Periodic Table & Periodicity

Need to classify Elements :

- At present 118 elements are known. Of them the recently discovered elements are man-made.
- With such a large no. of elements it is very difficult to study individually the chemistry of all these elements and their innumerable compounds individually.
- To ease out this problem, scientists searched for a systematic way to organize their knowledge by classifying the elements.
- It would rationalize known chemical facts about elements, but even predict new ones for undertaking further study.

Section (A) : Development of Periodic Table, Period, Group and Block

Development of Modern Periodic Table :

(a) Dobereiner's Triads :

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

Li Na K 7 23 39 7+39/2=23 Fe Co Ni

55.85 58.93 58.71 nearly same atomic masses

Limitation-Since it was restricted to few elements, so it was discarded.

(b) Newland's Law of Octave :

He was the first to correlate the chemical properties of the elements with their atomic masses. According to him if the elements are arranged in the order of their increasing atomic masses the every eighth element starting from given one is similar in properties to the first one.

This arrangement of elements is called as Newland's law of Octave.

Li	Be	В	С	Ν	0	F
Na	Mg	AI	Si	Р	S	CI
К	Ca					

Limitation–This classification worked quite well for the lighter elements but it failed in case of heavier elements and therefore, discarded

(c) Lother Meyer's Classification :

He determined the atomic volumes by dividing atomic masses with their densities in solid states. He plotted a graph between atomic masses against their respective atomic volumes for a number of elements. He found the following observations.

- Elements with similar properties occupied similar positions on the curve.
- Alkali metals having larger atomic volumes occupied the crests .
- Transitions elements occupied the troughs.
- The halogens occupied the ascending portions of the curve before the inert gases.

• Alkaline earth metals occupied the positions at about the mid points of the descending portions of the curve.

On the basis of these observations he concluded that the atomic volumes (a physical property) of the elements are the periodic functions of their atomic masses. It was discarded as it lacked practical utility.

(d) Mendeleev's Periodic Table :

Mendeleev's Periodic Law: According to him the physical and chemical properties of the elements are the periodic functions of their atomic masses. He arranged the then known elements in order of their increasing atomic masses considering the facts that elements with similar properties should fall in the same vertical columns and leaving out blank spaces where necessary for the element's which he thought when discovered in future will occupy these blank spaces.

This table was divided into nine vertical columns called groups and seven horizontal rows called periods.

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
$(7)^{\text{th}} n = 7$	32	Very long period

Table-1

The groups were numbered as I, II, III, IV, V, VI, VII, VIII and Zero group

Merits of Mendeleev's Periodic table:

- It simplified and systematised the study of elements and their compounds.
- It helped in predicting the discovery of new elements on the basis of the blank spaces given in left out intentionally in the periodic table.

Mendeleev's predicted the properties of those missing elements from the known properties of the other elements in the same group. Eka - aluminium and Eka-silicon names were given for gallium and germanium (not discovered at the time of Mendeleev's). Later on it was found that properties predicted by Mendeleev's for these elements and those found experimentally were almost similar.

Table-2									
Proeprty	eka-aluminum (predicted)	gallium (found)	eka-silicon (predicted)	germanium (found)					
Atomic Mass	68	70	72	72.6					
Density / (g/cm ³)	5.9	5.94	5.5	5.36					
Melting point (K)	Low	30.2	High	1231					
Formula of oxide	E ₂ O ₃	Ga ₂ O ₃	EO ₂	GeO ₂					
Formula of chloride	ECl ₃	GaCl₃	ECI ₄	GeCl ₄					

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

Demerits in Mendeleev's Periodic Table :

- Position of hydrogen is uncertain .It has been placed in IA and VIIA groups because of its resemblance with both the groups.
- Anomalous positions of lanthanides and actinides in the periodic table.
- Order of increasing atomic weights is not strictly followed in the arrangement of elements in the periodic table. For example Ar(39.94) is placed before K(39.08) and Te (127.6) is placed before I (126.9).
- Similar elements were placed in different groups for e.g. Cu in IB and Hg in IIB and similarly the elements with different properties were placed in same groups for e.g. alkali metals in IA and coinage metals in IB.
- It could not explain the cause of periodicity.

(e) Modern Periodic Law (Moseley's Periodic Law) :

According to the Moseley's periodic low the physical and chemical properties of the elements are the periodic functions of their atomic number. If the elements are arranged in order of their increasing atomic number, after a regular interval, elements with similar properties are repeated.

Periodicity :

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

Reason 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 ns¹, therefore, have similar properties.

This table is also referred to as Bohr's table since it follows Bohr's scheme of the arrangement 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 ns¹.
- Each period ends with a noble gas with the outermost electronic configuration ns²np⁶ except helium having outermost electronic configuration as 1s².
- Each period starts with the filling of a 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 e.g.

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

$n = 3, \ \ell = 0, 1, 2 \text{ and}$	m = 0,	3,	5
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.

Solved Examples

- **Ex-1.** What was the need to organise, the periodic table?
- **Sol.** The periodic table was organised to simplify and systematize the study of the properties of all the existing elements and their compounds. On the basis of similarities in chemical properties, the various elements have now been divided into different groups. This has made the study simple because the properties of elements are now studied in form of groups rather than individually.
- Ex-2. Why do elements in the same group have similar physical and chemical properties?
- **Sol.** Elements in the same group have similar electronic configuration and hence have similar physical and chemical properties.
- Ex-3. How would you justify the presence of 18 elements in the 5th period of the Periodic Table?
- Sol. When n = 5, I = 0, 1, 2, 3. The order in which the energy of the available orbitals 4d, 5s and 5p increases is 5s < 4d < 5p. The total number of orbitals available are 9. The maximum number of electrons that can be accommodated is 18; and therefore 18 elements are there in the 5th period.
 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.

