on going from top to bottom in a group in comparison to a period. Therefore, reactivity increases and the atom forms a cation by loss of electron.
- The elements having low value of ionisation potential readily lose electron and thus behave as strong reducing agents.
- The elements having low value of ionization potential readily lose electron and thus exhibit greater metallic property.
- The elements having low value of ionisation potential readily lose electron and thus oxide and hydroxides of these elements have basic property.

Ex.12 Which of the following should be the order of increasing values of second ionisation potential of C_6 , N_7 , O_8 and F_9

- (1) $C > N > F > O$ (2) $C < F < N < O$ (3) $C < F < N < O$ (4) $C < N < F < O$ **Ans. (4)**

Sol. The second ionisation potential means removal of electron from a cation

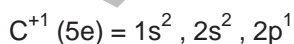
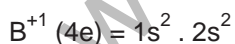
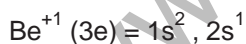


Therefore $C < N < F < O$

Ex.13 Which of the following should be the correct order of the second ionisation potential of Li, Be B and C

- (1) $Li < Be > B < C$ (2) $Be < B < C < Li$ (3) $Be < C < B < Li$ (4) $Li < C < B < Be$ **Ans. (3)**

Sol. $Li^{+1} (2e) = 1s^2$



Therefore $Be < C < B < Li$

Here, Li^{+1} has inert configuration and B^{+1} has fully filled s orbital. Hence these will have high value of second ionisation potential.

Ex.14 Which of the following should have least value of ionisation potential

- (1) Be^{+3} (2) H (3) Li^{+2} (4) He^{+1} **Ans. (2)**

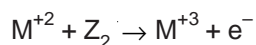
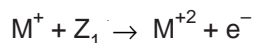
Sol. All the above four species have one electron each, H has least number of proton. Therefore, H will have least value of ionisation potential.

Ex.15 Which of the following should have less than 11.0 eV difference between the values of first and second ionisation potentials

- (1) Mg (2) Na (3) K (4) All of the above three **Ans. (1)**

Sol. Mg exhibits higher oxidation state (Mg^{+2}) than Na (Na^{+1}) and K (K^{+1})

Ex.16 Which of the following should be correct for Z_1 and Z_2 in the following two processes



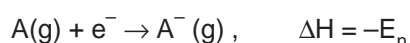
- (1) $\frac{1}{2} Z_1 = Z_2$ (2) $Z_1 = Z_2$ (3) $Z_1 = \frac{1}{2} Z_2$ (4) $Z_1 < Z_2$ **Ans. (4)**

Sol. Z_1 = second ionisation potential and Z_2 = Third ionisation potential.

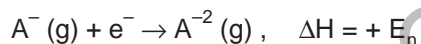
Second ionisation potential is always less than the third ionisation potential.

14. ELECTRON AFFINITY

- The energy released on adding up one mole of electron to one mole of neutral atom (A) in its gaseous state to form an anion (A^-) is called **electron affinity** of that atom. Since the electron adds up in the outermost orbit, energy is given out. Therefore, **electron affinity** is associated with an exothermic process.



- When one electron adds up to a neutral atom, it gets converted to a uninegative ion and energy is released. On adding one more electron to the mononegative anion, there is a repulsion between the negatively charged electron and anion. In order to counteract the repulsive forces, energy has to be provided to the system. Therefore, the value of the second electron affinity is positive.



14.1 FACTORS AFFECTING ELECTRON AFFINITY

- Atomic Size or Atomic Radius**

When the size or radius of an atom increases, the electron entering the outermost orbit is more weakly attracted by the nucleus and the value of electron affinity is lower.

- Effective Nuclear charge**

When effective nuclear charge is more then, the atomic size less. Then the atom can easily gain an electron and higher electron affinity.

- Stability of Fully-Filled and Half-Filled Orbitals**

The stability of the configuration having fully-filled orbitals (p^6, d^{10}, f^{14}) and half-filled orbital (p^3, d^5, f^7) is relatively higher than that of other configurations. Hence such type of atoms have less tendency to gain an electron, therefore their electron affinity values will be very low or zero.

14.2 TRENDS IN ELECTRON AFFINITY

14.2.1 In a period, atomic size decreases with increase in effective nuclear charge and hence increase in electron affinity.

