PERIODIC TABLE & PERIODICITY

Introduction :

The modern periodic table which is also known as long form periodic table grew out of the investigations of several chemists such as Dobereiner (1829), Newland (1864) and Mendeleev (1869). where the attempts of Dobereiner (Triads of elements) and Newland (law of octave) had no logical justification, Mendeleev's periodic law became the first systamatic attempt to classify the elements. Mendeleev's periodic law : "Properties of elements are periodic function of their atomic masses". At that time when this law was proposed only about 60 elements were known, which he successively arranged in the order of increasing atomic mass to form a periodic table as given below :

Н			117	V	M	1/11			
1.01	П	111	IV	V	VI	VII			
Li	Ве	В	С	Ν	0	F			
6.94	9.01	10.8	12.0	14.0	16.0	19.0			
Na	Mg	Al	Si	Р	S	Cl		1/111	
23.0	24.3	27.0	28.1	31.0	32.1	35.5		VIII	
K	Са		Ti	V	Cr	Mn			
39.1	40.1		47.9	50.9	52.0	54.9	Fe	Со	Ni
Cu	Zn			As	Se	Br	55.9	58.9	58.7
63.5	65.4			74.9	79.0	79.9			
Rb	Sr	Y	Zr	Nb	Мо				
85.5	87.6	88.9	91.2	92.9	95.9		Ru	Rh	Pd
Ag	Cd	In	Sn	Sb	Те	I.	101	103	108
108	112	115	119	122	128	127			
Ce	Ва	La		Та	W				
133	137	139		181	184		Os	Ir	Pt
Au	Hg	Ti	Pb	Bi			194	192	195
197	201	204	207	209					
			Th		U				
			292		238				

Table-1

Mendeleev's Original Periodic Table

But this table certainly did not exhibit much periodicity. Mendeleev argued with remarkable insight that not all the elements had been discovered and hence he left blanks for the undiscovered elements. Truely such elements were discovered later on for example the blank below AI which he predicted as eka-aluminum was discovered as galium (Ga). However this table suffered from many demerits like the position of isotopes.

The logical bases of the periodic table became clear after the discovery of Atomic number by Henry Moseley in 1913 and with the development of electronic structure of atom. This brought about a revision of the periodic low where atomic mass was replaced by atomic number, which is the root of the periodic behaviour of elements.

Modern Periodic Law :

"The properties of elements are a periodic function of the atomic number." The modern periodic table (Long form) as shown below is based on this law.

Table-2

S-Block	Elements													p–Bloo	ck Elen	nents	
	\wedge																
1 IA		Y															18 VIII A
1 H 1.007	2 II A				d	-Block	c Eler	nents				13 III A	14 IV A	15 V A	16 VI A	17 VII A	2 He 4.002
3 Li 6.941	4 Be 9.012										\frown	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 VIII	9 VIII	10 VIII	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 V 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 Pd 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 TI 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
							Inne	r - Tr	ansiti	ion M	etals (f-Blocl	k elem	ents)			
	*La	nthan	ides	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
	*	*Actini	des	90 Th 232.038	91 Pa 231	92 U 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

Periodicity : The regular variation in the properties of elements with the increasing atomic number and repetition of the same trend after regular intervals of atomic number is called periodicity. This is due to recurrence of similar valence shell electronic configurations, for example all alkali metals have same valence shell electronic configuration (ns') and therefore have similar properties.

Important Feature of Long Form Periodic Table :

Total number of elements = 118

Name of four elements with atomic numbers : 113, 115, 117 and 118 which have been recently approved by IUPAC (Jan. 2016) are yet to be finalised, till then these are to be known by their atomic numbers as :

113	Uut	i.e.	ununtrium
115	Uup	i.e.	un-unpentium
117	Uus	i.e.	un-unseptium
118	Uuo	i.e.	un-unoctium.

Period : Horizontal rows are called periods. There are 7 periods and all are complete. The number of elements in different periods are as :

Period NumberTotal elements in that period

1st	2
2nd	8
3rd	8
4th	18
5th	18
6th	32

7th

32

- Each period starts with an alkali metal having outermost electronic configuration ns¹.
- Each period ends with a noble gas with outermost electronic configuration ns²p⁶ (except He = 1s²)
- Each period starts with the filling of a new energy level.
- An element belonging to nth period may have electrons in ns and np, subshells in its ground state but not in 'nd' and 'nf' subshells. It may have (n-1) d and (n-2) f electrons.