S-Block	Elements													<u>ر</u>			
1 IA															~		18 VIII A
1 H 1.007	2 II A					А						13 III A	14 IV A	15 V A	16 VI A	17 VILA	2 He 4.002
3 Li 6.941	4 Be 9.012						<u> </u>				$\overline{}$	5 B 10.811	6 C 12.011	7 N 14.006	8 O 15.999	9 F 18.998	10 Ne 20.179
11 Na 22.98	12 Mg 24.30	3 III B	4 IV B	5 V B	6 VI B	7 VII B	8 ∨III	9 VIII	10 ∨III	11 I B	12 II B	13 Al 26.981	14 Si 28.085	15 P 30.973	16 S 32.006	17 Cl 35.452	18 Ar 39.948
19 K 39.08	20 Ca 40.078	21 Sc 44.959	22 Ti 47.88	23 ∨ 50.9415	24 Cr 51.996	25 Mn 54.938	26 Fe 55.84	27 Co 55.933	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.921	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.46	38 Sr 87.62	39 Y 88.905	40 Zr 91.224	41 Nb 92.906	42 Mo 95.94	43 Tc 98	44 Ru 101.07	45 Rh 102.905	46 Pol 106.42	47 Ag 107.868	48 Cd 112.411	49 In 114.82	50 Sn 118.710	51 Sb 121.757	52 Te 127.60	53 I 126.904	54 Xe 132.29
55 Cs 132.90	56 Ba 137.27	57 La* 138.905	72 Hf 178.49	73 Ta 180.947	74 W 183.85	75 Re 186.207	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.666	80 Hg 200.59	81 ∏ 204.383	82 Pb 207.2	83 Bi 207.980	84 Po 209	85 At 210	86 Rn 222
87 Fr 223	88 Ra 226	89 Ac** 227	104 Rf 261.11	105 Ha 262.114	106 Sg 263.118	107 Bh 262.12	108 Hs 265	109 Mt 266	110 Ds 269	111 Rg 272	112 Cn 285	113 Uut 284	114 Fl 289	115 Uup 288	116 Lv 292	117 Uus	118 Uuo 294
											4						
												•					
				58 Ce 140.115	59 Pr 140.907	60 Nd 144.24	61 Pm 145	62 Sm 150.36	63 Eu 151.965	64 Gd 157.25	65 Tb 158.925	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.967
				90 Th 232.038	91 Pa 231	92 ∪ 238.028	93 Np 237	94 Pu 244	95 Am 243	96 Cm 247	97 Bk 247	98 Cf 251	99 Es 252	100 Fm 257	101 Md 258	102 No 259	103 Lr 260

Table-3

Classification of the Elements :

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

(a) s-block elements

The element in which the shells upto (n - 1) are completely filled and the last electron enters the s-orbital of the outermost (n^{th}) shell, the elements of this type are called s-block elements.

- Element of group 1 & 2 constitute the s-block.
- General electronic configuration is [inert gas] ns¹⁻²
- s-block elements lie on the extreme left of the periodic table.
- This block includes Alkali metals, alkaline earth metals.

(b) p-block elements

The element in which the shells upto (n - 1) are completely filled and the differentiating electron enters the p-orbital of the nth orbit.

- Element from group 13 to 18 constitute the p-block.
- General electronic configuration is [inert gas] ns² np¹⁻⁶
- p-block elements lie on the extreme right of the periodic table.
- This block includes some metals, all nonmetals and metalloids.
- s-block and p-block elements are collectively called normal or representative elements.

(c) d-Block elements

The element in which outermost (n^{th}) and penultimate shells $(n-1)^{th}$ shells are incompletely filled and the differentiating electron enters the (n-1) d orbitals (i.e., d-orbital of penultimate shell) are called d-block elements.

- Element from group 3 to 12 constitute the d-block.
- ♦ General electronic configuration is [inert gas] (n-1) d¹⁻¹⁰ ns¹⁻² (except, palladium which has valence shell electron configuration 4d¹⁰ 5s⁰).
- All the transition elements are metals and most of them form coloured complexes or ions.
- d-block elements are classified into four series as given below :
- (1) Ist transition series i.e. 3d series which contains 10 elements and starts from ₂₁Sc ₃₀Zn. Filling of electrons takes place in 3d sub-shell.
- (2) IInd transition series i.e. 4d series which contains 10 elements and starts from $_{39}$ Y $_{48}$ Cd. Filling of electrons takes place in 4d sub-shell.
- (3) IIIrd transition series i.e. 5d series which contains 10 elements and starts from ${}_{57}La$, ${}_{72}Hf {}_{80}Hg$. Filling of electrons takes place in 5d sub-shell.
- (4) IV^{th} transition series i.e. 6d series which contains 10 elements and starts from ₈₉Ac, ₁₀₄Rf - ₁₁₂Uub. Filling of electrons takes place in 6d sub-shell (incomplete series).
- Those elements which have partially filled d-orbitals in neutral state or in any stable oxidation state are called transition elements. All d-block elements except Zn, Cd, Hg are transition elements.

(d) f-Block elements :

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

- All f-block elements belong to 3rd group.
- They are metals
- Within each series, the properties of the elements are quite similar.
- The elements coming after uranium are called transuranium elements.
- They are also called as inner-transition elements as they contain three outer most shell incomplete and were also referred to as rare earth elements since their oxides were rare in earlier days.

The elements of f-blocks have been classified into two series.

- **1.** Ist inner transition or 4 f-series or lanthanides, contains 14 elements ₅₈Ce to ₇₁Lu. Filling of electrons takes place from 4f subshell.
- 2. IInd inner transition or 5 f-series or actinides, contains 14 elements ₉₀Th to ₁₀₃Lr. Filling of electrons takes place from 5f subshell.
- The actinides and lanthanides have been placed at the bottom of the periodic table to avoid the undue expansion of the periodic table.

Nomenclature of the elements with atomic number > 100 (IUPAC) :

According to IUPAC, elements with atomic number greater than 100 are represented by three letter symbols.

- O These symbols are based on first letter of numbers from 0 to 9. The names of these number are derived from Greek and Latin languages.
- O The Latin words for various digits of the atomic number are written together in the order of which makes the atomic number and suffix 'ium' is added at the end. In cases of bi and tri 'i' is omitted.

NOTATION FOR IUPAC NOMENCLATURE OF ELEMENTS Table-4

Digit	Name	Abbreviation
0	nil	n
1	un	u
2	bi	b
3	tri	t
4	quad	q
5	pent	р
6	hex	h
7	sept	S
8	oct	0
9	enn	е

Prediction of period, group and block :

- O Period of an element corresponds to the principal quantum number of the valence shell.
- O The block of an element corresponds to the type of subshell which receives the last electron.
- O The group is predicted from the number of electrons in the valence shell or/and penultimate shell as follows.

- (a) For s-block elements, Group number = the number of valence electrons
- (b) For p-block elements, Group number = 10 + number of valence electrons
- (c) For d-block elements, Group number = number of electrons in (n 1) d sub shell + number of electrons in the valence shell.

(58-71) 4f

(d) For f block element - (92-103) 5f \int - All f block elements are belongs to group number 3.

Solved Examples.

- Ex-4. In terms of period and group where would you locate the element with Z = 114 ?
- **Sol.** The filling of the 6th period ends at $_{86}$ Rn. Thereafter, the filling of 7th period starts. Like in 6th, period, in 7th period also, the filling of four subshells, i.e., 7s, 7p, 6d and 5*f* occurs. But according to aufbau principle their energies increase in the order : 7s < 5*f* < 6d < 7p. Therefore, after $_{86}$ Rn, the next two elements with Z = 87 and Z = 88 are s-block elements, the next fourteen, i.e., Z = 90 103 are *f*-block elements, the next ten i.e., Z = 104 112 are d-block elements and the last six, i.e., Z = 113 118 are p-block elements. Therefore, the element Z = 114 is the second p-block element (i.e., group 14) of the 7th period. Thus, the location of the element with Z = 114 in the period table is **Period** = 7th **Block** : p- Block **Group** : 14
- Ex-5. 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

 $\mathsf{E}: 1\mathsf{s}^2\, 2\mathsf{s}^2\, 2\mathsf{p}^6\, 3\mathsf{s}^2\, 3\mathsf{p}^6\, 4\mathsf{s}^2$

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 as, ns² np¹.

