

PERIODIC TABLE & PERIODICITY

1. 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 systematic 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 :

Table-1

I	II	III	IV	V	VI	VII			
H 1.01									
Li 6.94	Be 9.01	B 10.8	C 12.0	N 14.0	O 16.0	F 19.0			
Na 23.0	Mg 24.3	Al 27.0	Si 28.1	P 31.0	S 32.1	Cl 35.5	VIII		
K 39.1	Ca 40.1		Ti 47.9	V 50.9	Cr 52.0	Mn 54.9	Fe 55.9	Co 58.9	Ni 58.7
Cu 63.5	Zn 65.4			As 74.9	Se 79.0	Br 79.9			
Rb 85.5	Sr 87.6	Y 88.9	Zr 91.2	Nb 92.9	Mo 95.9		Ru 101	Rh 103	Pd 108
Ag 108	Cd 112	In 115	Sn 119	Sb 122	Te 128	I 127			
Ce 133	Ba 137	La 139		Ta 181	W 184		Os 194	Ir 192	Pt 195
Au 197	Hg 201	Ti 204	Pb 207	Bi 209					
			Th 232		U 238				

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. Truly such elements were discovered later on for example the blank below Al 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 law where atomic mass was replaced by atomic number, which is the root of the periodic behaviour of elements.

2. 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		d-Block Elements										p-Block Elements					
1 IA H 1.007	2 II A He 4.002											13 III A	14 IV A	15 V A	16 VI A	17 VII A	18 VIII A
3 Li 6.941	4 Be 9.012	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	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 58.933	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	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	
87 Fr 223	88 Ra 226	89 Ac** 227	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	
Inner - Transition Metals (f-Block elements)																	
		*Lanthanides															
		**Actinides															

2.1 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. ununpentium
- 117 Uus i.e. ununseptium
- 118 Uuo i.e. ununoctium.

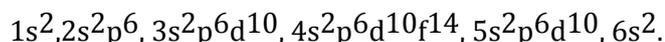


Period : Horizontal rows are called periods. There are 7 periods and all are complete. The number of elements in different periods are as :

Period Number Total 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 ns^1 .
- Each period ends with a noble gas with outermost electronic configuration ns^2p^6 (except $He = 1s^2$)
- Each period starts with the filling of a new energy level.
- An element belonging to n^{th} 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 :



2.2 Groups (vertical columns) :

There are 18 groups or vertical columns in this periodic table, which are numbered from left to right as 1, 2, 3, 4, 5,..... 13, 14, 15, 16, 17, 18. Previously vertical columns 1, 2, 13, 14, 15, 16, 17 & 18 were called IA, IIA, IIIA, IVA, VA, VIA, VIIA and Zero group respectively. And vertical columns 3, 4, 5, 6, 7, 11 and 12 were known as III B, IVB, VB, VIB, VIIB, VIII, IB and IIB respectively. Columns 8, 9, 10 collectively constituted VIII group. Lanthanides ($Z = 57$ to $Z = 71$) and Actinides ($Z = 89$ to $Z = 103$) belong to 3rd vertical column or IIIB subgroup. However these days groups are more known as vertical columns numbered from 1 to 18.

3. BLOCKS :

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

3.1 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] ns^{1-2}
- ◆ s -block elements lie on the extreme left of the periodic table.
- ◆ This block includes metals.
- ◆ Total number of elements including Hydrogen = 13.

3.2 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 n^{th} 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.

3.3 d-Block elements

The element in which outermost (n^{th}) and penultimate shells $(n - 1)^{\text{th}}$ 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] $(n - 1) d^{1-10} ns^{1-2}$ (except, palladium which has valence shell electron configuration $4d^{10} 5s^0$).
- ◆ 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 $_{21}\text{Sc} - _{30}\text{Zn}$. Filling of electrons takes place in 3d sub-shell.
 2. IInd transition series i.e. 4d series which contains 10 elements and starts from $_{39}\text{Y} - _{48}\text{Cd}$. Filling of electrons takes place in 4d sub-shell.
 3. IIIrd transition series i.e. 5d series which contains 10 elements and starts from $_{57}\text{La}, _{72}\text{Hf} - _{80}\text{Hg}$. Filling of electrons takes place in 5d sub-shell.
 4. IVth transition series i.e. 6d series which contains 10 elements and starts from $_{89}\text{Ac}, _{104}\text{Rf} - _{112}\text{Uub}$. 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