For example ground state electronic configuration of Hg (Z = 80) which belongs to 6th period is : $1s^2$, $2s^2p^6$, $3s^2p^6d^{10}$, $4s^2p^6d^{10}f^{14}$, $5s^2p^6d^{10}$, $6s^2$.

Groups (vertical columns) :

Blocks :

All the 118 elements can be considered to be further classified into 4 major blocks.

(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 those groups 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 metals.
- Total number of elements including Hydogen = 13.

(b) p-block elements

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

- Element from group 13 to 18 constitute the p-block.
- General electronic configuration is [inert gas] $ns^2 np^{1-6}$ (except He = $1s^2$)
- 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.
- Total number of elements = 37.

(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) I^{st} transition series i.e. 3d series which contains 10 elements and starts from ${}_{21}Sc {}_{30}Zn$. 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.
- Those elements which have partially filled d-orbitals in neutral state or in any stable oxidation state are called transition elements
- Total number of elements = 40

(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, contains 14 elements $_{58}$ Ce to $_{71}$ Lu. Filling of electrons takes place from 4f subshell.
- **2.** Ind inner transition or 5 f-series, contains 14 elements $_{90}$ Th to $_{103}$ Lr. 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.
- Total number of elements = 28

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.
- 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.

-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, l = 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.
- **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 afbaus 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 :

 $\begin{array}{lll} {\mathsf{A}}: 1{\mathsf{s}}^2\,2{\mathsf{s}}^2\,2{\mathsf{p}}^1 & {\mathsf{B}}: 1{\mathsf{s}}^2\,2{\mathsf{s}}^2\,2{\mathsf{p}}^6\,3{\mathsf{s}}^2\,3{\mathsf{p}}^1 \\ {\mathsf{C}}: 1{\mathsf{s}}^2\,2{\mathsf{s}}^2\,2{\mathsf{p}}^6\,3{\mathsf{s}}^2\,3{\mathsf{p}}^3 & {\mathsf{D}}: 1{\mathsf{s}}^2\,2{\mathsf{s}}^2\,2{\mathsf{p}}^6\,3{\mathsf{s}}^2\,3{\mathsf{p}}^5 \\ {\mathsf{E}}: 1{\mathsf{s}}^2\,2{\mathsf{s}}^2\,2{\mathsf{p}}^6\,3{\mathsf{s}}^2\,3{\mathsf{p}}^6\,4{\mathsf{s}}^2 \end{array}$

- 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. 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.
- Liquid elements at room-temperature are Ga(Z = 31), Br(Z = 35), Hg(Z = 80).

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.

Tools for predicting periodic trends.

(i) Nuclear Charge (Dominant factor in case of group) :

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.

(ii) 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.



(iii) 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 a 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.

(iv) 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) 1s², 2s² B (Z = 5) 1s², 2s², 2p¹

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

(v) Effective Nuclear Charge

Effective nuclear charge is the net positive charge experienced by electrons in a sub-shell. It can be approximated by the equation $Z_{eff} = Z - \sigma$, where Z is the atomic number and σ is a factor that depends on number of electrons which provide shielding.

The periodicity of atomic properties :

(i) 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,

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

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(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
К	231 pm	203 pm
Na	186 pm	154 pm

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

 $\rm r_{covalent} < \rm r_{crystal} < \rm r_{vander \, Walls}$

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	2. 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} . $r_n \propto \frac{1}{Z^*}$ Hence atomic radii decrease with increase in atomic number in a period from left to right.	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.
4. But number of orbitals (n) remains constant	4. But number of orbitals (n) increases.

Table-4

• The atomic radius of inert gas es (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.

(D) 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.

Table-5

Cation	Anion
(1) It is formed by the loss of one or more electrons	(1) It is formed by the gain of one or more electrons from
from the valence shell of an atom of an element.	in the valence shell of an atom of an element.
(2) Cations are smaller than the parent atoms	(2) Anions are larger than the parent atoms because
because,	(a) anion is formed by gain of one or more electrons in
(a) the whole of the outer shell of electrons is usually	the neutral atom and thus number of electrons increases
removed.	but magnitude of nuclear charge remains the same.
(b) in a cation, the number of positive charges on the	(b) nuclear charge per electrons is thus reduced and the
nucleus in greater than number of orbital electrons	electrons cloud is held less tightly by the nucleus leading
leading to incresed inward pull of remaining electrons	to the expansion of the outer shell. Thus size of anion is
causing contraction in size of the ion.	increased.

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

- 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.