Metals and nonmetals :

- The metals are characterised by their (i) shiny appearance (ii) nature of readily loosing the electrons to form cation. Metals comprises more than 78% of all known elements and appear on the left hand side of the periodic table. Metals are usually solids at room temperature (except mercury, gallium). They have high melting and boiling points and are good conductors of heat and electricity. Oxides of metals are generally basic in nature (some metals in their higher oxidation state form acid oxides e.g. CrO₃).
- Nonmetals do not lose electrons but rather accept electrons to form corresponding anions. Nonmetals are located at the top right hand side of the periodic table. Nonmetals are usually solids, liquids or gases at room temperature with low melting and boiling points. They are poor conductors of heat and electricity. Oxides of nonmetals are generally acidic in nature.

Metalloids (Semi metals) :

- It can be clear from the periodic table that nonmetallic character increases as we move from left to right across a row. It has been found that some elements which lie at the border of metallic and nonmetallic behaviour, possess the properties of both metals and nonmetals. Such elements are called semi metals or metalloids.
- The metalloids comprise of the elements Si, Ge, As, Sb and Te.
- Oxides of metalloids are generally amphoteric in nature.

Typical elements :

- Third period elements are called as typical elements. These include Na, Mg, Al, Si, P, S, Cl.
- The properties of all the elements belonging to a particular group resemble the properties of the corresponding typical element of that group. For example, the general properties of alkali metals (IA) can be predicted from the properties of Na, not Li, the first member of the group.
- The properties of the elements of second period differ in the many respect belonging to the same group due to the smaller atomic size and absence of vacant d-orbitals.

Diagonal relationship :

Some elements of certain groups of 2nd period resemble much in properties with the elements of third period of next group i.e. elements of second and third period are diagonally related in properties. This phenomenon is known as diagonal relationship. For example, the similarity between lithium (the first member of group 1) and magnesium (the second element in group 2) is called a diagonal relationship. Diagonal relationship also exist between other pairs of elements Be and AI, B and Si as shown in figure



(i) on descending a group, the atoms and ions increase in size. On moving from left to right in the periodic table, the size decreases. Thus on moving diagonally, the size remains nearly the same.

(ii) it is sometimes suggested that the diagonal relationship arises because of diagonal similarity in electronegativity values.

(Li = 1.0 & Mg = 1.2 ; Be = 1.5 & Al = 1.5 ; B = 2.0 & Si = 1.8)

(iii) Be and Al also show a diagonal relationship. In this case sizes are not so close (Be₂₊ = 0.45 Å and Al₃₊ = 0.535 Å) but the charge / ionic radius ratio is nearly similar because the charges are 2+ and 3+ (lonic charge)

respectively. Charge /Ionic radius ratio = (Ionic radius)

Section (B) : Shielding Effect & Zeff



Effective nuclear charge (Zeff) : The outer electron experiences two opposing force : -

(i) Force of attraction from nucleus.

(ii) Force of repulsion from inner electrons.

Suppose inner electrons are not present there then force of attraction experience by test electron must be greater than it actually experienced because electronic repulsion weakens the force of attraction. This reduced nuclear charge or nuclear actually experienced by an electron is termed as effective nuclear charge.

 $Z_{eff} = Z - \sigma$

Z_{eff} = Effective nuclear charge, Z = Atomic number

 σ = Screening constant or shielding constant or slater constant

Effective nuclear charge is not same for all the electrons present in an atom and it varies with distance between electron and nucleus. This effect in which inner electrons shield valence electron from the attraction due to nucleus is called shielding effect. It is also known as screening effect as inner electron create a protective screen between nucleus and valence electrons. Due to shielding effect valence electron experiences less attraction from the side of nucleus as if inner electrons are absent. s-orbitals have the largest screening effect for a given n value since s electrons are closer to the nucleus. Screening effect decreases as s-orbital's > p-orbital's > d-orbital's > f-orbital's.



Section (C) : Oxidation states & Inert pair effect Oxidation states :

The valence of representative elements is usually (though not necessarily) equal to the number of electrons in the outermost orbitals and / or equal to eight minus the number of outermost electrons. Nowadays the term oxidation state is frequently used for valence. Consider the two oxygen containing compounds : OF_2 and Na_2O . The order of electronegativity of the three elements involved in these compounds is F > O > Na. Each of the atoms of fluorine, with outer electronic configuration $2s^22p^5$, shares one electron with oxygen in the OF_2 molecule. Being highest electronegative element, fluorine is given oxidation state–1. Since there are two fluorine atoms in this molecule, oxygen with outer electronic configuration $2s^22p^4$ shares two electrons with fluorine atoms and thereby exhibits oxidation state +2. In Na_2O , oxygen being more electronegative accepts two electrons, one from each of the two sodium atoms and thus, shows oxidation state –2.

On the other hand sodium with electronic configuration 3s¹ loses one electron to oxygen and is given oxidation state +1. Thus, the oxidation state of an element in a particular compound can be defined as the charge acquired by its atom on the basis of electronegative consideration from other atoms in the molecule. There are many elements which exhibit variable valence. This is particularly characteristic of transition elements and actinoids.

Inert pair effect :

The outer shell 's' electrons (ns²) penetrate to (n - 1) d electrons and thus become closer to nucleus and are more effectively pulled towards the nucleus. This results in less availability of ns² electron pair for bonding of ns² electron pair becomes inert. The inert pair effect begins after $n \ge 4$ and increases with increases value of n.

(i) Dominance of lower oxidation state on moving down the group in boron, carbon & nitrogen family may also be explained by the inert pair effect.

(ii) The inert pair effect is not the explanation of why monovalency , bivalency and trivalency in group 13,

14 & 15. It merely describe what happens, i.e. two electrons do not participate in bonding. The reason that they do not take part in bonding in energy.

Order of stability of oxidation state according to Inert pair effect :

Table-5

List of common oxidation sate of an element in Periodic Table

1																	18
1 H +1 -1	* -																2 He
3 Li +1	2 3 Be +2											13 5 B +3 -3	14 6 C +4 +2 -4 etc.	15 7 +5 +4 +3 +1 -3 0	16 8 • 2 -1/2 -1 -1 -2	17 9 F –1	10 Ne
11 Na +1	12 Mg +2	3	4	5	6	7	8	9	10	11	12	13 Al +3	14 Si +4 -4	etc. 15 P +3 +1 -3	16 S +6 +4 +2 -2	17 CI +5 +7 +3 +1 -1 etc	18 Ar 0
19 K +1	20 Ca +2	21 Sc +2 +3	+2 +2 +3 +4	23 V +2 +3 +4 +5	24 Cr +1 +2 +3 +4 +5 +6	25 Mn +1 +2 +3 +4 +5 +6 +7	26 Fe +1 +2 +3 +4 +5 +6	27 Co +1 +2 +3 +4 +5	28 Ni +1 +2 +3 +4	29 Cu +1 +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 TI +3 +1	82 Pb +4 +2	83 Bi +5 +3	84 Po	85 At	86 Rn

Section (D) : Atomic and Ionic Radius

Atomic radius :

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

Atomic radius can be

(A) **Covalent radius :** It is one-half of the distance between the centres of two nuclei (of like atoms) bonded by a single covalent bond as shown in figure.