3.4 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., $(n - 2)^{\text{th}}$ 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. 1st inner transition or 4 f-series, contains 14 elements $_{58}\text{Ce}$ to $_{71}\text{Lu}$. Filling of electrons takes place from 4f subshell.
 2. 2nd inner transition or 5 f-series, contains 14 elements $_{90}\text{Th}$ to $_{103}\text{Lr}$. 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

3.5 Prediction of period, group and block :

- Period of an element corresponds to the principal quantum number of the valence shell.
- 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
 - (a) For d-block elements,
Group number = number of electrons in $(n - 1)$ d sub shell + number of electrons in the valence shell.

Solved Examples

Example 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.

Example 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.

Example 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.

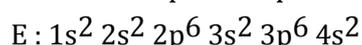
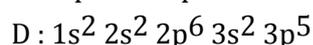
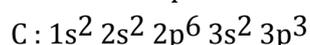
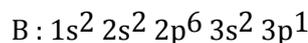
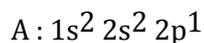
Example 4. In terms of period and group where would you locate the element with $Z = 114$?

Sol. The filling of the 6th period ends at $_{86}\text{Rn}$. 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 5f occurs. But according

to Aufbau principle their energies increase in the order : $7s < 5f < 6d < 7p$. Therefore, after ${}_{86}\text{Rn}$, the next two elements with $Z = 87$ and $Z = 88$ are s-block elements, the next fourteen, i.e., $Z = 89 - 102$ are f-block elements, the next ten i.e., $Z = 103 - 112$ 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 periodic table is

Period = 7th **Block** : p- Block **Group** : 14

Example 5. Elements A, B, C, D and E have the following electronic configurations :



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, $ns^2 np^1$.

4. METALS AND NONMETALS :

- ◆ The metals are characterised by their (i) shiny appearance (ii) nature of readily losing the electrons. Metals comprise 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. CrO_3).
- ◆ 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 ($Z = 80$).

4.1 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.

4.2 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 ($\square A$) 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 many respects belonging to the same group due to the smaller atomic size and absence of vacant d-orbitals.

4.3 Tools for predicting periodic trends.

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

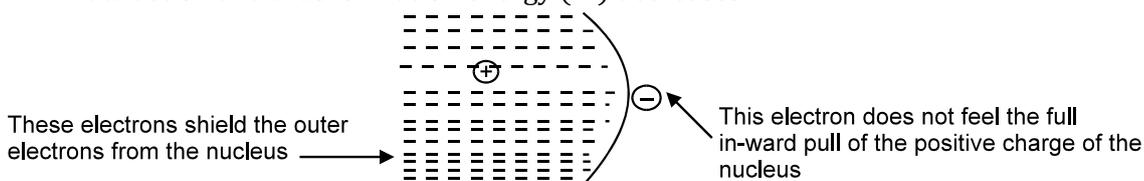
The ionisation energy increases with increase in the nuclear charge.

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.

4.3.2 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.



4.3.3 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 whereas in magnesium it will be removed from 3s-orbital (having larger penetration effect) of same energy level.

4.3.4 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.

4.3.5 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 $Z_{\text{eff}} = Z - \sigma$, where Z is the atomic number and σ is the shielding factor that depends on number of electrons which provide shielding.

5. THE PERIODICITY OF ATOMIC PROPERTIES :

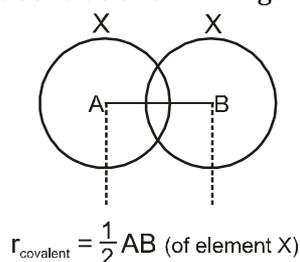
5.1 Atomic radius :

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.