- (i) S²⁻, Cl⁻, K⁺ , Ca⁺² , Sc⁺³ (ii) SO₂, NO₃⁻ , CO₃²⁻, (iii) N₂ , CO, CN⁻ (iv) NH₃ , H₃O⁺
- Pauling's empirical formula for ionic radius $\propto \frac{1}{\text{nuclear charge}}$ (only for isoelectronic species).

Solved Examples

- **Ex.6** Consider the following species :
 - $N^{\scriptscriptstyle 3-}$, $O^{\scriptscriptstyle 2-}$, $F^{\scriptscriptstyle -}$, $Na^{\scriptscriptstyle +}$, $Mg^{\scriptscriptstyle 2+}$, and $AI^{\scriptscriptstyle 3+}$
 - (a) What is common in them?
 - (b) Arrange them in the order of increasing ionic radii.
- *Sol.* (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[−], Na²⁺ and Al³⁺. 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⁻.

(ii) Ionisation Energy :

lonisation 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.

 $\begin{array}{c} \mathsf{M}(g) \xrightarrow{(\mathsf{IE}_1)} \mathsf{M}^+(g) + e^- ; & \mathsf{M}^+(g) + \mathsf{IE}_2 \longrightarrow \mathsf{M}^{2+}(g) + e^- \\ \mathsf{M}^{2+}(g) + \mathsf{IE}_2 \longrightarrow \mathsf{M}^{+3}(g) + e^- \end{array}$

 $\mathsf{M}^{_{2+}}(\mathsf{g}) + \mathsf{IE}_{_3} \longrightarrow \mathsf{M}^{_{+3}}(\mathsf{g}) + \mathsf{e}^{_-}$

 IE_1 , $IE_2 \& IE_3$ are the I^{st} , $II^{nd} \& III^{rd}$ 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.

 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.

- (i) Size of the Atom (Dominant factor in case of group) :
- (ii) Effective Nuclear charge
- (iii) Shielding and screening effect
- (iv) Electronic configuration
- 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.
- 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 = $-2.18 \times 10^{-15} \text{ J}$

Energy of the electron at infinite = 0.

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

- $= 0 (its energy in the ground state) = -(-2.18 \times 10^{-18} \text{ J}) = 2.18 \times 10^{18} \text{ J}.$
- $\therefore \qquad \text{Ionization enthalpy per mole of hydrogen atoms} = \frac{2.18 \times 10^{-18} \times 6.02 \times 10^{23}}{1000} \text{ kJ}$

= 1312.36 kJ mol⁻¹ = 1312.36 x 10³ 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 (1s² 2s²) the outermost electron is present in 2s-orbital while in B (1s² 2s² 2p¹) 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_iH of Be is higher than that Δ_iH 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 –



- (1) above Ne(3) below Li
- **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.

(iii) Electron Affinity :

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.

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

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 electron is added to the atom and the energy is released, the electron gain enthalpy is negative and when energy is needed to add an electron to the atom, the electron gain enthalpy is positive.

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.

 $E_a = E(X, g) - E(X^-, g),$

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

 $F + e^- \longrightarrow F^ E_a = 328 \text{ kJ mol}^{-1}$; $\Delta_{eq}H^{\Theta} = -328 \text{ kJ mol}^{-1}$

A positive electron gain enthalpy 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.

 $O(g) + e^{-}(g) \xrightarrow{\text{Exothermic}} O^{-}(g) ; \qquad E_a = 140.9 \text{ kJ mol}^{-1} ; \Delta_{eg}H^{\odot} = -140.9 \text{ kJ mol}^{-1}$

 $O^{-}(g) + e^{-}(g) \xrightarrow{\text{Endothermic}} O^{2-}(g) ; \qquad \mathsf{E}_{\mathsf{a}} = -744 \text{ kJ mol}^{-1} ; \Delta_{\mathsf{eg}} \mathsf{H}^{\Theta} = 744 \text{ kJ mol}^{-1}$

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

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

- As we move in a group from top to bottom, electron gain enthalpy becomes less negative because the size of the atom increases and the added electron would be at larger distance from the nucleus.
- 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.
- 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.
- 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. $-\Delta_{ea}H^{o}$)?
- **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 (A) 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.

(iv) 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.

- 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.
- The electronegativity also increases as the s-character in the hybrid orbitals increases.
 Hybrid orbital sp³ sp² sp
 s-character 25% 33% 50%

Electronegativity	increases
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l able-6	
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 negativity diff.

$$|X_{A} - X_{B}| = 0.102 \sqrt{\Delta E}$$

where $\Delta E = E_{A-B} - \frac{1}{2} \left[E_{A-A} + E_{B-B} \right]$
All E (bond energy)

are is kJ mol-1

(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).