Exception :

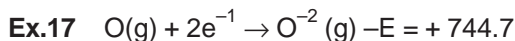
- On going from C^6 to N^7 in the second period, the values of electron affinity decreases instead of increasing. This is because there are half-filled ($2p^3$) orbitals in the outermost orbit of N, which are more stable. On the other hand, the outermost orbit in C has $2p^2$ configuration.
- In the third period, the value of electron affinity of Si is greater than that of P. This is because electronic configuration of the outermost orbit in P atom is $3p^3$, which being half-filled, is relatively more stable
- The values of electron affinity of inert gases are zero, because their outermost orbit has fully-filled p orbitals.
- In a period, the value of electron affinity goes on decreasing on going from group IA to group IIA. The value of electron affinity of the elements of group IIA is zero because ns orbitals are fully-filled and such orbitals have no tendency to accept electrons.

14.2.2 In a Group

- The values of electron affinity normally decrease on going from top to bottom in a group because the atomic size increases which decreases the actual force of attraction by the nucleus.

Exceptions

- The value of electron affinity of F is lower than that of Cl, because the size of F is very small and compact and the charge density is high on the surface. Therefore, the incoming electron experiences more repulsion in comparison to Cl. That is why the value of electron affinity of Cl is highest in the periodic table.
- The values of electron affinity of alkali metals and alkaline earth metals can be regarded as zero, because they do not have tendency to form anions by accepting electron.



The reason for the positive value of E is

- (1) endothermic reaction (2) exothermic reaction
 (3) both 1 and 2 (4) All of the above are wrong

Ans. (1)

Sol. When electron is brought near O^{-1} there will be repulsion between them, and therefore the energy will be positive i.e there will be absorption of energy during the process.

Ex.18 The increasing order of electron affinity of N, P and As is

- (1) $N < P < As$ (2) $As < P < N$ (3) $As < N < P$ (4) $As < N > P$

Ans. (3)

Sol. Phosphorus have vacant 'd' orbitals due to which it has higher electron affinity than Nitrogen.

Ex.19 The increasing order of electron affinity values of O, S and Se is

- (1) $O < S < Se$ (2) $S < O < Se$ (3) $Se < O < S$ (4) $Se < O > S$

Ans. (3)

Sol. Atomic size of Se is very large.

15. ELECTRONEGATIVITY

- The measure of the capacity or tendency of an atom to attract the shared pair of electrons of the covalent bond towards itself is called **electronegativity** of that atom.
- Electronegativity is a relative value that indicates the tendency of an atom to attract shared electrons more than the other atom bonded to it. Therefore it does not have any unit. Pauling was the first scientist to put forward the concept of electronegativity. On Pauling's scale the electronegativity is expressed in Paulings.
- The numerical value of electronegativity of an atom depends on its ionisation potential and electron affinity values.

15.1 Factors Affecting Electronegativity

- **Atomic size** – Electronegativity of a bonded atom decreases with increase in size with increase in size the forces of attractions on valence shell electrons decrease and hence electron negativity decreases.
- When effective nuclear charge is high the nucleus will attract the shared electrons with greater strength and the electronegativity will be high.
- This effect increases the atomic size which decreases the electronegativity value.
- **Hybridisation state of atom** – Electronegativity increases with increases in the s character of the hybrid orbital. This is because the s orbital is nearer to the nucleus and thus suffers greater attraction resulting in increase in electronegativity.
- The number of covalent bonds present between two bonded atoms is known as bond order. With increases in the bond order, the bond distance decreases, effective nuclear charge increases and thus electronegativity increases.

The increasing order of electronegativity is as follows : $C-C < C=C < C\equiv C$

- **Oxidation number** – The electronegativity value increases with increase in oxidation number because radius decreases with increase in oxidation number.

The increasing order of electronegativity is as follows : $Fe < Fe^{+2} < Fe^{+3}$

- Electronegativity does not depend on stability of fully-filled or half-filled orbitals because it is simply the capacity of nucleus to attract bonded pair of electrons.

15.2 Trends in Electronegativity

Atomic size decreases on going from left to right in a period thus electronegativity increases. Atomic size increases on going from top to bottom in a group thus electronegativity decreases.

