

## Development of Modern Periodic Table :

### (a) Dobereiner's Triads :

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

Li	Na	K	
7	23	39	$7+39 / 2 = 23$
Fe	Co	Ni	
55.85	58.93	58.71	nearly same atomic masses

It was restricted to few elements, therefore discarded.

### (b) Newland's Law of Octave :

He was the first to correlate the chemical properties of the elements with their atomic masses.

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

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

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca					

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

### (c) Lothar Meyer's Classification :

He determined the atomic volumes by dividing atomic masses with their densities in solid states.

He plotted a graph between atomic masses against their respective atomic volumes for a number of elements. He found the following observations.

- Elements with similar properties occupied similar positions on the curve.
- Alkali metals having larger atomic volumes occupied the crests.
- Transition elements occupied the troughs.
- The halogens occupied the ascending portions of the curve before the inert gases.
- Alkaline earth metals occupied the positions at about the mid points of the descending portions of the curve.

On the basis of these observations he concluded that the atomic volumes (a physical property) of the elements are the periodic functions of their atomic masses.

It was discarded as it lacks practical utility.

### (d) Mendeleev's Periodic Table :

#### Mendeleev's Periodic's Law

According to him the physical and chemical properties of the elements are the periodic functions of their atomic masses.

He arranged then known elements in order of their increasing atomic masses considering the facts that elements with similar properties should fall in the same vertical columns and leaving out blank spaces where necessary.

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

**Table-1**

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

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

### Merits of Mendeleev's Periodic table:

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

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

**Table-2**

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

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

### Demerits in Mendeleev's Periodic Table :

- Position of hydrogen is uncertain .It has been placed in IA and VIIA groups because of its resemblance with both the groups.
- No separate positions were given to isotopes.
- Anomalous positions of lanthanides and actinides in periodic table.
- Order of increasing atomic weights is not strictly followed in the arrangement of elements in the periodic table. For example Ar(39.94) is placed before K(39.08) and Te (127.6) is placed before I (126.9).
- Similar elements were placed in different groups e.g. Cu in IB and Hg in IIB and similarly the elements with different properties were placed in same groups e.g. alkali metals in IA and coinage metals in IB.
- It didn't explained the cause of periodicity.

### (e) Long form of the Periodic Table or Moseley's Periodic Table :

He studied (1909) the frequency of the X-ray produced by the bombardment of a strong beam of electrons on metal target. He found that the square root of the frequency of X-rays ( $\sqrt{\nu}$ ) is directly proportional to number of effective nuclear charge (Z) of metal i.e. to atomic number and not to atomic mass of the atom of that metal (as nuclear charge of metal atom is equal to atomic number), i.e.  $\sqrt{\nu} = a(Z - b)$ .

Where 'a' is the proportionality constant and 'b' is a constant for all the lines in a given series of X-rays. Therefore, he, concluded that atomic number was a better fundamental property of an element than its atomic weight. He then suggested that the atomic number (Z) instead of atomic weight should be basis of the classification of the elements.

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

Physical and chemical properties of the elements are the periodic functions of their atomic number. If the elements are arranged in order of their increasing atomic number, after a regular interval, elements with similar properties are repeated.

### Periodicity :

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

### Cause of Periodicity :

The periodic repetition of the properties of the elements is due to the recurrence of similar valence shell electronic configurations after certain regular intervals. For example, alkali metals have same valence shell electronic configuration  $ns^1$ , therefore, have similar properties.

The long form of periodic table is the contribution of Range, Werner, Bohr and Bury.

This table is also referred to as Bohr's table since it follows Bohr's scheme of the arrangements of elements into four types based on electronic configurations of elements.

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

### Periods :

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

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

- 1<sup>st</sup> period shortest period having only two elements. Filling of electrons takes place in the first energy shell, for which,  
 $n = 1, \ell = 0$  (s-subshell) and  $m = 0$ .  
Only one orbital (1s) is available and thus it contains only two elements.

- 3<sup>rd</sup> period short period having only eight elements. Filling of electrons takes place in the third energy level. For which,  
 $n = 3, \ell = 0, 1, 2$  and  
number of orbitals

$m = 0, 3, 5$		
1 3 5		
(3s) (3p) (3d)		
<hr/>		
Total number of orbitals	9	
	<hr/>	

But the energy of 3d orbitals are higher than 4s orbitals. Therefore, four orbitals (one 3s and three 3p orbitals) corresponding to  $n = 3$  are filled before filling in 4s orbital (next energy level). Hence 3<sup>rd</sup> period contains eight elements not eighteen elements.

### Groups :

There are eighteen groups numbered as 1, 2, 3, 4, 5, ..... 13, 14, 15, 16, 17, 18.

Group consists of a series of elements having similar valence shell electronic configuration.

**Table-3**

s-Block Elements		d-Block Elements																p-Block Elements					
1 IA																		18 VIII A					
1 H 1.007	2 II A																	13 III A	14 IV A	15 V A	16 VI A	17 VII A	18 He 4.002
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	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 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	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 Tl 204.383	82 Pb 207.2	83 Bi 208.980	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 Uun 269							114 Uuq							

**Inner - Transition Metals (f-Block elements)**

*Lanthanides	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
**Actinides	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

**Classification of the Elements :**

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

**(a) s-block elements**

When shells upto (n – 1) are completely filled and the last electron enters the s-orbital of the outermost (n<sup>th</sup>) shell, the elements of this class are called s-block elements.

- ◆ Group 1 & 2 elements constitute the s-block.
- ◆ General electronic configuration is [inert gas] ns<sup>1-2</sup>
- ◆ s-block elements lie on the extreme left of the periodic table.
- ◆ This block includes metals.

**(b) p-block elements**

When shells upto (n – 1) are completely filled and differentiating electron enters the p-orbital of the n<sup>th</sup> orbit, elements of this class are called p-block elements.

- ◆ Group 13 to 18 elements constitute the p-block.
- ◆ General electronic configuration is [inert gas] ns<sup>2</sup> np<sup>1-6</sup>
- ◆ p-block elements lie on the extreme right of the periodic table.
- ◆ This block includes some metals, all nonmetals and metalloids.
- ◆ s-block and p-block elements are collectively called normal or representative elements.

**(c) d-Block elements**

When outermost (n<sup>th</sup>) and penultimate shells (n – 1)<sup>th</sup> shells are incompletely filled and differentiating electron enters the (n – 1) d orbitals (i.e., d-orbital of penultimate shell) then elements of this class are called d-block elements.

- ◆ Group 3 to 12 elements constitute the d-block.
- ◆ General electronic configuration is [inert gas] (n – 1) d<sup>1-10</sup> ns<sup>1-2</sup> (except, palladium which has valence shell electron configuration 4d<sup>10</sup> 5s<sup>0</sup>).
- ◆ 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<sup>st</sup> transition series i.e. 3d series contains 10 elements and starts from  $_{21}\text{Sc} - _{30}\text{Zn}$ . Filling of electrons takes place in 3d sub-shell.
  - (2) II<sup>nd</sup> transition series i.e. 4d series contains 10 elements and starts from  $_{39}\text{Y} - _{48}\text{Cd}$ . Filling of electrons takes place in 4d sub-shell.
  - (3) III<sup>rd</sup> transition series i.e. 5d series 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) IV<sup>th</sup> transition series i.e. 6d series contains 10 elements and starts from  $_{89}\text{Ac}, _{104}\text{Rf} - _{112}\text{Uub}$ . Filling of electrons takes place in 6d sub-shell (incomplete series).
- ◆ Those elements which have partially filled d-orbitals in neutral state or in any stable oxidation state are called transition elements

#### (d) f-Block elements

When n, (n – 1) and (n – 2) shells are incompletely filled and last electron enters into f-orbital of antepenultimate i.e., (n – 2)th shell, elements of this class are called f-block elements. General electronic configuration is (n – 2) f<sup>1-14</sup> (n – 1) d<sup>0-1</sup> ns<sup>2</sup>

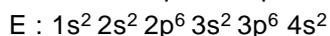
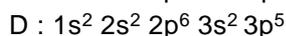
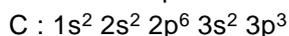
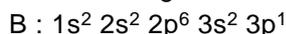
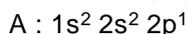
- ◆ All f-block elements belong to 3<sup>rd</sup> group.
- ◆ They are metals
- ◆ Within each series, the properties of the elements are quite similar.
- ◆ The elements coming after uranium are called transuranium elements.
- ◆ They are also called as inner-transition elements as they contain three outer 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. I<sup>st</sup> inner transition or 4 f-series, contains 14 elements  $_{58}\text{Ce}$  to  $_{71}\text{Lu}$ . Filling of electrons takes place in 4f subshell.
  2. II<sup>nd</sup> inner transition or 5 f-series, contains 14 elements  $_{90}\text{Th}$  to  $_{103}\text{Lr}$ . Filling of electrons takes place in 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.

## Solved Examples

**Example-1** 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 ?

**Solution** Out of these, elements A and B will belong to the same group of the periodic table because they have same outer electronic configuration, ns<sup>2</sup> np<sup>1</sup>.

### 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
  - (c) For d-block elements, Group number = number of electrons in (n – 1) d sub shell + number of electrons in valence shell.