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

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.

- -

- **Ex.15** Calculate the electronegativity of carbon from the following data : $E_{H-H} = 104.2 \text{ kcal mol}^{-1}$, $E_{C-C} = 83.1 \text{ kcal mol}^{-1}$ $E_{C-H} = 98.8 \text{ kcal mol}^{-1}$, $X_H = 2.1$
- *Sol.* Let the electronegativity of carbon be X_C, applying Pauling equation,

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

 $x_{C}^{-} - 2.1 = 0.208 [98.8 - \frac{1}{2} (83.1 + 104.2)]^{1/2}$ ⇒ $x_{C}^{-} = 2.5$ 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.

Sol. $EN_v - EN_v = 3.2 - 2.2 = 1.$

 $\Delta = 1$

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

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

Ans. (1)

(V) Periodicity of Valence or 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 $3s^1$ 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.

Table-7

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 N +5 +4 +3 +1	16 8 0 +2 -1/2 -1 -1 -1	17 9 F –1	10 Ne
11 Na +1	12 Mg +2											13 Al +3	14 Si +4 -4	-3 0 etc. 15 P +5 +3 +1	16 S +6 +4 +2	17 CI +5 +7 +3	18 Ar 0
19 K +1	20 Ca +2	3 21 Sc +2 +3	4 22 Ti +2 +3 +4	5 23 V +2 +3 +4 +5	6 24 Cr +1 +2 +3 +4 +5	7 25 Mn +1 +2 +3 +4 +5	8 26 Fe +1 +2 +3 +4 +5	9 27 Co +1 +2 +3 +4 +5	10 28 Ni +1 +2 +3 +4	11 29 Cu +1 +2	12 30 Zn +2	31 Ga +3	32 Ge +4 -4	-3 33 As +3 -3	-2 34 Se +4 -2	+1 -1 etc. 35 Br +5 +3 +1 -1	36 Kr +4 +2 0
37 Rb +1	38 Sr +2				+6	+6 +7	+6					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

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 (1) The metallic character of the elements is highest at the extremely left (low ionisation energies) and Sol. 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 > BSol. (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 ? (3) PH₃, SiH₄, HS⁻, Ar (1) CH₄, H₂O, NH₃, HF (4) H₂S, K⁺, Ar, Cl (3) OH⁻, H₂O, NH₂⁻, F⁻ (4) As all species do not have same number of electrons i.e., 18. H₂S, K⁺ and Ar have same number of Sol. electrons i.e. 18 but CI 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. [Ar]¹⁸ 4s⁰ 3d⁴ Electron configuration of M²⁺ is : [Ar]¹⁸ 4s¹ 3d⁵ (and not 4s² 3d⁴) *.*.. Electron configuration of M is : So total number of electrons = 24. Hence, metal M is chromium (Cr). 5. Following are the valence shell electronic configurations of some elements. (i) 3s² 3p⁵ (ii) 3d¹⁰ 4s² (iii) 2s² 3p⁶ 4s¹ (iv) 1s² 2s² Find 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. As last electron enters in d-subshell, therefore this belongs to d-block. For d-block element the group Ans. 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- ? Be²⁺ < Mg²⁺ < Na⁺ < Cl⁻ < S²⁻ < Br⁻ Ans. Be²⁺ is smaller than Mg²⁺ as Be²⁺ has one shell where as Mg²⁺ has two shells. Sol. Mg²⁺ and Na⁺ are isoelectronic species : Ionic radius \propto 1/nuclear charge. Cl⁻ and S²⁻ are isoelectronic species : Ionic radius \propto 1/nuclear charge. CI⁻ is smaller than Br⁻ as CI⁻ has three shells where as Br⁻ has four shells. The (IE₁) and the (IE₂) in kJ mol⁻¹ of a few elements designated by Roman numerals are shown below: 8. Ш Т Ш 403 549 IE₁ 1142 2640 1060 2080 IE, Which of the above elements is likely to be a (a) non-metal (b) alkali metal

- (c) alkaline earth metal ?
- Ans. (a) non-metal(III) – Due to highest ionisation energy, (IE_1) and (IE_2) .
 - (b) alkali metal (I) Due to lowest ionisation energy, (IE,) and there is guite high jump in (IE,) due to inert gas configuration.