5.1.1 Atomic radius can be

Covalent radius :

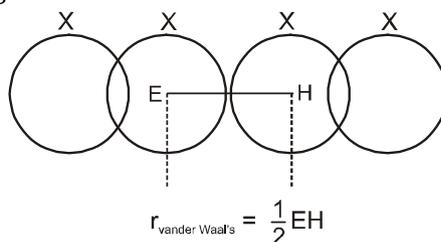
It is one-half of the distance between the centres of two nuclei (of like atoms) bonded by a single covalent bond as shown in figure.



- ◆ Covalent radius is generally used for nonmetals.

5.1.2 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

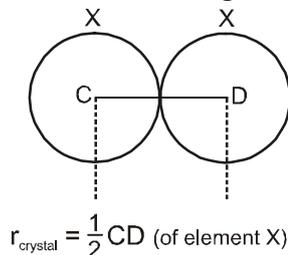
- The Vander Waal's force of attraction is weak, therefore, the internuclear distance in case of atoms held by Van der Waal's force is much larger than those of held covalent bond. Therefore Vander Waal's radii are always larger than covalent radii.
- A covalent bond is formed by the overlapping of two 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,

Table-3

Elements	H	O	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

5.1.3 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.

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

Table-4

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.

- ♦ 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.

5.2 Ionic radius :

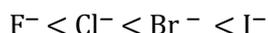
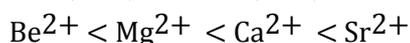
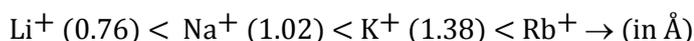
The effective distance from the centre of nucleus of the ion up to which it has an influence in the ionic bond is called ionic radius.

Table-5

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

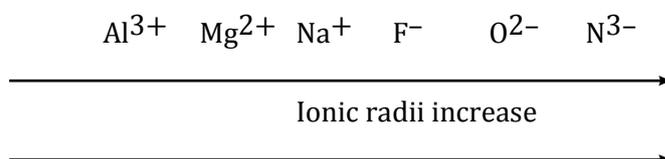
- ◆ The sizes of ions increases as we go down a group (considering the ions of same charge).

For example :



- ◆ For the cations same element the ionic radius decreases as more electrons are lost (i.e. the valency increases) For e.g. $\text{Cr}^{2+} = 0.80 \text{ \AA}$, $\text{Cr}^{3+} = 0.615 \text{ \AA}$, $\text{Cr}^{4+} = 0.55 \text{ \AA}$, $\text{Cr}^{5+} = 0.49 \text{ \AA}$ and $\text{Cr}^{6+} = 0.44 \text{ \AA}$.
- ◆ 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 ($\text{Hf}^{4+} = 0.71 \text{ \AA}$) and Zr ($\text{Zr}^{4+} = 0.72 \text{ \AA}$), and Nb ($\text{Nb}^{3+} = 0.72 \text{ \AA}$) and Ta ($\text{Ta}^{3+} = 0.72 \text{ \AA}$) 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, N^{3-} , O^{2-} , F^- , Ne, Na^+ , Mg^{2+} and Al^{3+} 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.



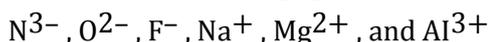


As effective nuclear charge decrease.

- ◆ Following are the examples of isoelectronic series
(i) S^{2-} , Cl^- , 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 $\propto \frac{1}{\text{nuclear charge}}$ (only for isoelectronic species).

Solved Examples

Example 1. Consider the following species :



- (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^- , Na^{2+} and Al^{3+} . All these ions 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 Al^{3+} (+ 13). Therefore, their ionic radii decrease in the order : $N^{3-} > O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$.

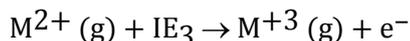
Example 2. Mg^{2+} is smaller than O^{2-} in size, though both have same electronic configuration. Explain ?

Sol.

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

5.3 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.



IE_1 , IE_2 & IE_3 are the 1st, 2nd & 3rd 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.
- ◆ 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

Example 1. 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^{-18} J

Energy of the electron at infinite = 0.