## Solved Examples

**Example-2** An element X with  $Z = 112$  has been recently discovered. What is the electronic configuration of the element? To which group and period will it belong?

- Solution**
- The electronic configuration of element X is  $[\text{Rn}]^{86} 5f^{14} 6d^{10} 7s^2$
  - It belongs to d-block as last electron enters in d subshell.
  - As number of electrons in  $(n - 1)d$  subshell and valence shell is equal to twelve i.e.  $10 + 2$ . So it belongs to group 12.
  - It belongs to period 7 of the periodic table as principal quantum number of valence shell is 7 (i.e.,  $7s^2$ ).

### Metals and nonmetals :

- The metals are characterised by their nature of readily giving up the electron(s) and from shining lustre. 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.  $\text{CrO}_3$ ).
- Nonmetals do not lose electrons but take up electrons to form corresponding anions. Nonmetals are located at the top right hand side of the periodic table. Nonmetals are usually solids, liquids or gases at room temperature with low melting and boiling points. They are poor conductors of heat and electricity. Oxides of nonmetals are generally acidic in nature.

### Metalloids (Semi metals) :

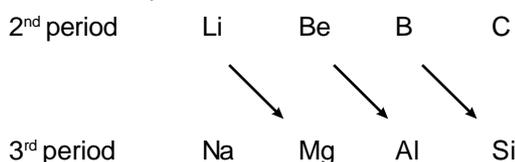
- It can be understood from the periodic table that nonmetallic character increases as we move from left to right across a row. It has been found that some elements which lie at the border of metallic and nonmetallic behavior, possess the properties that are characteristic of both metals and nonmetals.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

### Typical elements :

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

### Diagonal relationship :

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



Diagonal relationship arises because of ;

- on descending a group, the atoms and ions increase in size. On moving from left to right in the periodic table, the size decreases. Thus on moving diagonally, the size remains nearly the same.  
( $\text{Li} = 1.23 \text{ \AA}$  &  $\text{Mg} = 1.36 \text{ \AA}$  ;  $\text{Li}^+ = 0.76 \text{ \AA}$  &  $\text{Mg}^{2+} = 0.72 \text{ \AA}$ )

- (ii) it is sometimes suggested that the diagonal relationship arises because of diagonal similarity in electronegativity values.  
(Li = 1.0 & Mg = 1.2 ; Be = 1.5 & Al = 1.5 ; B = 2.0 & Si = 1.8)
- (iii) Be and Al also show a diagonal relationship. In this case sizes are not so close ( $\text{Be}^{2+} = 0.45 \text{ \AA}$  and  $\text{Al}^{3+} = 0.535 \text{ \AA}$ ) but the charge per unit area is nearly similar ( $\text{Be}^{2+} = 2.36$  and  $\text{Al}^{3+} = 2.50$ ) because the charges are 2+ and 3+ respectively.

$$\text{Charge per unit area} = \frac{(\text{ionic charge})}{\frac{4}{3} \cdot \pi \cdot (\text{ionic radius})^2}$$

- ◆ Similarities between properties of Li and Mg are as follows.
  - (a) Li and Mg both reacts directly with nitrogen to form lithium nitride ( $\text{Li}_3\text{N}$ ) and magnesium nitride ( $\text{Mg}_3\text{N}_2$ ) whereas other alkali metals of IA group does not form nitride.
  - (b) Fluoride carbonate and phosphate of Li and Mg are insoluble in water whereas these compounds of other alkali metals are soluble.
  - (c) Li and Mg both are hard metals, whereas other metals of IA group are soft.
  - (d)  $\text{LiOH}$  and  $\text{Mg}(\text{OH})_2$  both are weak bases, whereas hydroxides of other elements of IA group are strong base.
  - (f) Metallic bond in Li and Mg both are strong compare to other alkali metals .
  - (g) Their melting and boiling points are high.
  - (h) By thermal disintegration of  $\text{LiNO}_3$  and  $\text{Mg}(\text{NO}_3)_2$  ;  $\text{Li}_2\text{O}$  and  $\text{MgO}$  is obtained respectively.
  - (i) Thermal stability of  $\text{Li}_2\text{CO}_3$  and  $\text{MgCO}_3$  is very less compare to other alkali metals and they liberates  $\text{CO}_2$  gas easily.
- ◆ Similarly Be shows similarity to Al of IIIA group compare to other elements of IIA group which are as follows.
  - (a) These both elements do not provide colour to Bunsen burner.
  - (b) They both are comparatively stable in air.
  - (c) Both are insoluble in  $\text{NH}_3$  therefore do not form blue coloured solution.
  - (d) There is no tendency of making peroxide and superoxide by them.
  - (e) Reducing power is very less due to low value of standard electrode potential in the form of oxidation potential.
  - (f) The oxides and hydroxides of Be and Al both are amphoteric in nature.

### The periodicity of atomic properties :

- (i) **Effective nuclear charge :**  
Between the outer most valence electrons and the nucleus of an atom, there exists number of shells containing electrons. Due to the presence of these intervening electrons, the valence electrons are unable to experience the attractive pull of the actual number of protons in the nucleus. These intervening electrons act as shield between the valence electrons and protons in the nucleus. Thus, the presence of intervening (shielding) electrons reduces the electrostatic attraction between the protons in the nucleus and the valence electrons because intervening electrons repel the valence electrons. The concept of effective nuclear charge helps in understanding the effects of shielding on periodic properties. The effective nuclear charge ( $Z_{\text{eff}}$ ) is the charge actually felt by the valence electron.  $Z_{\text{eff}}$  is given by  $Z_{\text{eff}} = Z - \sigma$ , (where  $Z$  is the actual nuclear charge (atomic number of the element) and  $\sigma$  is the shielding (screening) constant). The value of  $\sigma$  i.e. shielding effect can be determined using the Slater's rules.

## Solved Examples

**Example-3** What is the effective nuclear charge at the periphery of nitrogen atom when an extra electron is added during the formation of an anion. Compare the value of  $Z_{\text{eff}}$  when the atom is ionized to  $\text{N}^+$ .

**Solution** Ground state electron configuration of  $\text{N}(Z = 7) = 1s^2 2s^2 2p^3$

Electron configuration of  $\text{N}^- = (1s^2) (2s^2 2p^4)$

Shielding constant for the last 2p electron,

$$\sigma = [(5 \times 0.35) + (2 \times 0.85)] = 3.45$$

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

So  $Z_{\text{eff}}$  for last p electron on  $\text{N}^- = 7 - 3.45 = 3.55$

Electron configuration of  $\text{N}^+ = (1s^2) (2s^2 2p^2)$

Shielding constant for the last 2p electron,

$$\sigma = [(3 \times 0.35) + (2 \times 0.85)] = 2.75$$

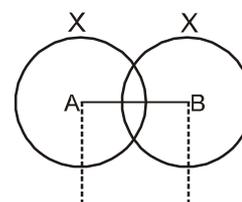
So  $Z_{\text{eff}}$  for last electron on  $\text{N}^+ = 7 - 2.75 = 4.25$

**(ii) 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.



$$r_{\text{covalent}} = \frac{1}{2} AB \text{ (of element X)}$$

◆ Covalent radius is generally used for nonmetals.

**Single Bond Covalent Radius, SBCR (bond length) -**

(a) For homodiatom molecules  $d_{A-A} = r_A + r_A$  or  $2r_A$

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

(b) For heterodiatom molecules in which electronegativity remains approximately same.

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

◆ For heteronuclear diatomic molecule, A-B, where difference between the electronegativity values of atom A and atom B is relatively larger,

$$d_{A-B} = r_A + r_B - 9.0 \Delta\chi$$

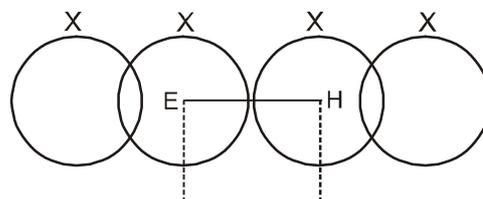
This formula was given by **Stevenson & Schomaker**. Electronegativity values are given in Pauling units and radius in picometers.

$\Delta\chi = X_A - X_B$  where  $X_A$  and  $X_B$  are electronegativity values of high electronegative element A and less electronegative element B.

Later on modified and more accurate formula was proposed by **Porterfield** as given below ;

$$d_{A-B} = r_A + r_B - 7.0 (\Delta\chi)^2$$

**(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.



$$r_{\text{vander Waal's}} = \frac{1}{2} EH$$

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

### Comparison of covalent radius and van der Waal's radius

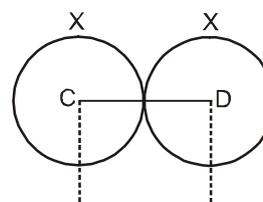
- The vander Waal's force of attractions are weak, therefore, the internuclear distances in case of atoms held by van der Waal's forces are much larger than those of between covalently bonded atoms. Therefore van der Waal's radii are always larger than covalent radii.
- A covalent bond is formed by the 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-6**

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

**(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 figure.



$$r_{\text{crystal}} = \frac{1}{2} CD \text{ (of element X)}$$

- ◆ 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 covalent, vander Wall's and metallic radius magnitude wise follows the order,

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

**Table-7**

Variation in a Period	Variation in a Group
In a period left to right :	In a group top to bottom :
Nuclear charge (Z) increases by one unit	Nuclear charge (Z) increases by more than one unit
Effective nuclear charge ( $Z_{\text{eff}}$ ) also increases	Effective nuclear charge ( $Z_{\text{eff}}$ ) almost remains constant because of increased screening effect of inner shells electrons.
But number of orbitals (n) remains constant	But number of orbitals (n) increases.
As a result, the electrons are pulled closer to the nucleus by the increased $Z_{\text{eff}}$ .  $r_n \propto \frac{1}{Z^*}$ Hence atomic radii decrease with increase in atomic number in a period from left to right.	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.