- **Inert Gases** – The electronegativity value of inert gases is zero, because they do not form covalent bonds
- In a period, the electronegativity value of halogen is maximum, while the electronegativity value of alkali metal is minimum.
- F has maximum electronegativity value in the periodic table, while Cs has minimum electronegativity.
- According to Pauling scale, the electronegativity value of F is 4.0, O is 3.5 N is 3.0 and Cl is 3.1.

Exceptions -

- The elements of group II B i.e. Zn, Cd and Hg show increase in electronegativity value on going from top to bottom in the group.
- The elements of group III A, i.e. Al to Ga show increase in electronegativity value on going from top to bottom in the group.
- The elements of group IV A, show no change in electronegativity value on going from top to bottom in the group from Si on words.

16. MEASUREMENT OF ELECTRONEGATIVITY

16.1 Pauling Scale

If two atoms, A and B, having different electronegativity values, get bonded to form a molecule, AB, then the bond between A and B in A – B will have both covalent and ionic properties.

$$\Delta_{A-B} = \text{Observed bond energy} - \text{Energy of 100\% covalent} \text{ or } \Delta_{A-B} = D - E_{A-B}$$

where D = Observed bond energy

$$E_{A-B} = \text{Bond energy of pure covalent bond of A - B.}$$

$$\text{The value of } E_{A-A} \text{ and } E_{B-B} \text{ is } E_{A-B} = 1/2 [E_{A-A} + E_{B-B}]$$

$$= 0.208 \sqrt{\Delta_{A-B}} = X_A - X_B \text{ where } X_A > X_B \text{ or } 0.043 \times \Delta_{A-B} = (X_A - X_B)^2$$

16.2 Mulliken Scale

Mulliken suggested that the value of electronegativity of an element as an average of the values of ionisation potential and electron affinity of the element.

$$X_M = \frac{\text{I.P.} + \text{E.A.}}{2} \text{ (in eV)}$$

where X_M = Electronegativity value as given by Mulliken

$$X_p = \frac{X_M}{2.8} = \frac{\text{I.P.} + \text{E.A.}}{5.6}$$

where X_p = Electronegativity value as given by Pauling or $X_p = 0.336 (X_M - 0.615)$

16.3 Allred-Roschov's Scale

$$X_{AR} = \frac{Z_{\text{eff}} \cdot e^2}{r^2}$$

$$X_p = 0.359 \frac{Z_{\text{eff}} \cdot e^2 + 0.744}{r^2}$$

$$Z_{\text{eff}} = Z - \sigma$$

where Z = Nuclear charge

σ = Shielding constant

$$\text{or } X_p = 0.359 X_{AR} + 0.744$$

16.4 Senderson's Scale

In Senderson scale the stability ratio of an atom itself has been regarded as its electronegativity.

$$X_s \text{ or S.R.} = \frac{\text{Average electron density of an atom}}{\text{Electron density of the isoelectronic inert gas}}$$

This is related to Pauling scale as follows

$$\sqrt{X_p} = \frac{0.2}{\text{S.R. (or } X_s)} + 0.77$$

17. APPLICATIONS OF ELECTRONEGATIVITY**(i) Nomenclature**

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

Ex.20 Write the correct formula and name of the following

(a) ICl or ClI

(b) FCl or ClF

(c) BrCl or ClBr

(d) BrI or IBr

(e) OF₂ or F₂O

(f) Cl₂O or OCl₂

Sol. Correct formula

Name

(a) I⁺ Cl⁻

Iodine chloride

(b) Cl⁺ F⁻

Chlorine fluoride

(c) Br⁺ Cl⁻

Bromine chloride

(d) IBr

Iodine bromide

(e) OF₂

Oxygen difluoride

(f) Cl₂O

Dichlorine oxide

(ii) Nature of Bond

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

(iii) Metallic and Nonmetallic Nature

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

17.1 Partial Ionic Character in Covalent bonds

Partial ionic characters are generated in covalent compounds by the difference of electronegativities. **Hanny and smith** calculated percentage of ionic character from the difference of electronegativity.

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

(Here X_A X_B)

X_A is electronegativity of element A

X_B is electronegativity of element B

$$\Delta = X_A - X_B$$

17.2 Bond Length

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

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

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

Here $X_A > X_B$

17.3 Bond Strength & Stability

Bond strength and stability of A-B increases on increases in difference of electronegativities of atoms A and B bonded A-B.