(c) alkaline earth metal (II) – There is little difference in (IE,) and (IE,) and the value of (IE,) is slightly greater than(I) due to stable configuration(ns²).

- lonisation energy and electron affinity of fluorine are respectively 17.42 and 3.45 eV. Calculate 9. electronegativity of fluorine atom.
- According to Mulliken's electronegativity $(\chi_M) = \frac{\text{Ionisation energy} + \text{Electron affinity}}{2}$ (in eV) Sol. \mathbf{c}

$$=\frac{17.42+3.45}{2}=10.435$$

Therefore, electronegativity on Pauling's scale $(\chi_P) = \frac{10.435}{2.8} = 3.726$

Ans. $\chi_{\rm P} = 3.726$

- 10. Are the oxidation state and covalency of AI in [AICI (H₂O)₂]²⁺ same ?
- Sol. No. The oxidation state of AI is +3 and the covalency is 6.
- 11. Why the electron gain enthalpy values of alkaline earth metals are lower (i.e. less negative) or positive ?
- Sol. The general valence shell electron configuration of alkaline earth metals is ns² (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 given electronic configurations of some elements. Match these with the correct metals given in Column-II: Column-II

С	ol	u	m	r)-l

(1)	ns², np⁵	(p)	Chromium
(2)	(n – 1) d ¹⁰ , ns ¹	(q)	Copper
(3)	(n – 1) d⁵ , ns¹	(r)	Krypton
(4)	(n – 1) d ¹⁰ , ns ² , np ⁶	(s)	Bromine
(1)	(a) + (2) + (a) + (2) + (a) + (4) + (a)		

- Ans. $(1) \rightarrow (s)$; $(2) \rightarrow (q)$; $(3) \rightarrow (p)$; $(4) \rightarrow (r)$.
- (1) ns²np⁵ is general valence shell electron configuration of halogens. So this configuration belongs to Sol. bromine.

(2) $(n - 1) d^{1-10} ns^{1-2}$; This is electron configuration of d-block elements. As it contains $(n - 1) d^{10}ns^{1}$ configuration it belongs to copper.

(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 chromium.

(4) Noble gases has valence shell electron configuration ns² np⁶, so it belongs to krypton.

Match the metals given in Column-II with their type given in Column-I: 13.

Match the metals given in Column in with their type given in Column 1.			
	Column-I		Column-II
(1)	Metalloid	(p)	Sulphur
(2)	Radioactive	(q)	Gold
(3)	Transition metal	(r)	Arsenic
(4)	Chalcogen	(s)	Uranium

- Ans. $(1) \rightarrow (r)$; $(2) \rightarrow (s)$; $(3) \rightarrow (q)$; $(4) \rightarrow (p)$
- (1) Arsenic is a metalloid because it behaves as metal (forming cation, As3+ -AsCl₂) as well as nonmetal Sol. (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, 5d⁸6s⁰.

(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:

	0	<i>J</i> 1 0	
	Column-I		Column-II
(1)	Representative element	(p)	Cerium
(2)	Lanthanide	(q)	Aluminium
(3)	Coinage metal	(r)	Thorium

(4) Actinide

- (s) Gold
- **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.

Col	um	n-I
001	um	11-1

(1)

Ans.

- Isoelectronic species(p) $A^+(g)$ + energy $\rightarrow A^{++}(g)$ + $e^-(g)$ Half filled orbital(q)Ar, K⁺, Ca⁺⁺
- (2) Half filled orbital
 (3) Second ionisation energy
 (7) Lutetium
- (4) Inner transition element (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-I		Column-II
(1)	Highest 1 st ionisation energy	(p)	Technitium
(2)	Highest electronegativity	(q)	Lithium
(3)	Synthetic element	(r)	Helium
(4)	Strongest reducing agent	(s)	Fluorine
(1) →	(\mathbf{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^0 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) 8 th group
(2) An element whose valence shell contains one unpaired p-electron	(q) 12 th group
(3) An element which receives last electron in $(n - 1)$ d-subshell	(r) 14 th group
(4) An element with the ground-state electron configuration $[Ar]4s^23d^{10}$	(s) 17 th group

 $\textbf{Ans.} \quad (1) \rightarrow (r) \ ; \ (2) \rightarrow (s) \ ; \ (3) \rightarrow (p, \, q) \ ; \ (4) \rightarrow (q).$

Sol. (1) [Ar] $3d^{10}4s^{2}4p^{2}$: 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 ns^2np^5 and there is one unpaired electron in p-subshell i.e., 1/1/1/1

(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.