- The atomic radius of inert gas (zero group) is given largest in a period because it is represented as vander Waals's radius which is generally larger than the covalent radius. The vander Waal's radius of inert gases also increases from top to bottom in a group.

- In the transition series (e.g. in first transition series), the covalent radii of the elements decrease from left to right across a row until near the end when the size increases slightly. On passing from left to right, extra protons are placed in the nucleus and the extra electrons are added. The orbital electrons shield the nuclear charge incompletely. Thus the nuclear charge attracts all the electrons more strongly, hence a contraction in size occurs. The radii of the elements from Cr to Cu, are very close to one another because the successive addition of d-electrons screen the outer electrons (4s) from the inward pull of the nucleus. As a result of this, the size of the atom does not change much in moving from Cr to Cu.

**Table-8**

Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
<b>Atomic radius (Å)</b>	1.44	1.32	1.22	1.18	1.17	1.17	1.16	1.15	1.17	1.25

- The elements of group 3 belonging to d-block show the expected increase in size as that found in case of s-block and p-block elements. However, in the subsequent groups that is 4<sup>th</sup> onwards (upto 12<sup>th</sup> group), there is increase in size between the first and second member, but hardly any increase between the second and third elements of the given group. There are 14 lanthanide elements between lanthanum and hafnium, in which the antepenultimate 4f shell of electrons (exert very poor shielding effect) is filled. There is a gradual decrease in size of the 14 lanthanide elements from cerium to lutetium. This is called lanthanide contraction. This lanthanide contraction cancels out the normal size increase on descending a group in case of transition elements.

**(iii) 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-9**

Cation	Anion
It is formed by the lose of one or more electrons from the valence shell of an atom of an element. Cations are smaller than the parent atoms because, (i) the whole of the outer shell of electrons is usually removed. (ii) 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.	It is formed by the gain of one or more electrons in the valence shell of an atom of an element. Anions are larger than the parent atoms because (i) 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. (ii) 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.

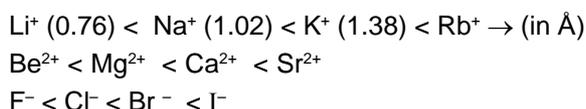
For example :	Na (metallic radius 1.86 Å)	Na <sup>+</sup> (ionic radius 1.02 Å)
Number of electrons :	11	10
Number of protons :	11	11
Electronic configuration :	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup> 3s <sup>1</sup>	1s <sup>2</sup> 2s <sup>2</sup> 2p <sup>6</sup>

There is complete removal of one shell leading to decrease in size.

	Cl (covalent radius 0.99 Å)	Cl <sup>-</sup> (ionic radius 1.84 Å)
Number of electrons :	17	18
Number of protons :	17	17

So, there is reduction in effective nuclear charge and hence the electron cloud expands in case of Cl<sup>-</sup>.

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

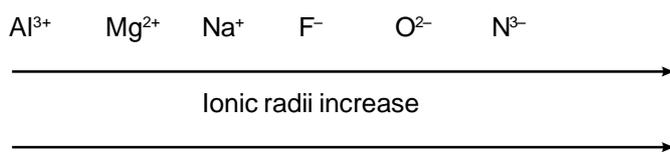


- ◆ The ionic radius decreases as more electrons are ionized off that is as the valency increases e.g. Cr<sup>2+</sup> = 0.80 Å, Cr<sup>3+</sup> = 0.615 Å, Cr<sup>4+</sup> = 0.55 Å, Cr<sup>5+</sup> = 0.49 Å and Cr<sup>6+</sup> = 0.44 Å.

◆ 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 ( $Hf^{4+} = 0.71 \text{ \AA}$ ) and Zr ( $Zr^{4+} = 0.72 \text{ \AA}$ ), and Nb ( $Nb^{3+} = 0.72 \text{ \AA}$ ) and Ta ( $Ta^{3+} = 0.72 \text{ \AA}$ ) 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-4** X – X bond length is 1.00 Å and C–C bond length is 1.54 Å. If electronegativities of X and C are 3.0 and 2.0 respectively, then C–X bond length is likely to be ? (using **Stevenson & Schomaker** formula).

**Solution**  $r_{C-X} = r_C + r_X - 0.09 \Delta\chi$   
 $= \frac{1.00}{2} + \frac{1.54}{2} - 0.09[\Delta\chi = 1] = 1.27 - 0.09 ;$

**Ans. C–X bond length = 1.18 Å.**

**Example-5** Atomic radius of Li is 1.23 Å and ionic radius of  $Li^+$  is 0.76 Å. Calculate the percentage of volume occupied by single valence electron in Li.

**Solution** Volume of Li =  $\frac{4}{3} \times 3.14 \times (1.23)^3 = 7.79 \text{ \AA}^3$  (– Li =  $1s^2 2s^1$ )

Volume of  $Li^+$  =  $\frac{4}{3} \times 3.14 \times (0.76)^3 = 1.84 \text{ \AA}^3$  (–  $Li^+$  =  $1s^2$ )

∴ Volume occupied by 2s subshell =  $7.79 - 1.84 = 5.95 \text{ \AA}^3$ .

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

**Example-6** Select from each group the species which has the smallest radius stating appropriate reason.

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

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

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

**Example-7**  $Mg^{2+}$  is smaller than  $O^{2-}$  in size, though both have same electronic configuration. Explain ?

**Solution**  $Mg^{2+}$  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  $O^{2-}$ .

**(iv) 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 1<sup>st</sup>, 2<sup>nd</sup> & 3<sup>rd</sup> ionization energies to remove electron from a neutral atom, monovalent and divalent cations respectively.

In general,  $(IE)_1 < (IE)_2 < (IE)_3 < \dots$  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  $\text{kJ mol}^{-1}$ ,  $\text{k Cal mol}^{-1}$ ,  $\text{eV}$  (electron volt).

♦ **Factors Influencing Ionisation energy**

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

**(A) Size of the Atom :**

Ionisation energy decreases with increase in atomic size.

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

For example, ionisation energy decreases in a group from top to bottom with increase in atomic size.

**(B) Nuclear Charge :**

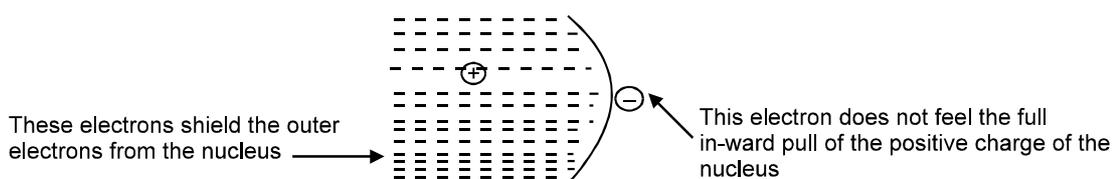
The ionisation energy increases with increase in the nuclear charge.

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

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

**(C) Shielding or screening effect :**

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



**(D) Penetration effect of the electron :**

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

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

(E) **Electronic Configuration :**

If an atom has exactly half-filled or completely filled orbitals, then such an arrangement has 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 outer most completely filled outer most 2s orbital while boron has partially filled less stable outer most 2p-orbital.



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

- 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 gaseous phase increases as the value of ionisation energy decreases. Amongst alkali metals, the lithium is strongest reducing agent in aqueous solution.

## Solved Examples

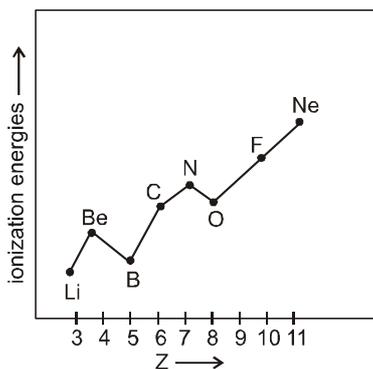
**Example-8** From each set, choose the atom which has the largest ionization enthalpy and explain your answer with suitable reasons. (a) F, O, N (b) Mg, P, Ar

**Solution** (a) Fluorine (F) has the largest ionization enthalpy because in moving from left to right in a period, atomic size decreases and electrons are held more tightly. Since F has the smallest size and maximum nuclear charge. It has the largest ionization enthalpy among these elements.  
(b) Argon (Ar) has the largest ionization enthalpy as argon has extra stable fully filled configuration.