- Covalent radius is generally used for nonmetals.
- (B) Vander Waal's radius (Collision radius) : It is one-half of the internuclear distance between two adjacent atoms in two nearest neighbouring molecules of the substance in solid state as shown in figure.



 Vander Waal's radius does not apply to metal. Its magnitude depends upon the packing of the atoms when the element is in the solid state.

Comparison of Covalent radius and Vander Waal's radius

- The Vander Waal's force of attraction is weak, therefore, the internuclear distance in case of atoms held by Van der Waal's force is much larger than those of held covalent bond. Therefore Vander Waal's radii are always larger than covalent radii.
- A covalent bond is formed by the overlapping of two half-filled atomic orbitals and in this a part of the orbital becomes common. Therefore, covalent radii are always smaller than the Van der Waals radii. For example,

		lable	e-6		
Elements	н	0	F	S	Br
Covalent radius (Å)	0.37	0.66	0.64	1.04	1.11
van der Waal's radius (Å)	1.20	1.40	1.35	1.85	1.95

(C) Metallic radius (Crystal radius) :

It is one-half of the distance between the nuclei of two adjacent metal atoms in the metallic crystal lattice as shown in the figure.



Metallic radius of an element is always greater than its covalent radius. It is due to the fact that metallic bond (electrical attraction between positive charge of an atom and mobile electrons) is weaker than covalent bond and hence the internuclear distance between the two adjacent atoms in a metallic crystal is longer than the internuclear distance between the covalently bonded atom.

For example :		Metallic radius	Covalent radius
	K	231 pm	203 pm
	Na	186 pm	154 pm

Thus the magnitude of the above three radii follow the below order.

 $r_{covalent} < r_{crystal} < r_{vander Walls}$

lä	adie-7
Variation in a Period (Left to Right)	Variation in a Group (Top to Bottom)
1. Nuclear charge (Z) increases by one unit	1. Nuclear charge (Z) increases by more than one unit
2. Effective nuclear charge (Z_{eff}) also increases	 Effective nuclear charge (Z_{eff}) almost remains constant because of increased screening effect of inner shells electrons.
3. As a result, the electrons are pulled closer to the nucleus by the increased $Z_{eff}.$ $r_n \propto \frac{1}{Z^*}$	3. The effect of increased number of atomic shells overweighs the effect of increased nuclear charge. As a result of this the size of atom increases from top to bottom in a given group.
Hence atomic radii decrease with increase in atomic number in a period from left to right.	
4. But number of orbitals (n) remains constant	4. But number of orbitals (n) increases.

abla 7

Ο The atomic radius of inert gases (zero group) is quoted as highest in magnitude in their respective period because it is sine for then Vander Waal's radius is quoted. The Vander Waal's radius of inert gases also increases from top to bottom in a group.

Ionic radius :

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

Та	ble-8
Cation	Anion
 (1) It is formed by the loss of one or more electrons from the valence shell of an atom of an element. (2) Cations are smaller than the parent atoms because, (a) the whole of the outer shell of electrons is usually removed. (b) in a cation, the number 	 (1) It is formed by the gain of one or more electrons from in the valence shell of an atom of an element. (2) Anions are larger than the parent atoms because (a) anion is formed by gain of one or more electrons in the neutral atom and thus number of el

The sizes of ions increases as we go down a group (considering the ions of same charge). For example :

> $Li^{+}(0.76) < Na^{+}(1.02) < K^{+}(1.38) < Rb^{+} \rightarrow (in Å)$ Be²⁺ < Mg²⁺ < Ca²⁺ < Sr²⁺ $F^- < CI^- < Br^- < I^-$

- ♦ For the cations same element the ionic radius decreases as more electrons are lost (i.e. the valency increases) For e.g. Cr²⁺ = 0.80 Å, Cr³⁺ = 0.615 Å, Cr⁴⁺ = 0.55 Å, Cr⁵⁺ = 0.49 Å and Cr⁶⁺ = 0.44 Å.
- The d and f orbitals do not shield the nuclear charge very effectively .Therefore there is significant reduction in the size of the ions, just after d or f orbitals have been filled completely. This is called a lanthanide contraction. Atomic radii of Hf (Hf⁴⁺ = 0.71 Å) and Zr (Zr⁴⁺ = 0.72 Å), and Nb (Nb³⁺ = 0.72 Å) and Ta (Ta³⁺ = 0.72 Å) are almost identical due to lanthanide contraction.
- ♦ The species containing the same number of electrons but differ in the magnitude of their nuclear charges are called as isoelectronic species. For example, N³⁻, O²⁻, F⁻, Ne, Na⁺, Mg²⁺ and Al³⁺ are all isoelectronic species with same number of electrons (i.e 10) but different nuclear charges of +7, +8, +9, +10, +11, +12 and +13 respectively.

Within a series of isoelectronic species as the nuclear charge increases, the force of attraction by the nucleus on the electrons also increases. As a result, the ionic radii of isoelectronic species decrease with increases in the magnitude of nuclear charges. For example, as shown in figure.

Al³⁺ Mg²⁺ Na⁺ F⁻ O²⁻ N³⁻

Ionic radii increase

As effective nuclear charge decrease.

Following are the examples of isoelectronic series

(i) S^{2-} , CI^- , K^+ , Ca^{+2} , Sc^{+3} (ii) SO_2 , NO_3^- , CO_3^{2-} , (iii) N_2 , CO, CN^- (iv) NH_3 , H_3O^+

• Pauling's empirical formula for ionic radius \propto ^{nuclear charge} (only for isoelectronic species).

Solved Examples.

- **Ex-6.** Consider the following species :
 - N³⁻ , O²⁻ , F⁻ , Na⁺ , Mg²⁺ , and Al³⁺
 - (a) What is common in them?
 - (b) Arrange them in the order of increasing ionic radii.
 - (a) Each one of these ions contains 10 electrons and hence all are isoelectronic ions.

(b) The ionic radii of isoelectronic ions decrease with the increase in the magnitude of the nuclear charge. For example, consider the isoelectronic ions : N^{3-} , O^{2-} , F^- , Na^{2+} and AI^{3+} . All these ionse have 10 electrons but their nuclear charges increase in the order :

 N^{3-} (+ 7), O^{2-} (+ 8), F^{-} (+ 9), Na^{+} (+ 11), Mg^{2+} (+ 12) and AI^{3+} (+ 13). Therefore, their ionic radii decrease in the order : $N^{3-} > O^{2-} > F^{-} > Na^{+} > Mg^{2+} > AI^{3+}$.