The energy required to remove an electron in the ground state of hydrogen atom = $0 - (\text{its energy in the ground state}) = -(-2.18 \times 10^{-18} \text{ J}) = 2.18 \times 10^{-18}$ J.

\therefore Ionization enthalpy per mole of hydrogen atoms = $\frac{2.18 \times 10^{-18} \times 6.02 \times 10^{23}}{1000}$ kJ

= $1312.36 \text{ kJ mol}^{-1} = 1312.36 \times 10^3 \text{ J mol}^{-1}$.

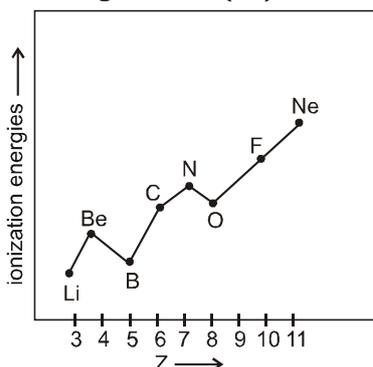
Example 2. Among the second period elements the actual ionization enthalpies are in the order $\text{Li} < \text{B} < \text{Be} < \text{C} < \text{O} < \text{N} < \text{F} < \text{Ne}$. Explain why :

- (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, therefore, lesser amount of energy is required to knock out a 2p-electron than a 2s-electron. Consequently, $\Delta_i H$ of Be is higher than that $\Delta_i H$ of B.

- (ii) The electronic configuration 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 ($1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$) 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, $\Delta_i H$ 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 outweighs the effect of stability due to exactly half-filled orbitals, therefore, the $\Delta_i H$ of N and O are lower than that of F.

Example 3. Following graph shows variation of ionization energies with atomic number in second period (Li-Ne). Value of ionization energies of Na(11) will be –



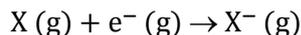
- (1) above Ne
 (2) below Ne but above O.
 (3) below Li
 (4) between N and O.

Sol.

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

5.4 Electron Affinity :

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



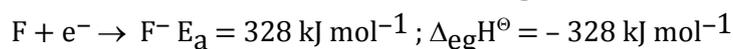
Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form anion. Electron gain may be either exothermic or endothermic depending on the elements.

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.

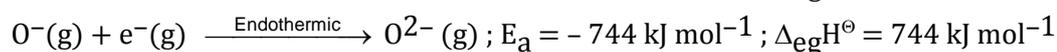
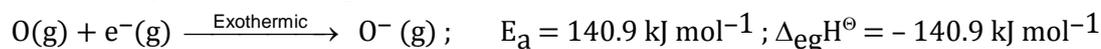
Although the electron gain enthalpy is the thermodynamically appropriate term, much of inorganic chemistry is discussed in terms of a closely related property, the electron affinity, E_a of an element which is the difference in energy between the gaseous atoms.

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

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



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.



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.
- 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.
- (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

Example 1. 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.

Example 2. Why do halogens have high electron gain enthalpies (i.e. $-\Delta_{\text{eg}}H^\ominus$) ?

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

Example 3. 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 O^{2-} . So option (A) is correct.



Example 4. 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 $[\text{He}]2s^1$ and $[\text{He}]2s^2$, respectively. The additional electron enters the 2s orbital of Li but the 2p orbital of Be and hence is much less tightly bound. In fact, the nuclear charge is so well shielded in Be that electron gain is endothermic.

5.5 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.
- 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 \square electronegativity of the atom.

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

Hybrid orbital	sp^3	sp^2	sp
s-character	25%	33%	50%

$\xrightarrow{\hspace{10em}}$
 Electronegativity increases

Table-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.