**Example-9** First and second ionisation energies of magnesium are 7.646 eV and 15.035 eV respectively. The amount of energy in kJ needed to convert all the atoms of magnesium into  $\text{Mg}^{2+}$  ions present in 12 mg of magnesium vapour will be ? [Given  $1 \text{ eV} = 96.5 \text{ kJ mol}^{-1}$ ].

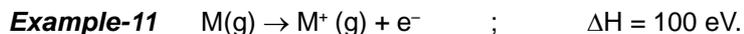
**Solution** Total energy needed to convert one Mg atom into  $\text{Mg}^{2+}$  gas ion,  
 $= \text{IE}_I + \text{IE}_{II} = 22.681 \text{ eV} = 2188.6 \text{ kJ mol}^{-1}$ .  
 $\Rightarrow 12 \text{ mg of Mg} = 0.5 \times 10^{-3} \text{ mole}$ .  
 $\therefore \text{Total energy} = 0.5 \times 10^{-3} \times 2188.6 = 1.0943 \text{ kJ}$  **Ans.**

**Example-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 –

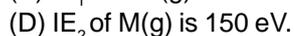
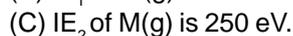
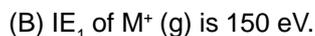
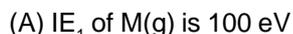


- (A) above Ne
- (B) below Ne but above O.
- (C) below Li
- (D) between N and O.

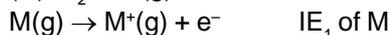
**Solution** Na is 3rd period element and is bigger than Li. The distance between the nucleus and outer most electron is more as compared to Li. Thus the outer most electron is loosely bound with nucleus and removal of electron is easier. So option (C) is correct.



Which is/are correct statement(s) ?



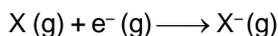
**Solution**



**Ans.** (A,B,D)

**(V) 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.



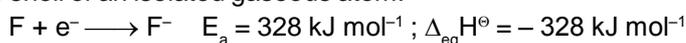
Electron gain enthalpy provides a measure of the ease with which an atom adds an electron to form an anion. Electron gain may be either exothermic or endothermic depending on the elements.

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 and the gaseous ions at  $T = 0$ .

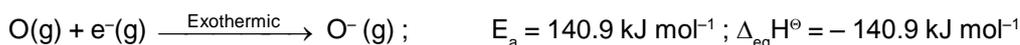
$$E_a = E(X, g) - E(X^-, g),$$

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



Although the precise relation is  $\Delta_{eg}H^\ominus = -E_a - \frac{5}{2}RT$ , the contribution  $\frac{5}{2}RT$  is commonly ignored.

A positive electron affinity indicates that the ion  $X^-$  has a lower, more negative energy than the neutral atom X. The second electron gain enthalpy, the enthalpy change for the addition of a second electron to an initially neutral atom, invariably positive because the electron repulsion outweighs 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 ( $Z_{\text{eff}}$ )
- (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 a system is difficult and hence the electron affinity value decreases.

## Solved Examples

**Example-12** Consider the elements N, P, O and S and arrange them in order of increasing negative electron gain enthalpy.

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

**Example-13** Why do halogens have high electron gain enthalpies (i.e.  $-\Delta_{\text{eg}}H^\ominus$ ) ?

**Solution** 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-14** Which will have the maximum value of electron affinity  $O^x, O^y, O^z$  [x, y and z respectively are 0, -1 and -2]?

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

**Solution** 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-15** The amount of energy when million atoms of iodine are completely converted into  $I^-$  ions in the vapour state according to the equation,  $I(g) + e^-(g) \rightarrow I^-(g)$  is  $5.0 \times 10^{-13}$  J.

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

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

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

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

**Example-16** Account for the large decrease in electron affinity between Li and Be despite the increase in nuclear charge.

**Solution** 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.

### (VI) 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  $\propto$  electronegativity of the atom.

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

Hybrid orbital	sp <sup>3</sup>	sp <sup>2</sup>	sp
s-character	25%	33%	50%

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

**Table-10**

Variation of electronegativity in a group	Variation of electronegativity in a period
On moving down the groups, Z increases but $Z_{\text{eff}}$ 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, $Z_{\text{eff}}$ 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

$$\Delta = X_A - X_B = 0.208 \sqrt{E_{A-B} - \sqrt{E_{A-A} \times E_{B-B}}}$$

$E_{A-B}$  = Bond enthalpy/ Bond energy of A – B bond.

$E_{A-A}$  = Bond energy of A – A bond

$E_{B-B}$  = Bond energy of B – B bond

**(All bond energies are in kcal / mol)**

$$\Delta = X_A - X_B = 0.1017 \sqrt{E_{A-B} - \sqrt{E_{A-A} \times E_{B-B}}}$$

All bond energies are in kJ / mol.

- (b) **Mulliken's scale** : Electronegativity  $\chi$  (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  $\chi_P$  is related to Mulliken's electronegativity  $\chi_M$  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.

### APPLICATION OF ELECTRONEGATIVITY :

- (a) **Nomenclature** :

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

Correct formula	Name
(a) I <sup>+</sup> Cl <sup>-</sup>	Iodine chloride
(b) Cl <sup>+</sup> F <sup>-</sup>	Chlorine fluoride
(c) Br <sup>+</sup> Cl <sup>-</sup>	Bromine chloride
(d) IBr	Iodine bromide
(e) OF <sub>2</sub>	Oxygen difluoride
(f) Cl <sub>2</sub> O	Dichlorine oxide

- (b) **Metallic and Nonmetallic Nature** :

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

(c) **Partial Ionic Character in Covalent bonds :**

Partial ionic characters are generated in covalent compounds by the difference of electronegativities.

**Hanney and Smith** calculated percentage of ionic character from the difference of electronegativity.

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

$X_A$  is electronegativity of element A (Higher)

$X_B$  is electronegativity of element B (Lower)

$$\Delta = X_A - X_B$$

(d) **Bond Strength & Stability :**

As the difference in the electronegativities between two bonded atoms in a molecule increases, the bond strength and stability increase. For example  $\text{H-F} > \text{H-Cl} > \text{H-Br} > \text{H-I}$

(VII) **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 :  $\text{OF}_2$  and  $\text{Na}_2\text{O}$ . The order of electronegativity of the three elements involved in these compounds is  $\text{F} > \text{O} > \text{Na}$ . Each of the atoms of fluorine, with outer electronic configuration  $2s^2 2p^5$ , shares one electron with oxygen in the  $\text{OF}_2$  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  $\text{Na}_2\text{O}$ , 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.

## Solved Examples

**Example-17** Calculate the electronegativity of carbon from the following data :

$$\begin{aligned}E_{\text{H-H}} &= 104.2 \text{ kcal mol}^{-1}, & E_{\text{C-C}} &= 83.1 \text{ kcal mol}^{-1} \\ E_{\text{C-H}} &= 98.8 \text{ kcal mol}^{-1}, & X_{\text{H}} &= 2.1\end{aligned}$$

**Solution** Let the electronegativity of carbon be  $X_C$ , applying Pauling equation,

$$\begin{aligned}X_C - X_H &= 0.208 [E_{\text{C-H}} - \frac{1}{2} (E_{\text{C-C}} + E_{\text{H-H}})]^{1/2} \\ X_C - 2.1 &= 0.208 [98.8 - \frac{1}{2} (83.1 + 104.2)]^{1/2} \\ \Rightarrow X_C &= 2.5 & \text{Ans. } & 2.5\end{aligned}$$

**Example-18** If electronegativity of x be 3.2 and that of y be 2.2, the percentage ionic character of xy is –  
(A) 19.5 (B) 18.5 (C) 9.5 (D) 29.5.

**Solution**  $EN_x - EN_y = 3.2 - 2.2 = 1$ .  
 $\Delta = 1$

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

$$\% \text{ ionic character} = 16\Delta + 3.5\Delta^2 = 19.5. \quad \text{Ans. (A)}$$

**Example-19** Which of the following relation is correct ?

- (A)  $2 \text{ IP} - \text{E.A.} - \text{EN} = 0$  (B)  $2\text{EN} - \text{IP} - \text{EA} = 0$   
(C)  $2\text{EA} - \text{IP} - \text{EA} = 0$  (D)  $\text{EN} - \text{IP} - \text{EA} = 0$

**Solution** On Mulliken's scale :  $\text{E.N.} = \frac{\text{I.P.} + \text{E.A.}}{2}$   
 $2\text{E.N.} - \text{I.P.} - \text{E.A.} = 0 \quad \text{Ans. (B)}$

## (VIII) Periodic Trends and Chemical Reactivity :

As the periodicity is related to electronic configuration, all chemical and physical properties are a manifestation of the electronic configuration of elements.