- Ex-7. Mg₂₊ is smaller than O₂₋ in size, though both have same electronic configuration. Explain?
- **Sol.** Mg₂₊ and O₂₋ both are isoelectronic i.e., have same number of electrons. But Mg₂₊ having 12 protons in its nucleus exerts higher effective nuclear charge than O₂₋ having 8 protons and thus valence shell as well as inner shells electrons are more strongly attracted by the nucleus in Mg₂₊ resulting smaller size than O₂₋.

Section (E) : Ionisation Energy

Ionisation energy (IE) is defined as the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom to form a cation.

 $\mathsf{M}(g) \xrightarrow{\quad (\mathsf{IE}_1) \quad} \mathsf{M}^+(g) + e^- \hspace{0.2cm} ; \hspace{0.2cm} \mathsf{M}^+ + \mathsf{IE}_2 \xrightarrow{\quad } \mathsf{M}^{2+} + e^- \hspace{0.2cm} ; \hspace{0.2cm} \mathsf{M}^{2+} + \mathsf{IE}_3 \xrightarrow{\quad } \mathsf{M}^{+3} + e^-$

 IE_1 , $IE_2 \& IE_3$ are the Ist, $II^{nd} \& III^{rd}$ ionization energies to remove electron from a neutral atom, monovalent and divalent cations respectively.

In general for the same element, $(IE)_1 < (IE)_2 < (IE)_3 < \dots$ because, as the number of electrons decreases, the attraction between the nucleus and the remaining electrons increases considerably and hence subsequent ionization energies increase.

Sol.

Ionization energies are determined from spectra and are measured in kJ mol⁻¹, k Cal mol⁻¹, eV (electron volt).

• Factors Influencing Ionisation energy

Variation in ionization energies in a period and group may or may not be regular and can be influenced by the following factors.

(A) Size of the Atom (Dominant factor in case of group) :

Generally ionisation energy decreases with increase in the atomic size.

As the distance between the outer most electrons and the nucleus increases, the force of attraction between the valence shell electrons and the nucleus decreases. As a result, outer most electrons are held less firmly and lesser amount of energy is required to knock them out.

For example, ionisation energy decreases continuously in a group from top to bottom with increase in atomic size. But in 13th & 14th group the ionisation energy does not decrease continuously and order is as follows:

For 13th group: B > TI > Ga > AI > InFor 14th group: C > Si > Ge > Pb > Sn

(B) Nuclear Charge (Dominant factor in case of period) :

The ionisation energy increases with increase in the nuclear charge.

This is due to the fact that with increase in the nuclear charge, the electrons of the outer most shell are more firmly held by the nucleus and thus greater amount of energy is required to pull out an electron from the atom.

For example, ionisation energy increases as we move from left to right along a period due to increase in nuclear charge.

(C) Shielding or screening effect :

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



(D) Penetration effect of the electron :

The ionization energy also depends on the type of the electron which is to be removed. Reason being : s, p, d and f electrons have orbitals with different shapes. An s electron penetrates closer to the nucleus, and is therefore more tightly held than a p electron. Similarly p-orbital electron is more tightly held than a d-orbital electron and a d-orbital electron is more tightly held than an f-orbital electron. If other factors being equal, ionisation energies are in the order s > p > d > f.

For example, ionisation energy of aluminium is comparatively less than magnesium because outer most electron is to be removed from 3p-orbital (having lesser penetration effect) in aluminium where as in magnesium it will be removed from 3s-orbital (having larger penetration effect) of same energy level.

(E) Electronic Configuration :

If an atom has exactly half-filled or completely filled orbitals, then such an arrangement is said to have extra stability.

The removal of an electron from such an atom requires more energy than expected. For example, first ionisation energy of beryllium is greater than boron because beryllium has extra stable completely filled outer most 2s orbital while boron has partially filled less stable outer most 2p-orbital.

Be (Z = 4) $1s^2$, $2s^2$ B (Z = 5) $1s^2$, $2s^2$, $2p^1$

Similarly noble gases have completely filled electronic configurations and hence they have highest ionisation energies in their respective periods.

- O Metallic or electropositive character of elements increases as the value of ionisation energy decreases.
- O The relative reactivity of the metals in gaseous phase increases with the decrease in ionisation energy.
- O The reducing power of elements in the gaseous phase increases as the value of ionisation energy decreases. Among alkali metals, lithium is the strongest reducing agent in aqueous solution.

Solved Examples

- **Ex-8.** Energy of an electron in the ground state of the hydrogen atom is -2.18×10^{-18} J. Calculate the ionization enthalpy of atomic hydrogen in terms of J mol⁻¹.
- **Sol.** Ionization energy is the amount of energy required to remove the electron from the ground state to infinity. Now, energy of the electron in the ground state = -2.18×10^{-15} J Energy of the electron at affinity = 0.

The energy required to remove an electron in the ground state of hydrogen atom

 $= 0 - (its energy in the ground state) = -(-2.18 \times 10^{-18} J) = 2.18 \times 10^{18} J.$

$$2.18 \times 10^{18} \times 6.02 \times 10^{23}$$

- 1000 kJ
- lonization enthalpy per mole of hydrogen atoms = $= 1312.36 \text{ kJ mol}^{-1} = 1312.36 \times 10^3 \text{ J mol}^{-1}$.
- **Ex-9.** Among the second period elements the actual ionization enthalpies are in the order Li < B < Be < C < O < N < F < Ne.

Explain why :

:..

(i) Be has higher $\Delta_i H$ than B

(ii) O has lower $\Delta_I H$ than N and F?

Sol. (i) The ionization enthalpy, among other things depends upon the type of electron to be removed from the same principal shell. In case of Be $(1s^2 2s^2)$ the outermost electron is present in 2s-orbital while in B $(1s^2 2s^2 2p^1)$ it is present in 2p- orbital. Since 2s-electrons are more strongly attracted by the nucleus than 2p-electrons, therfore, lesser amount of energy is required to knock out a 2p-electron than a 2s-electron. Consequently, Δ_i H of Be is higher than that Δ_i H of B.

(ii) The electronic configration of N $(1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1)$ in which 2p-orbitals are exactly half-filled is more stable than the electronic configuration of O $(1s^2 2s^2 p_x^2 2p_y^1 2p_z^1)$ in which the 2p- orbitals are neither exactly half-filled nor completely filled. Therefore, it is difficult to remove an electron from N than from O. As result, Δ_i H of N is higher than that of O. Further, the electronic configuration of F is $1s^2 2s^2$ $2p_x^2 2p_y^2 2p_z^1$. Because of higher nuclear charge (+9) the first ionization enthalpy of F is higher than that of O. Further, the effect of increased nuclear charge outweights the effect of stability due to exactly halffilled orbitals, therefore, the Δ_i H of N and O are lower than that of F.