Therefore $H-F > H-Cl > H-Br > H-I$

18. NATURE OF OXIDES

If difference of the two electronegativities ($X_O - X_A$) is 2.3 or more then oxide will be basic in nature. Similarly if value of $X_O - X_A$ is lower than 2.3 then the compound will be first amphoteric then acidic in nature.

Oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₂ O ₅	SO ₃	Cl ₂ O ₇
($X_O - X_A$)	2.6	2.3	2.1	1.8	1.5	1.1	0.5
Nature	Strong	Basic	Amphoteric	weak	Acidic	Strong	Strongest
	basic			acidic		acidic	acidic

19. NATURE OF HYDROXIDES

According to Gallis if electronegativity of A in a hydroxide (AOH) is more than 1.7 then it will be acidic in nature whereas it will be basic in nature if electronegativity is less than 1.7

For example NaOH and ClOH

Electronegativity (X_A) 0.9 3.00

Nature Basic Acidic

If the value is more than $X_O - X_H$, then that hydroxide will be basic otherwise it will be acidic.

Ex.21 Increasing order of electronegativities of F, Cl, Br and I is

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

Sol. Electronegativity decreases in a group on going from top to bottom. Therefore increasing electronegativity order is $I < Br < Cl < F$.

Ex.22 Electronegativity of which of the following is high ?

- (1) $-CH_3 (sp^3)$ (2) $H_2C = CH_2 (sp^2)$ (3) $CH \equiv CH (sp)$ (4) Equal in all **Ans. (3)**

Sol. Electro negativity increases with increase in % of 's' character as in $HC \equiv CH$ 50% 's' character is present hence to electro negativity is higher.

Ex.23 Which of the following is the most polar bond

- (1) N–H (2) Cl–H (3) O–H (4) Br–H **Ans. (3)**

Sol. Difference of electronegativities of O and H is very high.

Ex.24 Which of the following formula is incorrectly written

- (1) OF_2 (2) Cl_2O (3) $BrCl$ (4) None of these **Ans. (4)**

Sol. In all the formulae less electronegative element (cation) could be indicated followed by the more electronegative element (anion)

Ex.25 CF_3NH_2 is not a base, whereas CH_3NH_2 is a base. What is the reason ?

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

Ex.26 OF_2 is called oxygen difluoride, whereas Cl_2O is called dichlorine monoxide. Why ?

Sol. Electronegativity of O in OF_2 is less than F. Therefore, there will be positive charge on oxygen and negative charge on fluorine. Whereas in Cl_2O , electronegativity of Cl is less than that of O therefore there will be positive charge on Cl and negative charge on O. Positive charge is written first followed by negative charge.

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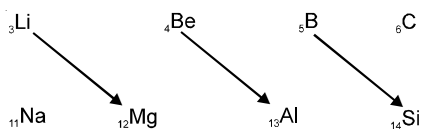
Negative N is written first in NH_3 because it became very common from the very beginning. There is no scientific basis for that. Otherwise according to rule it should have been written as H_3N in the form of trihydrogen nitride.

19.1 Acidic and Basic Nature of Hydroxides of Elements

Acidic and basic nature of hydroxide of an element AOH depends on ionisation potential of A. If ionisation potential of A is low then it will give its electron to oxygen easily thus AOH will be basic.

20. DIAGONAL RELATIONSHIP

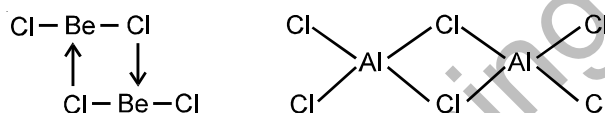
- Some elements of second period Li, Be, B shows dissimilarities with other elements of this group but shows similarities with elements of third group like Mg, Al, Si situated diagonally to them. It is called diagonal relationship.



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

- Li and Mg both reacts directly with nitrogen to form lithium nitride (Li_3N) and magnesium nitride (Mg_3N_2) whereas other alkali metals of IA group does not form nitride.
- Fluoride carbonate and phosphate of Li and Mg are insoluble in water whereas these compounds of other alkali metals are soluble.
- Li and Mg both are hard metals, whereas other metals of IA group are soft.