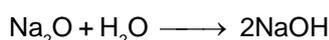
The atomic and ionic radii, as we know, generally decrease in a period from left to right. As a consequence, the ionization enthalpies generally increase (with some exceptions) and electron gain enthalpies become more negative across a period. In other words, the ionization enthalpy of the extreme left element in a period is the least and the electron gain enthalpy of the element on the extreme right is the highest negative (except noble gases which having completely filled shells have rather positive electron gain enthalpy values). This results in high chemical reactivity at the two extremes and the lowest in the centre. Thus, the maximum chemical reactivity at the extreme left (among alkali metals) is shown by the loss of an electron leading to the formation of cation and at the extreme right (among halogens) shown by the gain of an electron forming an anion. The loss and gain of electron can be related with the reducing and oxidizing behaviour of the elements respectively. However, it can also be directly related to the metallic and nonmetallic character of elements. Thus, the metallic character of an element, which is highest at the extremely left decreases and the nonmetallic character increases while moving from left to right across the period. The chemical reactivity of an element can be best understood by its reactions with oxygen and halogens. Elements on two extremes of a period easily combine with oxygen to form oxides. The normal oxide formed by the element on extreme left is the most basic (e.g.  $\text{Na}_2\text{O}$ ), whereas that formed by the element on extreme right is the acidic (e.g.  $\text{Cl}_2\text{O}_7$ ). Oxides of elements in the centre are amphoteric (e.g.  $\text{Al}_2\text{O}_3$ ,  $\text{As}_2\text{O}_3$ ) or neutral (e.g.,  $\text{CO}$ ,  $\text{NO}$ ,  $\text{N}_2\text{O}$ ). Amphoteric oxides behave as acidic with bases and as basic with acids, whereas neutral oxides have no acidic or basic properties.

In general, metallic oxides ( $\text{O}^{2-}$ ), peroxides ( $\text{O}_2^{2-}$ ) and super oxides ( $\text{O}_2^{1-}$ ) are ionic solids.

The tendency of group IA metals (alkali metals) to form oxygen rich compounds increases from top to bottom i.e. with increasing cation radii and decreasing charge density on the metal ion.

IIA metals also show the similar trend. Except Be, the IIA metals react with oxygen at normal conditions to form normal ionic oxides and at high pressure of  $\text{O}_2$ , they form peroxides ( $\text{CaO}_2$ ,  $\text{SrO}_2$ ,  $\text{BaO}_2$ ). Oxides of metals are called as basic anhydrides as most of them combine with water forming hydroxides with no change in oxidation state of metals.

Oxides of IA and IIA dissolve in water forming basic solutions where as other oxides do not dissolve in water.



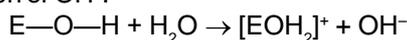
Oxygen combines with many nonmetals to form covalent oxides such as  $\text{CO}$ ,  $\text{CO}_2$ ,  $\text{SO}_2$ ,  $\text{SO}_3$ ,  $\text{P}_4\text{O}_{10}$ ,  $\text{Cl}_2\text{O}_7$  etc. Nonmetals with limited supply of oxygen usually form oxides in which nonmetals are present in lower oxidation states where as with excess of oxygen, oxides with higher oxidation state are formed. Oxides of nonmetals are called as acid anhydrides as most of them dissolve in water forming acids or oxy-acids.



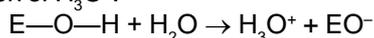
- In a group, basic nature of oxides increases or acidic nature decreases. Oxides of the metals are generally basic and oxides of the nonmetals are acidic. The oxides of the metalloids are generally amphoteric in nature. The oxides of Be, Al, Zn, Sn, As, Pb and Sb are amphoteric.
- In a period the nature of the oxides varies from basic to acidic.

$\text{Na}_2\text{O}$	$\text{MgO}$	$\text{Al}_2\text{O}_3$	$\text{SiO}_2$	$\text{P}_4\text{O}_{10}$	$\text{SO}_3$	$\text{Cl}_2\text{O}_7$
Strongly basic	Basic	amphoteric	Weakly acidic	Acidic	Acidic	Strongly acidic

- Oxides of the elements E in  $\text{H}_2\text{O}$  produce EOH.  
If electronegativities of E and O is greater than that of H and O in  $\text{H}_2\text{O}$  then EOH is basic due to the formation of  $\text{OH}^-$ .



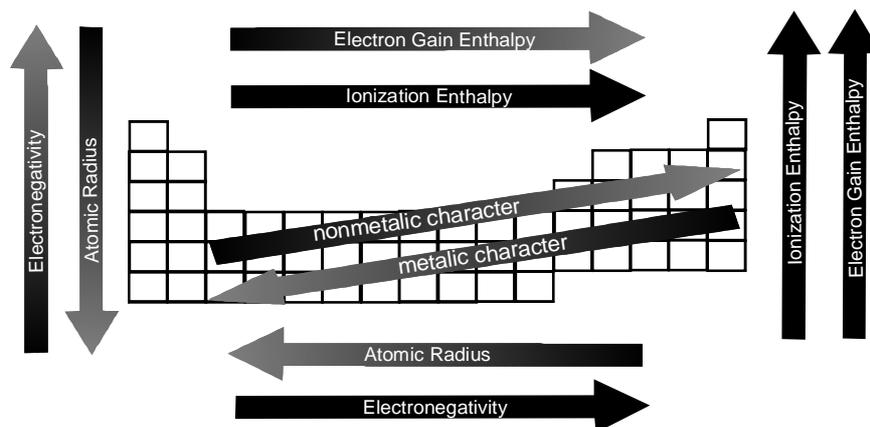
If electronegativities of E and O is less than that of H and O in  $\text{H}_2\text{O}$  then EOH is acidic due to the formation of  $\text{H}_3\text{O}^+$ .



Across the 3d-series from left to right, the change in atomic radii is much smaller as compared to those of representative elements. The change in atomic radii is still smaller among inner-transition metals (4f series). The ionization enthalpies of transition metals are intermediate between those of s- and p-blocks. As a result, they are less electropositive than group 1 and 2 metals.

In a particular group, the increase in atomic and ionic radii with increase in atomic number generally results in a gradual decrease in ionization enthalpies and a regular decrease (with exception in some third period elements) in electron gain enthalpies in the case of main group elements. Thus the metallic character increases down the group and nonmetallic character decreases. This trend can be correlated with their reducing and oxidizing property.

In the case of transition elements, however, a reverse trend is observed. This can be explained in terms of atomic size and ionization enthalpy as shown in figure.



## Solved Examples

**Example-20** Arrange the following in decreasing basic nature LiOH, NaOH, KOH, CsOH.  
LiOH, NaOH, KOH, CsOH

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

**Example-21** Arrange the following compounds in increasing order of acidic strength Al<sub>2</sub>O<sub>3</sub>, SiO<sub>2</sub>, P<sub>2</sub>O<sub>3</sub> and SO<sub>2</sub>.

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

**Ans.** Al<sub>2</sub>O<sub>3</sub> < SiO<sub>2</sub> < P<sub>2</sub>O<sub>3</sub> < SO<sub>2</sub>.

## MISCELLANEOUS SOLVED PROBLEMS

1. Which of the following group does not represent the isoelectronic species ?

(A) CH<sub>4</sub>, H<sub>2</sub>O, NH<sub>3</sub>, HF

(B) PH<sub>3</sub>, SiH<sub>4</sub>, HS<sup>-</sup>, Ar

(C) OH<sup>-</sup>, H<sub>2</sub>O, NH<sub>2</sub><sup>-</sup>, F<sup>-</sup>

(D) H<sub>2</sub>S, K<sup>+</sup>, Ar, Cl

**Ans.** (D)

**Sol.** As all species do not have same number of electrons i.e., 18. H<sub>2</sub>S, K<sup>+</sup> and Ar have same number of electrons i.e. 18 but Cl has 17 electrons.

2. Which of the following compound is isoelectronic with [NH<sub>3</sub> + BH<sub>3</sub>] ?

(A) B<sub>2</sub>H<sub>6</sub>

(B) C<sub>2</sub>H<sub>6</sub>

(C) C<sub>2</sub>H<sub>4</sub>

(D) C<sub>3</sub>H<sub>6</sub>

**Ans.** (B)

**Sol.** C<sub>2</sub>H<sub>6</sub> (12 + 6 = 18) has 18 electrons and [NH<sub>3</sub> + BH<sub>3</sub>] = 10 + 8 = 18. So, both are isoelectronic with each other.

3. A M<sup>2+</sup> 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  $M^{2+}$  is :  $[Ar]^{18} 4s^0 3d^4$   
 $\therefore$  Electron configuration of M is :  $[Ar]^{18} 4s^1 3d^5$  (and not  $4s^2 3d^4$ )  
 So total number of electrons = 24.  
 Hence, metal M is chromium (Cr).

4. Following are the valence shell electronic configurations of some elements.  
 (i)  $3s^2 3p^5$  (ii)  $3d^{10} 4s^2$  (iii)  $2s^2 3p^6 4s^1$  (iv)  $1s^2 2s^2$   
 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.

5. Find out the group of the element having the electronic configuration,  $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$ .

**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 =  $6 + 2 = 8$ .

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

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

**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.

7. The ( $IE_1$ ) and the ( $IE_2$ ) in  $\text{kJ mol}^{-1}$  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( $ns^2$ ).

8. 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 ( $\chi_M$ ) =  $\frac{\text{Ionisation energy} + \text{Electron affinity}}{2}$   
 $= \frac{17.42 + 3.45}{2} = 10.435$

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

**Ans.**  $\chi_P = 3.726$

9. Select neutral, acidic, basic and amphoteric oxides from the following.  
 $CO$ ,  $BeO$ ,  $Na_2O$ ,  $N_2O_5$

**Sol.**  $BeO$  is amphoteric and reacts with acids and bases forming salts.

$CO$  is neutral as it is neutral towards litmus.