Ex-10. Following graph shows variation of ionization energies with atomic number in second period (Li–Ne). Value of ionization energies of Na (11) will be :





(4) between N and O.

(1) above Ne

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 (3) is correct.

Section (F) : Electron gain enthalpy (Electron affinity)

Electron affinity is conventionally defined as the energy released when an electron is added to the valence shell of an isolated gaseous atom.

 $F + e_{-} - \rightarrow F_{-}$ $E_{a} = 328 \text{ kJ mol}_{-1}$

A positive electron affinity indicates that the ion X_- has a lower, more negative energy than the neutral atom X. The second electron gain enthalpy, the enthalpy change for the addition of a second electron to an initially neutral atom, invariably positive because the electron repulsion out weighs the nuclear attraction.

 $O(g) + e_{-}(g) \xrightarrow{\text{Exothermic}} O_{-}(g) \text{ ; } E_a = 140.9 \text{ kJ mol}_{-1} \text{ ; } \Delta_{eg}H_{\Theta} = -140.9 \text{ kJ mol}_{-1}$

 $O_{-}(g) + e_{-}(g) \xrightarrow{\text{Endothermic}} O_{2-}(g) ; \qquad E_a = -744 \text{ kJ mol}_{-1} ; \Delta_{eg}H_{\Theta} = 744 \text{ kJ mol}_{-1}$

An element has a high electron affinity if the additional electron can enter a shell where it experiences a strong effective nuclear charge.

The electron gain enthalpy $\Delta_{eg}H_{\Theta}$, is the change in standard molar enthalpy when a neutral gaseous atom gains an electron to form an anion.

 $X (g) + e_{-} (g) - \rightarrow X_{-} (g)$

Now a days both electron affinity and electron gain enthalpy terms are used.

Both are same in magnitude but opposite in sign.

E.A. = +ve (Exothermic):- $\Delta_{eg}H_{\Theta} < 0$.

E.A. = -ve (Endothermic):- $\Delta_{eg}H_{\Theta} > 0$.

The units of EA and Electron gain enthalpy are ev/atom or Kcal / mole or KJ / Mole.

Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion. Electron gain may be either exothermic or endothermic depending on the elements. When an elec tron is added to the atom and the energy is released, the electron gain enthalpy is negative and when energy is needed to add an electron to the atom, the electron gain enthalpy is positive.

Although the electron gain enthalpy is the thermodynamically appropriate term, much of inorganic chemistry is discussed in terms of a closely related property, the electron affinity, E_a of an element which is the difference in energy between the gaseous atoms and the gaseous ions.

$$E_a = E(X, g) - E(X_{-}, g)$$

- O Group 17 elements (halogens) have very high negative electron gain enthalpies (i.e. high electron affinity) because they can attain stable noble gas electronic configuration by picking up an electron.
- O Across a period, with increase in atomic number, electron gain enthalpy becomes more negative because left to right across a period effective nuclear charge increases and consequently it will be easier to add an electron to a small atom.
- O As we move in a group from top to bottom, electron gain enthalpy becomes less negative because the size of the atom increases and the added electron would be at larger distance from the nucleus. In general the electron gain enthalpies for some third period element (e.g. P, S, Cl) are more negative than the corresponding second period members (e.g. N, O, F). This is due to the smaller size of the atoms of the second period elements which would produced larger electron-electron repulsions for the additional electron.

PERIODIC TABLE & PERIODICITY

Group 1	$\Delta_{e.g.}$ H	Group 16	$\Delta_{e.g.}$ H	Group 17	$\Delta_{e.g.}$ H	Group 0	$\Delta_{e.g.}$ H
Н	-73					He	+48
Li	-60	0	-141	F	-328	Ne	+116
Na	-53	S	-200	CI	-349	Ar	+96
K	-48	Se	-195	Br	-325	Kr	+96
Rb	-47	Те	-190	l	-295	Xe	+77
Cs	-46	Po	-174	At	-270	Rn	+68

- O Noble gases have large positive electron gain enthalpies because the electron has to enter the next higher energy level leading to a very unstable electronic configuration.
- O Negative electron gain enthalpy of O or F is less than S or Cl. This is due to the fact that when an electron is added to O or F, the added electron goes to the smaller n = 2 energy level and experiences significant repulsion from the other electrons present in this level. In S or Cl, the electron goes to the larger n = 3 energy level and consequently occupies a larger region of space leading to much less electron-electron repulsion.
- O Electron gain enthalpies of alkaline earth metals are very less or positive because the extra electron is to be added to completely filled s-orbitals in their valence shells.
- O Nitrogen has very low electron affinity because there is high electron repulsion when the incoming electron enters an orbital that is already half filled.
- O (i) Electron affinity $\propto \frac{1}{\text{Atomic size}}$

(ii) Electron affinity \propto Effective nuclear charge (z_{eff})

(iii) Electron affinity $\propto \frac{1}{\text{Screening effect}}$

(iv) Stability of half filled and completely filled orbitals of a subshell is comparatively more and the addition of an extra electron to such an system is difficult and hence the electron affinity value decreases.

Solved Examples.

- **Ex-11.** 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-12.** Why do halogens have high electron gain enthalpies (i.e. $-\Delta_{eg}H^{\Theta}$)?
- **Sol.** The valence shell electronic configuration of halogens is ns²np⁵ and thus they require one electron to acquire the stable noble gas configuration ns²np⁶. Because of this they have strong tendency to accept an additional electron and hence have high electron gain enthalpies.
- **Ex-13.** 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]?
 - (1) O^x (2) O^y (3) O^z (4) 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 O²⁻. So option (1) is correct.
- **Ex-14.** 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]2s¹ and [He]2s², 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.

Section (G) : Electronegativity

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

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

- O With increase in atomic size the distance between nucleus and valence shell electrons increases, therefore, the force of attraction between the nucleus and the valence shell electrons decreases and hence the electronegativity values also decrease.
- O With increase in nuclear charge force of attraction between nucleus and the valence shell electrons increases and, therefore, electronegativity value increases

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

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

Hybrid orbital	sp ³	sp²	sp
s-character	25%	33%	50%

Electronegativity increases

Electronegativity of some elements according to pauling scale.

Elements	Н	Li	Be	В	С	Ν	0	F	Ne	Ρ	S	CI	Br	Ι
Electronegativity	2.1	1.0	1.5	2.0	2.5	3.0	3.5	4.0	0.9	2.1	2.5	3.0	2.8	2.5

Table-9						
Variation of electronegativity in a group	Variation of electronegativity in a period					
On moving down the groups, Z increases but Z_{eff} almost remains constant, number of shells (n) increases, r_n (atomic radius) increases. Therefore, electronegativity decreases moving down the groups.	While moving across a period left to right, Z, Z _{eff} increases & r _n decreases. Therefore, electronegativity increases along a period.					