5.5.1 Pauling's scale :

Linus Pauling developed a method for calculating relative electronegativities of most elements. According to Pauling negativity diff.

$$|X_A - X_B| = 0.102 \sqrt{\Delta E}$$

$$\text{where } \Delta E = E_{A-B} - \frac{1}{2}[E_{A-A} + E_{B-B}]$$

All E (bond energy)

are in kJ mol^{-1}

5.5.2 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).

$$\chi_M = \frac{IE + EA}{2}$$

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

$$\chi_P = 1.35 (\chi_M)^{1/2} - 1.37$$

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

Solved Examples

Example 1. Calculate the electronegativity of carbon from the following data :

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

$$E_{\text{C-H}} = 98.8 \text{ kcal mol}^{-1}, \quad X_{\text{H}} = 2.1$$

Sol. Let the electronegativity of carbon be X_{C} , applying Pauling equation,

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

$$X_{\text{C}} - 2.1 = 0.208 \left[98.8 - \frac{1}{2} (83.1 + 104.2) \right]^{1/2}$$

$$\Rightarrow X_{\text{C}} = 2.5$$

Ans. 2.5

Example 2. 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. $EN_x - EN_y = 3.2 - 2.2 = 1.$

$$\Delta = 1$$

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

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

Ans. (1)

5.6 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 $F > O > \text{Na}$. Each of the atoms of fluorine, with outer electronic configuration $2s^2 2p^5$, 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 $2s^2 2p^4$ 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 state of an element in Periodic Table

1 H +1 -1																	18 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 -3 0 etc.	16 8 O +2 -1 -2	17 9 F -1	10 Ne
11 Na +1	12 Mg +2											13 13 Al +3	14 14 Si +4 -4	15 15 P +5 +3 +1 -3	16 16 S +6 +4 +2 -2	17 17 Cl +5 +7 +3 +1 -1 etc.	18 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 +6	7 25 Mn +1 +2 +3 +4 +5 +6 +7	8 26 Fe +1 +2 +3 +4 +5 +6	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 31 Ga +3	32 32 Ge +4 -4	33 33 As +5 +3 -3	34 34 Se +6 +4 -2	35 35 Br +5 +3 +1 -1	36 36 Kr +4 +2 0
37 Rb +1	38 Sr +2											49 49 In +3 +1	50 50 Sn +4 +2	51 51 Sb +5 +3 -3	52 52 Te +6 +4 -2	53 53 I +7 +5 +1 0 -1	54 54 Xe +6 +4 +2 0
55 Cs +1	56 Ba +2											81 81 Tl +3 +1	82 82 Pb +4 +2	83 83 Bi +5 +3	84 84 Po	85 85 At	86 86 Rn

MISCELLANEOUS SOLVED PROBLEMS

Problem 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).

Problem 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).

Problem 3 Which of the following group does not represent the isoelectronic species ?

- (1) CH₄, H₂O, NH₃, HF (3) PH₃, SiH₄, HS⁻, Ar
 (2) OH⁻, H₂O, NH₂⁻, F⁻ (4) H₂S, K⁺, Ar, Cl

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.

Problem 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 : [Ar]¹⁸ 4s⁰ 3d⁴
 ∴ Electron configuration of M is : [Ar]¹⁸ 4s¹ 3d⁵ (and not 4s² 3d⁴)
 So total number of electrons = 24.
 Hence, metal M is chromium (Cr).

Problem 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.

Problem 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.



Problem 7 Arrange the following ions in the increasing order of their size : Be^{2+} , Cl^- , S^{2-} , Na^+ , Mg^{2+} , Br^- ?

Ans. $\text{Be}^{2+} < \text{Mg}^{2+} < \text{Na}^+ < \text{Cl}^- < \text{S}^{2-} < \text{Br}^-$

Sol. 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 : Ionic radius $\propto 1/\text{nuclear charge}$.
 Cl^- and S^{2-} are isoelectronic species : Ionic radius $\propto 1/\text{nuclear charge}$.
 Cl^- is smaller than Br^- as Cl^- has three shells where as Br^- has four shells.

Problem 8 The (IE_1) and the (IE_2) in kJ mol^{-1} of a few elements designated by Roman numerals are shown below:

	I	II	III
IE_1	403	549	1142
IE_2	2640	1060	2080

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_1) and there is quite high jump in (IE_2) due to inert gas configuration.
 (c) alkaline earth metal (II) – There is little difference in (IE_1) and (IE_2) and the value of (IE_1) is slightly greater than(I) due to stable configuration(ns^2).