$Na_2O$  is basic and dissolves in water forming base ( $NaOH$ ).

$N_2O_5$  is acidic and dissolves in water forming acid ( $HNO_3$ )

10. 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.

11. 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
(A) $ns^2, np^5$	(p) Chromium
(B) $(n-1)d^{10}, ns^1$	(q) Copper
(C) $(n-1)d^5, ns^1$	(r) Krypton
(D) $(n-1)d^{10}, ns^2, np^6$	(s) Bromine

**Ans.** (A)  $\rightarrow$  (s) ; (B)  $\rightarrow$  (q) ; (C)  $\rightarrow$  (p) ; (D)  $\rightarrow$  (r).

**Sol.** (A)  $ns^2np^5$  is general valence shell electron configuration of halogens. So this configuration belongs to bromine.  
 (B)  $(n-1)d^{10}ns^1$ ; This is electron configuration of d-block elements. As it contains  $(n-1)d^{10}ns^1$  configuration it belongs to copper.

(C)  $(n-1)d^5ns^1$ ; This is electron configuration of d-block elements. As it contains  $(n-1)d^5ns^1$  configuration it belongs to chromium.

(D) Noble gases has valence shell electron configuration  $ns^2 np^6$ , so it belongs to krypton.

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

Column-I	Column-II
(A) Metalloid	(p) Sulphur
(B) Radioactive	(q) Gold
(C) Transition metal	(r) Arsenic
(D) Chalcogen	(s) Uranium

**Ans.** (A)  $\rightarrow$  (r) ; (B)  $\rightarrow$  (s) ; (C)  $\rightarrow$  (q) ; (D)  $\rightarrow$  (p)

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

(B) Uranium is a radioactive element.

(C) 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  $[Xe]^{54}, 5d^86s^0$ .

(D) 16<sup>th</sup> group elements like oxygen and sulphur are ore forming elements and therefore are called as chalcogens.

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

Column-I	Column-II
(A) Representative element	(p) Cerium
(B) Lanthanide	(q) Aluminium
(C) Coinage metal	(r) Thorium
(D) Actinide	(s) Gold

**Ans.** (A)  $\rightarrow$  (q) ; (B)  $\rightarrow$  (p) ; (C)  $\rightarrow$  (s) ; (D)  $\rightarrow$  (r)

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

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

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

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

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

Column-I	Column-II
(A) Isoelectronic species	(p) $A^+(g) + \text{energy} \rightarrow A^{++}(g) + e^-(g)$
(B) Half filled orbital	(q) Ar, $K^+$ , $Ca^{++}$
(C) Second ionisation energy	(r) Lutetium
(D) Inner transition element	(s) Antimony

**Ans.** (A)  $\rightarrow$  (q) ; (B)  $\rightarrow$  (s) ; (C)  $\rightarrow$  (p) ; (D)  $\rightarrow$  (r)

**Sol.** (A) 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.

(B)  $np^3$ ,  $(n-1)d^5$  and  $(n-2)f^7$  represent half filled orbitals. Antimony has  $([Kr]^{36}4d^{10}5s^25p^3)$ .

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

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

15. 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
(A) Highest 1 <sup>st</sup> ionisation energy	(p) Technitium
(B) Highest electronegativity	(q) Lithium
(C) Synthetic element	(r) Helium
(D) Strongest reducing agent	(s) Fluorine

**Ans.** (A) → (r) ; (B) → (s) ; (C) → (p) ; (D) → (q).

**Sol.** (A) Helium has highest 1<sup>st</sup> ionisation energy amongst all the elements of periodic table because of ns<sup>2</sup> valence electron configuration and its small size of atom.

(B) Fluorine has highest electronegativity i.e. 4.0 on Pauling scale on account of its small size.

(C) Technitium is a man made element.

(D) Lithium is a strongest reducing agent because of its highest negative value of E<sup>o</sup> due to its higher hydration energy on account of its small size of atom.

16. 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)
(A) An element whose fourth shell contains two p-electrons	(p) 8 <sup>th</sup> group
(B) An element whose valence shell contains one unpaired p-electron	(q) 12 <sup>th</sup> group
(C) An element which receives last electron in (n – 1) d-subshell	(r) 14 <sup>th</sup> group
(D) An element with the ground-state electron configuration [Ar]4s <sup>2</sup> 3d <sup>10</sup>	(s) 17 <sup>th</sup> group

**Ans.** (A) → (r) ; (B) → (s) ; (C) → (p, q) ; (D) → (q).

**Sol.** (A) [Ar]3d<sup>10</sup>4s<sup>2</sup>4p<sup>2</sup> : Fourth shell contains two electron in 4p-sub shell i.e., 4p<sup>2</sup>. Therefore, group number = 10 + 4 = 14.

(B) Halogens (i.e. group number 17) have valence shell electronic configuration ns<sup>2</sup>np<sup>5</sup> and there is one unpaired electron in p-subshell i.e.,  $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$

(C) 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<sup>2</sup>(n – 1)d<sup>6</sup>. Therefore, group number = 2 + 6 = 8. Like wise, group 12 is ns<sup>2</sup>(n – 1)d<sup>10</sup>. Therefore, group number = 2 + 10 = 12.

So in group 8 and 12 last electron enters in d-subshell.

(D) For electronic configuration. [Ar]4s<sup>2</sup>3d<sup>10</sup> the group number = 2 + 10 = 12.

## Board Level Exercise

### Type (I) : Very Short Answer Type Questions :

[01 Mark Each]

1. What is the basic theme of organisation in the periodic table ?
2. What is the basic difference in approach between the Mendeleev's Periodic Law and the Modern Periodic Law?
3. Write the atomic number of the element present in the fourth period and eighth group of the Modern periodic table.
4. What would be the IUPAC name and symbol for the element with atomic number 109 ?
5. Which element do you think would have been named by :  
(i) Lawrence Berkeley Laboratory (ii) Seaborg's group
6. Why do elements in the same period have different physical and chemical properties ? What about elements in the same group ?
7. Are the oxidation state and covalency of Be in  $[\text{Be}(\text{H}_2\text{O})_3(\text{OH})]^+$  same ?
8. Considering the atomic number and position in the Modern periodic table, arrange the following elements in the increasing order of metallic character : P, Mg, Ca, K, Si.

### Type (II) : Short Answer Type Questions :

[02 Marks Each]

9. In terms of block, period and group, where would you locate the element with  $Z = 114$  in Modern periodic table?
10. Write the general outer electronic configuration of s-, p-, d- and f-block elements.
11. Use the Modern periodic table to answer the following questions :  
(i) Identify an element with six electrons in the outermost shell.  
(ii) Identify an element that would tend to lose one electron.  
(iii) Identify an element that would tend to gain three electrons.  
(iv) Identify the group having metal, non-metal, liquid as well as gas at  $30^\circ\text{C}$  temperature.
12. Explain why cations are smaller and anions larger in radii than their parent atoms ?
13. Consider the following species :  
 $\text{P}^{3-}$ ,  $\text{S}^{2-}$ ,  $\text{Cl}^-$ ,  $\text{K}^+$ ,  $\text{Ca}^{2+}$  and  $\text{Sc}^{3+}$   
(i) What is common in them ?  
(ii) Arrange them in the order of decreasing ionic radii.
14. Which of the following species will have the largest and the smallest size ? K,  $\text{K}^+$ , Ca,  $\text{Ca}^{2+}$ . Explain.
15. Energy of an electron in the ground state of the hydrogen atom is  $-5.19 \times 10^{-19}$  Cal. Calculate the ionization enthalpy of atomic hydrogen in terms of KCal/mol.  
**Hint:** Apply the idea of mole concept to derive the answer.
16. What are the various factors due to which the ionization enthalpy of the main group elements tends to decrease down a group ?
17. The first ionization enthalpy ( $\Delta_i H$ ) values of the second period elements Li, Be and C are respectively 520, 899 and  $1086 \text{ kJmol}^{-1}$ . Predict whether the first  $\Delta_i H$  value for B will be more close to 690 or  $990 \text{ kJmol}^{-1}$  ? Justify your answer.
18. Would you expect the first ionization enthalpies for two isotopes of the same element to be the same or different ? Justify your answer.
19. Would you expect the second electron gain enthalpy of S as positive, more negative or less negative than the first ? Justify your answer.
20. What is the basic difference between the terms 'electron gain enthalpy' and 'electronegativity' ?

21. Using the Modern periodic table, predict the formulas of compounds which might be formed by the following pairs of elements : (i) Boron & Oxygen (ii) Calcium & Fluorine.
22. Show by a chemical reaction with water that BaO is a basic oxide and  $N_2O_5$  is an acidic oxide.
23. Which important property did Mendeleev use to classify the elements in his periodic table and did he stick to that ?
24. On the basis of quantum numbers, justify that the fourth period of the Modern periodic table should have 18 elements.
25. What does atomic radius and ionic radius really mean to you ?
26. How do atomic radius generally vary in a period and in a group for s- and p-block elements ? How do you explain the variation ?