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

(a) **Pauling's scale :** Linus Pauling developed a method for calculating relative electronegativities of most elements. According to Pauling

$$\begin{split} &\Delta = X_A - X_B = O.208 \ \sqrt{E_{\cdot A - B} - \sqrt{E_{A - A} \times E_{B - B}}} \\ &E_{A \cdot B} = \text{Bond enthalpy/ Bond energy of } A - B \text{ bond.} \\ &E_{A \cdot A} = \text{Bond energy of } A - A \text{ bond} \\ &E_{B - B} = \text{Bond energy of } B - B \text{ bond} \end{split}$$

(All bond energies are in kcal / mol)

$$\Delta = X_A - X_B = O.1017 \sqrt{E_{\cdot A - B} - \sqrt{E_{A - A} \times E_{B - B}}}$$

All bond energies are in kJ / mol.

(b) Mulliken's scale : Electronegativity χ (chi) can be regarded as the average of the ionisation energy (IE) and the electron affinity (EA) of an atom (both expressed in electron volts).

$$\chi_{\rm M} = \frac{\rm IE + EA}{2}$$

Paulings's electronegativity χ_P is related to Mulliken's electronegativity χ_M as given below.

$$\chi_{\rm P} = 1.35 \ (\chi_{\rm M})^{1/2} - 1.37$$

Mulliken's values were about 2.8 times larger than the Pauling's values.

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

Xc-XH = 0.208 [Ec-H -
$$\frac{1}{2}$$
 (Ec-c + EH-H)]^{1/2}
xc - 2.1 = 0.208 [98.8 - $\frac{1}{2}$ (83.1 + 104.2)]^{1/2}
⇒ xc = 2.5

- **Ans.** 2.5
- Ex-16. If electronegativity of x be 3.2 and that of y be 2.2, the percentage ionic character of xy is (1) 19.5 (2) 18.5 (3) 9.5 (4) 29.5.

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

 $[\Delta = difference of electronegativity values between x and y].$

% ionic character = $16\Delta + 3.5\Delta^2 = 19.5$.

Ans. (1)

Sol.

Table-10 Periodic trends



MISCELLANEOUS SOLVED PROBLEMS

1. The correct order of the metallic character is :

(1) Na > Mg > Al > Si (2) Mg > Na > Al > Si (3) Al > Mg > Na > Si (4) Si > Al > Na > Mg

- **Sol.** (1) The metallic character of the elements is highest at the extremely left (low ionisation energies) and then decreases across the period from left to right (ionisation energies increases across the period).
- 2. The correct order of the non-metallic character is :

(1) B > C > Si > N > F	(2) Si > C > B > N > F
(3) F > N > C > B > Si	(4) F > N > C > Si > B

- **Sol.** (3) The non-metallic character of the elements is highest at the extremely right (high ionisation energies and high negative values of electron gain enthalpies) and then decreases from right to left across the period (ionisation energies decrease and also negative values of electron gain enthalpies decrease from right to left).
- 3. Which of the following group does not represent the isoelectronic species ?
 - (1) CH₄, H₂O, NH₃, HF (3) PH₃, SiH₄, HS⁻, Ar
 - (3) OH⁻, H₂O, NH₂⁻, F⁻ (4) H₂S, K⁺, Ar, CI
- **Sol.** (4) As all species do not have same number of electrons i.e., 18. H₂S, K⁺ and Ar have same number of electrons i.e. 18 but Cl has 17 electrons.
- **4.** A M²⁺ ion derived from a metal in the first transition metal series has four electrons in 3d subshell. What element might M be ?
- Sol.Electron configuration of M2+ is :
Electron configuration of M is :
So total number of electrons = 24.
Hence, metal M is chromium (Cr).[Ar]18 4s⁰ 3d⁴
[Ar]18 4s¹ 3d⁵ (and not 4s² 3d⁴)
- 5.Following are the valence shell electronic configurations of some elements.
(i) 3s2 3p5(ii) 3d10 4s2(iii) 2s2 3p6 4s1(iv) 1s2 2s2Find out the blocks to which they belong in the periodic table ?
- Ans. (i) p-block (ii) d-block (iii) s-block (iv) s-block
- **Sol.** The block of the elements depend on the type of sub-shell which receive the last electron. In case of (i) it enters in 3p-subshell, (ii) it enters 3d-subshell, (iii) it enters 4s-subshell and (iv) it enters 2s-subshell.
- 6. Find out the group of the element having the electronic configuration, 1s₂ 2s₂ 2p₆ 3s₂ 3p₆ 3d₆ 4s₂.
- **Ans.** As last electron enters in d-subshell, therefore this belongs to d-block. For d-block element the group number is equal to the number of valence shell electrons + number of electrons in (n-1) d-subshell. So, group number = 6 + 2 = 8.
- 7. Arrange the following ions in the increasing order of their size : Be₂₊ , Cl₋, S₂₋ , Na₊ , Mg₂₊ , Br₋ ?

Ans. Sol.	Be ₂₊ < Mg ₂₊ < Na ₊ < Cl ₋ < S ₂₋ < Br ₋ Be ₂₊ is smaller than Mg ₂₊ as Be ₂₊ has one shell where as Mg ₂₊ has two shells. Mg ₂₊ and Na ₊ are isoelectronic species : lonic radius \propto 1/nuclear charge. Cl ₋ and S ₂₋ are isoelectronic species : lonic radius \propto 1/nuclear charge. Cl ₋ is smaller than Br ₋ as Cl ₋ has three shells where as Br ₋ has four shells.							
8.	The (IE1) and the (IE2) in kJ mol-1 of a few elements designated by Roman numerals are shown below:IIIIIIIE14035491142							
Ans.	IE2 2640 1060 2080 Which of the above elements is likely to be a							
9.	lonisation energy and electron affinity of fluorine are respectively 17.42 and 3.45 eV. Calculate electronegativity of fluorine atom.							
Sol.	According to Mulliken's e 10.435	lectronegativity (ҳм)=	$\frac{17.42 + 3.45}{2} = \frac{17.42 + 3.45}{2} = \frac{17.42 + 3.45}{2}$				
			$\frac{10.435}{2.8}$					
Anc	I herefore, electronegativity $v_{\rm P} = 2.726$	on Pauling's scale ($\chi_{\rm P}$) = 2.0 = $\chi_{\rm P}$	3.726				
10. Sol.	$\chi P = 3.726$ Are the oxidation state and No. The oxidation state of A	covalency of AI in [A I is +3 and the cova	ICI (H₂O)₅]²+ sa lency is 6.	ame ?				
11. Sol.	Why the electron gain enthalpy values of alkaline earth metals are lower (i.e. less negative) or positive ? The general valence shell electron configuration of alkaline earth metals is ns ² (stable configuration). The extra electron must enter np subshell, which is effectively shielding by the two ns electrons and the inner electrons. Consequently, the alkaline earth metals have little or no tendency to pick up an extra electron.							
12.	In Column-I, there are give metals given in Column-II :	n electronic configu	rations of some	e elements. Match these with the correct				
	Column-l			Column-II				
	(1) ns^2 , np^5		(p)	Chromium				
	(2) $(II - I) d^{5}$, IIS ¹ (3) $(n - 1) d^{5} ns^{1}$		(q) (r)	Krypton				
	(4) $(n-1) d^{10}$, ns ² , np	6	(r) (s)	Bromine				
Ans.	$(1) \rightarrow (s)$; $(2) \rightarrow (q)$; $(3) \rightarrow$	(p) ; (4) → (r).						
Sol.	(1) ns ² np ⁵ is general valen	ce shell electron cor	figuration of ha	alogens. So this configuration belongs to				
	(2) $(n - 1) d^{1-10} ns^{1-2}$; This is electron configuration of d-block elements. As it contains $(n - 1) d^{10}ns^{1-10}$							
	(3) $(n - 1) d^{1-10} ns^{1-2}$; This is electron configuration of d-block elements. As it contains $(n - 1) d^5 ns^1$							
	configuration it belongs to c	hromium. e shell electron conf	iouration ns ² nn	6 so it belongs to knypton				
12	Match the metale given in C		iyaraliyan in Cal					
Column-I Column-I with their type given in Column-I Column-I								