- (d) LiOH and $\text{Mg}(\text{OH})_2$ both are weak base, whereas hydroxides of other elements of IA group are strong base.
- (f) Metallic bond in Li and Mg both are strong compare to other alkali metals .
- (g) Their melting and boiling points are high.
- (h) By thermal disintegration of LiNO_3 and $\text{Mg}(\text{NO}_3)_2$ Li_2O and MgO is obtained respectively.
- (i) Thermal stability of Li_2CO_3 and MgCO_3 is very less compare to other alkali metals and they liberates CO_2 gas easily.
- Similarly Be shows similarity to Al of IIIA group compare to other elements of IIA group which are as follows.
 - These both elements do not provide colour to Bunsen burner.
 - They both are comparatively stable in air.
 - Both are insoluble in NH_3 therefore do not form blue coloured solution.
 - There is no tendency of making peroxide and superoxide in them.
 - Reducing power is very less due to low value of standard electrode potential in the form of oxidation potential.
 - Be and Al both forms halogen bridge halides.



21. SOME IMPORTANT POINTS

- Triad rule – Dobereiner
- Octet rule – Newland
- Study of atomic volume – Lothar Mayer
- Inventor of atomic number – Moseley
- Godfather of periodic table – Mandeleef
- Maker of modern periodic table – Bohr
- Mg is bridge element, which joins metals of IIA and II B groups.
- Elements after atomic number 92 are transuranic elements.
- Artificial element is Tc^{43} .
- Liquid non-metal – Br
- Liquid metal – Hg, Ga, Cs, Fr
- Solid volatile non-metal – Iodine
- Lightest metal – Li
- Heaviest metal – Ir
- Hardest metal – W
- Nobel metals – Pc, Pt, Au, Ag
- Element most found on earth – Al
- Gaseous elements – 11 (He , Ne , Ar , Kr , Xe , Rn , H_2 , N_2 , O_2 , Cl_2 , F_2)
- Liquid elements – 5 (Br , Hg , Ga , Cs , Fr)
- Submetals – 5 (B , Si , As , Te , At)
- Inert gases – 6