**Type (III) : Long Answer Type Questions:**

**[03 Marks Each]**

27. Among the third period elements, the actual ionization enthalpies are in the order  $Na < Al < Mg < Si < S < P < Cl < Ar$ . Explain why :  
(i) Mg has higher  $\Delta_f H$  than Al.  
(ii) S has lower  $\Delta_f H$  than P and Cl.
28. How would you explain the fact that the first ionization enthalpy of Lithium is lesser than that of Beryllium but its second ionization enthalpy is greater than that of Beryllium ?
29. The first ionization enthalpy values (in  $\text{kJmol}^{-1}$ ) of group 13 elements are :  

B	Al	Ga	In	Tl
801	577	579	558	589

 How would you explain this deviation from the general trend ?
30. What is the significance of the terms — 'isolated gaseous atom' and 'ground state' while defining the ionization enthalpy and electron gain enthalpy ?  
[Hint : Requirements for comparison purposes]
31. Which of the following will have the most negative electron gain enthalpy and which the least negative ? O, S, F and Cl. Explain.
32. How would you react to the statement that the electronegativity of C on Pauling scale is 2.5 in all the carbon compounds ?

**Type (IV) : Very Long Answer Type Questions :**

**[05 Marks Each]**

33. Assign the position of the element having outer electronic configuration  
(i)  $ns^2np^2$  for  $n = 2$  (ii)  $(n - 1)d^5ns^1$  for  $n = 4$ , and (iii)  $(n - 2)f^{14}(n - 1)d^0ns^2$  for  $n = 6$ , in the Modern periodic table.
34. Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements :  

(a) Sodium and sulphur	(b) Calcium and phosphorus	(c) Boron and chlorine
(d) Carbon and sulphur	(e) Nitrogen and bromine	(f) Element 71 and fluorine

## Exercise # 1

### OBJECTIVE QUESTIONS

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

- A-1. Atomic weight of Cl = 35.5 and of I = 127. According to Doberiner triad rule, atomic. weight of Br will be:  
(1) 80.0 (2) 162.5 (3) 81.25 (4) 91.5
- A-2. The atomic volume was chosen as the basis of periodic classification of elements by :  
(1) Niels Bohr (2) Mendeleev (3) Lothar Maeyer (4) Newlands

- A-3.** Elements which occupied position in the Lothar Meyer curve, on the peaks, were :  
 (1) Alkali metals (2) Highly electro positive elements  
 (3) Elements having large atomic volume (4) All
- A-4.** According to modern periodic law the elements with similar chemical and physical properties repeat at regular intervals when the elements are arranged in order of :  
 (1) decreasing atomic number. (2) increasing atomic weight.  
 (3) increasing atomic number. (4) decreasing atomic weight.
- A-5.** How many periods and vertical columns are there in the long form of the periodic table?  
 (1) 8, 12 (2) 6, 8 (3) 7, 18 (4) 6, 18
- A-6.** An element has atomic number 37. The block and group of this element are respectively :  
 (1) s and 1<sup>st</sup> (2) p and 17<sup>th</sup> (3) s and 2<sup>nd</sup> (4) p and 13<sup>th</sup>
- A-7.** What is the characteristic valence shell electron configuration of 11<sup>th</sup> group metals ?  
 (1)  $ns^2 np^6$  (2)  $(n-1)d^2 ns^2$  (3)  $nd^9 ns^2$  (4)  $(n-1)d^{10} ns^1$
- A-8.** Which one of the following pairs of atomic numbers represents elements belonging to the same group?  
 (1) 13 and 31 (2) 11 and 20 (3) 14 and 33 (4) 12 and 30
- A-9.** Which of the following pairs of elements belongs to representative group of elements in the periodic table?  
 (1) Aluminium and Magnesium (2) Chromium and Zinc  
 (3) Argentum and Astatine (4) Lanthanum and Thorium
- A-10.** There are 10 neutrons in the nucleus of the elements  ${}_Z M^{19}$ . It belongs to :  
 (1) f-block (2) s-block (3) d-block (4) None of these
- A-11.** In the general electronic configuration  $(n-2)f^{1-14} (n-1)d^{0-1} ns^2$ , if value of  $n = 7$  the configuration will be:  
 (1) lanthanides (2) actinides  
 (3) transition elements (4) none of these
- A-12.** Pt, Ni, Au and Ti belongs to :  
 (1) f-block (2) d-block (3) p-block (4) s-block
- A-13.** Which of the following element is a metalloid ?  
 (1) Bi (2) Sn (3) Ge (4) C
- A-14.** In the periodic table, where are non-metals located?  
 (1) Between groups II A and III A (2) On the lower left hand side  
 (3) On the upper left hand side (4) On the upper right hand side
- A-15.** Modern periodic table is based on atomic number experiments which proved importance of atomic number was :  
 (1) Braggs work on X-ray diffraction  
 (2) Moseley's work on X-ray spectrum  
 (3) Mulliken's oil drop experiment  
 (4) Lothar Meyer curve plotted between atomic volume & atomic weight
- A-16.** Which one of the following statements related to the modern periodic table is **incorrect** ?  
 (1) The p-block has 6 columns, because a maximum of 6 electrons can occupy all the orbitals in a p-subshell.  
 (2) The d-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a d-subshell.  
 (3) Each block contains a number of columns equal to the number of electrons that can occupy that subshell.  
 (4) The block indicates value of azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration.
- A-17.** Element with electronic configuration as  $[Ar]^{18} 3d^5 4s^1$  is placed in :  
 (1) IA, s-block (2) VIA, s-block (3) VIB, s-block (4) VIB, d-block

- A-18.** Atomic number of Pd is 46. In the same group the atomic number of elements placed above and below Pd will be :  
 (1) 37, 67 (2) 28, 78 (3) 39, 69 (4) 18, 28
- A-19.** The first element of a group in many ways differs from the other heavier members of the group. This is due to:  
 (1) the small size (2) the high electronegativity and high ionisation potential  
 (3) the unavailability of d-orbitals (4) all of the above
- A-20.** The statement that is not correct for the periodic classification of elements is :  
 (1) In d-block elements, the last electron enters in  $(n - 1)d$  sub-shell.  
 (2) non-metallic elements are lesser in number than metallic elements.  
 (3) the third period contains 8 elements and not 18 as 4th period contains.  
 (4) for transition elements the d-subshells are filled with electrons monotonically with increase in atomic number.
- A-21.** The atomic numbers of the metallic and non-metallic elements which are liquid at room temperature respectively are :  
 (1) 55, 87 (2) 33, 87 (3) 35, 80 (4) 80, 35
- A-22.**  $\text{Na}^+$ ,  $\text{Mg}^{2+}$ ,  $\text{Al}^{3+}$ ,  $\text{Si}^{4+}$  are isoelectronic. Their ionic size follows the order :  
 (1)  $\text{Na}^+ < \text{Mg}^{2+} < \text{Al}^{3+} < \text{Si}^{4+}$  (2)  $\text{Na}^+ > \text{Mg}^{2+} < \text{Al}^{3+} < \text{Si}^{4+}$   
 (3)  $\text{Na}^+ < \text{Mg}^{2+} > \text{Al}^{3+} > \text{Si}^{4+}$  (4)  $\text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+} > \text{Si}^{4+}$

### Section (B) : Atomic and Ionic Radius

- B-1.** The atom larger in size as compared to oxygen is :  
 (1) F (2) He (3) Ne (4) none of these
- B-2.** Which of the following has the largest ionic radius ?  
 (1)  $\text{Na}^+$  (2)  $\text{Cs}^+$  (3)  $\text{Ca}^+$  (4)  $\text{Mg}^+$
- B-3.** Which one of the following is the smallest in size ?  
 (1)  $\text{N}^{3-}$  (2)  $\text{O}^{2-}$  (3)  $\text{F}^-$  (4)  $\text{Na}^+$
- B-4.** In which pair, the second atom is larger than first :  
 (1) Br, Cl (2) Na, Mg (3) Sr, Ca (4) N, P
- B-5.** Which of the following order of radii is correct ?  
 (1)  $\text{Li} < \text{Be} < \text{Mg}$  (2)  $\text{O}^+ < \text{O}^{2-} < \text{N}^{3-}$  (3)  $\text{O} < \text{F} < \text{Ne}$  (4)  $\text{Na}^+ > \text{F}^- > \text{O}^{2-}$
- B-6.** Which of the following is not isoelectronic series ?  
 (1)  $\text{Cl}^-$ ,  $\text{P}^{3-}$ , Ar (2)  $\text{N}^{3-}$ , Ne,  $\text{Mg}^{2+}$  (3)  $\text{B}^{3+}$ , He,  $\text{Li}^+$  (4)  $\text{F}^-$ ,  $\text{S}^{2-}$ ,  $\text{N}^{3-}$
- B-7.** Which of the following order of atomic / ionic radius is not correct ?  
 (1)  $\text{I}^- > \text{I} > \text{I}^+$  (2)  $\text{N}^{3-} < \text{N}^{5+}$  (3)  $\text{P}^{5+} < \text{P}^{3+}$  (4)  $\text{Li} > \text{Be} > \text{B}$
- B-8.** Among  $\text{Cl}^-$ ,  $\text{F}^-$ ,  $\text{Br}^-$  and  $\text{I}^-$  the correct order of increasing atomic radii is :  
 (1)  $\text{I}^- > \text{F}^- > \text{Cl}^- > \text{Br}^-$  (2)  $\text{I}^- > \text{Br}^- > \text{Cl}^- > \text{F}^-$  (3)  $\text{F}^- > \text{Br}^- > \text{Cl}^- > \text{I}^-$  (4)  $\text{F}^- > \text{Cl}^- > \text{Br}^- > \text{I}^-$
- B-9.** Atomic radii of F & Ne in Angstrom are respectively given by :  
 (1) 0.72, 1.60 (2) 1.60, 1.60 (3) 0.72, 0.72 (4) 1.60, 0.72.
- B-10.** Match list – I with list – II and select the correct answer using the codes given below –