Problem 9 Ionisation energy and electron affinity of fluorine are respectively 17.42 and 3.45 eV. Calculate electronegativity of fluorine atom.

Sol. According to Mulliken's electronegativity (χ_M) = $\frac{\text{Ionisation energy} + \text{Electron affinity}}{2}$ (in eV)

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

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

Ans. $\chi_P = 3.726$

Problem 10 Are the oxidation state and covalency of Al in $[\text{AlCl}(\text{H}_2\text{O})_5]^{2+}$ same ?

Sol. No. The oxidation state of Al is +3 and the covalency is 6.

Problem 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 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.

Problem 12 In Column-I, there are given electronic configurations of some elements. Match these with the correct metals given in Column-II :

Column-I	Column-II
(1) ns^2, np^5	(p) Chromium
(2) $(n-1)d^{10}, ns^1$	(q) Copper
(3) $(n-1)d^5, ns^1$	(r) Krypton
(4) $(n-1)d^{10}, ns^2, np^6$	(s) Bromine

Ans. (1) → (s) ; (2) → (q) ; (3) → (p) ; (4) → (r).

Sol. (1) ns^2np^5 is general valence shell electron configuration of halogens. So this configuration belongs to bromine.
 (2) $(n-1)d^{10}ns^1$; This is electron configuration of d-block elements. As it contains $(n-1)d^{10}ns^1$ configuration it belongs to copper.
 (3) $(n-1)d^5ns^1$; This is electron configuration of d-block elements. As it contains $(n-1)d^5ns^1$ configuration it belongs to chromium.
 (4) Noble gases has valence shell electron configuration ns^2np^6 , so it belongs to krypton.

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

Column-I	Column-II
(1) Metalloid	(p) Sulphur
(2) Radioactive	(q) Gold
(3) Transition metal	(r) Arsenic
(4) Chalcogen	(s) Uranium

Ans. (1) → (r) ; (2) → (s) ; (3) → (q) ; (4) → (p)

Sol. (1) Arsenic is a metalloid because it behaves as metal (forming cation, $As^{3+} - AsCl_3$) as well as nonmetal (forming anion, $As^{3-} - AsH_3$).
 (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.

Problem 14 Match the metals given in Column-II with their type given in Column-I :

Column-I	Column-II
(1) Representative element	(p) Cerium
(2) Lanthanide	(q) Aluminium
(3) Coinage metal	(r) Thorium
(4) Actinide	(s) Gold

Ans. (1) → (q) ; (2) → (p) ; (3) → (s) ; (4) → (r)

Sol. (1) s-block and p-block elements are collectively called as representative elements. As in aluminium last electron enters in p-subshell ($[Ne]^{10}3s^23p^1$).
 (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.

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

Column-I	Column-II
(1) Isoelectronic species	(p) $A^+(g) + \text{energy} \rightleftharpoons A^{++}(g) + e^-(g)$
(2) Half filled orbital	(q) Ar, K^+ , Ca^{++}
(3) Second ionisation energy	(r) 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) np^3 , $(n-1) d^5$ and $(n-2) f^7$ represent half filled orbitals. Antimony has $([Kr]^{36} 4d^{10} 5s^2 5p^3)$.
- (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.

Problem 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

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 ns^2 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^\ominus due to its higher hydration energy on account of its small size of atom.

Problem 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 (element / elements)	Column-II (group number)
(1) An element whose fourth shell contains two p-electrons	(p) 8 th group
(2) An element whose valence shell	(q) 12 th group

contains one unpaired p-electron

- (3) An element which receives last electron in $(n - 1)$ d-subshell (r) 14th group
- (4) An element with the ground-state electron configuration $[\text{Ar}]4s^23d^{10}$ (s) 17th group

Ans.

(1) \rightarrow (r) ; (2) \rightarrow (s) ; (3) \rightarrow (p, q) ; (4) \rightarrow (q).

Sol.

(1) $[\text{Ar}]3d^{10}4s^24p^2$: Fourth shell contains two electron in 4p-sub shell i.e., $4p^2$. 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., $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$

(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. $[\text{Ar}]4s^23d^{10}$ the group number = $2 + 10 = 12$.