PERIODIC TABLE & PERIODICITY

- (1) Metalloid
- (2) Radioactive
- (3) Transition metal
- (4) Chalcogen

- Sulphur
- (q) Gold
- (r) Arsenic(s) Uranium
- (s) Uranii

(p)

- **Ans.** (1) \rightarrow (r); (2) \rightarrow (s); (3) \rightarrow (q); (4) \rightarrow (p)
- **Sol.** (1) Arsenic is a metalloid because it behaves as metal (forming cation, As³⁺-AsCl₃) as well as nonmetal (forming anion, As³⁻-AsH₃).
 - (2) Uranium is a radioactive element.
 - (3) Those elements which in their neutral atoms or in most common oxidation state have partially filled d-orbitals are called as transition elements. Gold in its +3 oxidation state has electron configuration $[Xe]^{54}$, $5d^86s^0$.

(4) 16th group elements like oxygen and sulphur are ore forming elements and therefore are called as chalcogens.

- 14. Match the metals given in Column-II with their type given in Column-I:
 - Column-I Column-II (1) Representative element Cerium (p) (2) Lanthanide (q) Aluminium (3) Coinage metal (r) Thorium (4) Actinide Gold (s)
- **Ans.** (1) \rightarrow (q); (2) \rightarrow (p); (3) \rightarrow (s); (4) \rightarrow (r)
- **Sol.** (1) s-block and p-block elements are collectively called as representative elements. As in aluminium last electron enters in p-subshell ([Ne]¹⁰3s²3p¹).

(2) Lanthanide series follows lanthanum (atomic number 57) and starts from cerium (atomic number 58) to lutetium (atomic number 71), fourteen 4f- series elements.

(3) Group 11- transition elements copper, silver & gold are known as coinage metals (used for making the coins).

(4) Actinides series follows actinium (atomic number 89) and starts from thorium (atomic number 90) to lawrencium (atomic number 103), fourteen 5f- series elements.

15. Match the particulars given in Column-I with the process/metal / species given in Column-II.

Column-l

- (1) Isoelectronic species
- (2) Half filled orbital
- (3) Second ionisation energy
- (4) Inner transition element

Ar, K+, Ca++

Column-II

 $A^+(g)$ + energy $\rightarrow A^{++}(g)$ + $e^-(g)$

(r) Lutetium

(p)

(q)

- (s) Antimony
- **Ans.** (1) \rightarrow (q) ; (2) \rightarrow (s) ; (3) \rightarrow (p) ; (4) \rightarrow (r)
- Sol. (1) Species having same number of electrons but different nuclear charge are called isoelectronic species. Ar, K⁺ & Ca⁺⁺ have same number of electrons i.e. 18 but 18, 19 & 20 number of protons respectively.

(2) np³, (n–1) d⁵ and (n–2) f⁷ represent half filled orbitals. Antimony has ([Kr]³⁶ 4d¹⁰5s²5p³).

(3) The energy required to remove an electron from an univalent cation(g) is called second ionisation energy.

(4) 4f and 5f- series elements are called inner transition elements because they have three outer most shells incomplete.

16. Match the type of elements / characteristic of the elements listed in Column-I with the correct element listed in Column-II.

	Column-l		Column-II
(1)	Highest 1st ionisation energy	(p)	Technitium
(2)	Highest electronegativity	(q)	Lithium

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(3) Synthetic element

(r) Helium

(4) Strongest reducing agent

- (s) Fluorine
- **Ans.** (1) \rightarrow (r); (2) \rightarrow (s); (3) \rightarrow (p); (4) \rightarrow (q).
- **Sol.** (1) Helium has highest 1st ionisation energy amongst all the elements of periodic table because of ns² valence electron configuration and its small size of atom.
 - (2) Fluorine has highest electronegativity i.e. 4.0 on Pauling scale on account of its small size.
 - (3) Technitium is a man made element.

(4) Lithium is a strongest reducing agent because of its highest negative value of E^o due to its higher hydration energy on account of its small size of atom.

17. The Column-I has certain details about the elements of s-, p- and d-block elements. Match those with the group number of the elements listed in Column-II.

Column-I Column-II (element / elements) (group number) (1) An element whose fourth shell contains two p-electrons (p) 8th group (2) An element whose valence shell contains one unpaired p-electron (q) 12th group (3) An element which receives last electron in (n - 1) d-subshell (r) 14th group (4) An element with the ground-state electron configuration [Ar]4s₂3d₁₀ (s) 17th group Ans. (1) \rightarrow (r); (2) \rightarrow (s); (3) \rightarrow (p, q); (4) \rightarrow (q). Sol. (1) [Ar]3d¹⁰4s²4p² : Fourth shell contains two electron in 4p-sub shell i.e., 4p². Therefore, group number = 10 + 4 = 14. (2) Halogens (i.e. group number 17) have valence shell electronic configuration ns²np⁵ and there is one 1111 unpaired electron in p-subshell i.e..

(3) The element in which last electron enters in d-subshell belongs to d-block. For d-block elements the

group number = number of electrons in valence shell + number of electrons in (n - 1) d-subshell. Group number 8. Valence shell electronic configuration is $ns^2(n-1)d^6$. Therefore, group number = 2 + 6 = 8.

Like wise, group 12 is $ns^2(n-1)d^{10}$. Therefore, group number = 2 + 10 = 12. So in group 8 and 12 last electron enters in d-subshell.

(4) For electronic configuration. [Ar] $4s^23d^{10}$ the group number = 2 + 10 = 12.