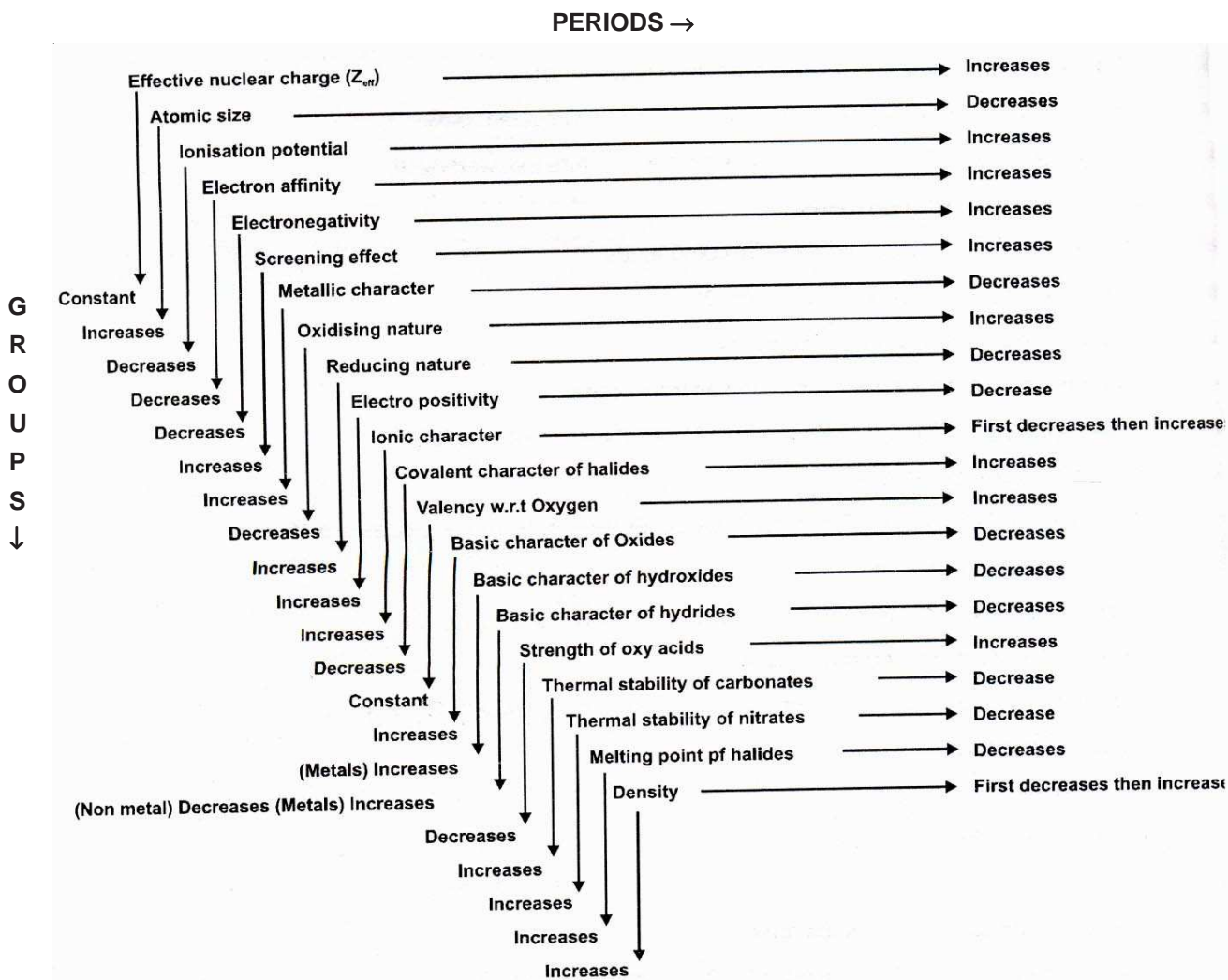
(22) Lowest electronegativity	:	Cs
(23) Highest electronegativity	:	F
(24) Highest ionisation potential	:	He
(25) Lowest ionisation potential	:	Cs
(26) Highest electron affinity	:	Chlorine (Cl)
(27) Least electropositive element	:	Fluorine (F)
(28) Most reactive solid element	:	Li
(29) Most reactive liquid element	:	Cs
(30) Most stable element	:	Te
(31) Largest atomic size	:	Cs
(32) Most electropositive element	:	Cs ; Fr (in stable element) (In all element)
(33) Group containing maximum no. of gaseous elements in periodic table	:	Zero gp ; next to zero gp is VII gp or halogen gp (F ₂ and Cl ₂)
(34) Total number of gaseous elements in periodic table	:	11 (H ₂ , He, N ₂ , O ₂ , F ₂ , Ne, Cl ₂ , Ar, Kr, Xe, Rn)
(35) Total number of liquid elements in periodic table	:	4 (Ga, Br, Cs, Hg) (Fr and Eka are also liquid)
(36) Volatile d-block elements	:	Zn, Cd, Hg
(37) Most abundant element on earth	:	Oxygen followed with Si
(38) Most stable carbonate	:	Cs ₂ CO ₃
(39) Strongest alkali	:	Cs(OH)
(40) Element kept in water	:	P
(41) Elements kept in kerosene oil	:	Na, K, I, Cs
(42) Liquid non metal	:	Br ₂
(43) Bridge metals	:	Na, Mg
(44) Noble metals	:	Au, Pt
(45) Lightest element	:	H
(46) Poorest conductor of current	:	Pb(metal), S (non metal)
(47) Most abundant gas	:	N ₂
(48) Lightest solid metal	:	Li
(49) Heaviest solid metal	:	Os(highest density 22.6 g/cm ³)
(50) Natural explosive	:	NCl ₃
(51) Dry ice	:	CO ₂
(52) First Nobel prize of chemistry was given to	:	vant Hoff

* **Core charge** - Atomic number – Kernel of electron

* **Penultimate shell** – Shell present inside one shell (n – 1) from outermost shell, is called penultimate shell.

* **Prepenultimate shell** – Shell present inside two shells (n – 2) from outermost shell, is called prepenultimate shell

GENERAL TREND OF DIFFERENT PROPERTIES IN THE PERIOD AND GROUPS



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