List – I  
Ion

- (I)  $\text{Li}^+$   
 (II)  $\text{Na}^+$   
 (III)  $\text{Br}^-$   
 (IV)  $\text{I}^-$

List – II  
Radius

- (a) 216  
 (b) 195  
 (c) 60  
 (d) 95

**Codes :**

- |     | I | II | III | IV |     | I | II | III | IV |
|-----|---|----|-----|----|-----|---|----|-----|----|
| (1) | a | b  | d   | c  | (2) | b | c  | a   | d  |
| (3) | c | d  | b   | a  | (4) | d | c  | b   | a  |

- B-11.** Select correct statement about radius of an atom.  
 (1) Values of Vander Waal's radii are larger than those of covalent radii because the Vander Waal's forces are much weaker than the forces operating between atoms in a covalently bonded molecule.  
 (2) The metallic radii are smaller than the vander Wall's radii, since the bonding forces in the metallic crystal lattice are much stronger than the vander Wall's forces.  
 (3) Both are correct.  
 (4) None is correct.
- B-12.** The smallest among these species is :  
 (1) lithium (2) lithium ion (3) hydrogen (4) helium
- B-13.** In the ions  $P^{3-}$ ,  $S^{2-}$  and  $Cl^-$  the increasing order of size is :  
 (1)  $Cl^- < S^{2-} < P^{3-}$  (2)  $P^{3-} < S^{2-} < Cl^-$  (3)  $S^{2-} < Cl^- < P^{3-}$  (4)  $S^{2-} < P^{3-} < Cl^-$
- B-14.** Among O,C,F,Cl,Br the correct order of increasing atomic radii is :  
 (1)  $F < O < C < Cl < Br$  (2)  $F < C < O < Br < Cl$   
 (3)  $F < Cl < Br < O < C$  (4)  $C < O < F < Cl < Br$
- B-15.** Which series of elements should have nearly the same atomic radii ?  
 (1) Na, K, Rb (2) Fe, Co, Ni (3) Li, Be, B (4) F, Cl, Br
- B-16.** Which of the following has largest radius -  
 (1)  $1s^2, 2s^2, 2p^6, 3s^2$  (2)  $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$   
 (3)  $1s^2, 2s^2, 2p^6, 3s^2, 3p^2$  (4)  $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$
- B-17.** Arrange the following in order of increasing atomic radii Na, Si, Al, Ar  
 (1)  $Na < Si < Al < Ar$  (2)  $Si < Al < Na < Ar$  (3)  $Ar < Al < Si < Na$  (4)  $Na < Al < Si < Ar$
- B-18.** Consider the isoelectronic series :  
 $K, S^{2-}, Cl^-$  and  $Ca^{2+}$ , the radii of the ions decrease as -  
 (1)  $Ca^{2+} > K^+ > Cl^- > S^{2-}$  (2)  $Cl^- > S^{2-} > K^+ > Ca^{2+}$   
 (3)  $S^{2-} > Cl^- > K^+ > Ca^{2+}$  (4)  $K^+ > Ca^{2+} > S^{2-} > Cl^-$

### Section (C) : Ionisation Energy

- C-1.** In a given energy level, the order of penetration effect of different orbitals is :  
 (1)  $f < d < p < s$  (2)  $s = p = d = f$  (3)  $s < p < d < f$  (4)  $p > s > d > f$
- C-2.** The ionization energy will be higher when the electron is removed from ..... if other factors being equal.  
 (1) s-orbital (2) p-orbital (3) d-orbital (4) f-orbital
- C-3.** The first ionisation energy in eV of N & O are respectively given by :  
 (1) 14.6, 13.6 (2) 13.6, 14.6 (3) 13.6, 13.6 (4) 14.6, 14.6
- C-4.** Which electronic configuration of neutral atoms will have the highest first ionisation energy ?  
 (1)  $1s^2 2s^2 2p^4$  (2)  $1s^2 2s^2 2p^3$  (3)  $1s^2 2s^2 2p^2$  (4)  $1s^2 2s^2 2p^1$
- C-5.** The first ionization energy is smallest for the atom with electronic configuration :  
 (1)  $ns^2 np^6$  (2)  $ns^2 np^4$  (3)  $ns^2 np^5$  (4)  $ns^2 np^3$
- C-6.** The second ionization energies of elements are always higher than their first ionization energies because:  
 (1) the cation is smaller than its parent atom.  
 (2) it is easier to remove electron from cation.  
 (3) ionization is an endothermic process.  
 (4) cation formed always have stable half filled or completely filled valence shell electron configuration.
- C-7.** The first ionisation energy will be maximum for :  
 (1) Be (2) He (3) Li (4) Fe
- C-8.** Which of the following is incorrect ?  
 (1) 1<sup>st</sup> ionisation energy of Li > 1<sup>st</sup> ionisation energy of Be  
 (2) 1<sup>st</sup> ionisation energy of Li < 1<sup>st</sup> ionisation energy of Be  
 (3) 1<sup>st</sup> ionisation energy of Li > 1<sup>st</sup> ionisation energy of Na  
 (4) 1<sup>st</sup> ionisation energy of He > 1<sup>st</sup> ionisation energy of Ne

- C-9.** The first ionisation energy of Na, Mg, Al and Si are in the order :  
 (1) Na < Mg < Al < Si    (2) Na < Al < Mg < Si    (3) Mg < Na < Al < Si    (4) Si < Mg < Al < Na
- C-10.** Correct orders of 1<sup>st</sup> ionisation energies are :  
 (i) Li < B < Be < C    (ii) O < N < F    (iii) Be < N < Ne  
 (1) (i), (ii)    (2) (ii),(iii)    (3) (i),(iii)    (4) (i),(ii),(iii)
- C-11.** Ionisation energy :  
 (1) increases with an increase in atomic radii.  
 (2) is independent of atomic radii.  
 (3) decreases with an increase in atomic radii.  
 (4) remains constant with an increase or decrease in atomic radii.
- C-12.** Which among the following element have lowest value of first ionisation energy ?  
 (1) Pb    (2) Sn    (3) Si    (4) C
- C-13.** Which of the following isoelectronic ion has the highest 1<sup>st</sup> ionization energy ?  
 (1) Na<sup>+</sup>    (2) F<sup>-</sup>    (3) Mg<sup>2+</sup>    (4) O<sup>2-</sup>.
- C-14.** Which of the following orders are correct for the ionization energies ?  
 (i) Ba < Sr < Ca    (ii) S<sup>2-</sup> < S < S<sup>2+</sup>    (iii) C < O < N    (iv) Mg < Al < Si  
 (1) i, ii and iv    (2) i, iii and iv    (3) i, ii and iii    (4) i, ii, iii and iv
- C-15.** The first ionization energy of O is less than that of N because :  
 (1) the former is more electronegative than later one.  
 (2) the former has partially filled electron configuration while later one has half filled electron configuration.  
 (3) the former is bigger than that of later one.  
 (4) the former has less electron affinity than that of later one.
- C-16.** Among the following elements (whose electronic configuration is give below) the one having the highest ionisation energy is :  
 (1) [Ne] 3s<sup>2</sup> 3p<sup>3</sup>    (2) [Ne] 3s<sup>2</sup> 3p<sup>4</sup>    (3) [Ne] 3s<sup>2</sup> 3p<sup>5</sup>    (4) [Ne] 3s<sup>2</sup>
- C-17.** The values of IE(I), IE(II), IE(III), and IE(IV), of an atom are respectively 7.5 eV, 25.6 eV, 48.6 eV and 170.6 eV. The electronic configuration of the atom will be :  
 (1) 1s<sup>2</sup>, 2s<sup>2</sup>, 2p<sup>6</sup>, 3s<sup>1</sup>    (2) 1s<sup>2</sup>, 2s<sup>2</sup>, 2p<sup>6</sup>, 3s<sup>2</sup>, 3p<sup>1</sup>  
 (3) 1s<sup>2</sup>, 2s<sup>2</sup>, 2p<sup>6</sup>, 3s<sup>2</sup>, 3p<sup>3</sup>    (4) 1s<sup>2</sup>, 2s<sup>2</sup>, 2p<sup>6</sup>, 3s<sup>2</sup>
- C-18.** IE<sub>1</sub> and IE<sub>2</sub> of Mg are 178 and 348 K, cal mol<sup>-1</sup>. The enthalpy required for the reaction Mg → Mg<sup>2+</sup> + 2e<sup>-</sup> is  
 (1) +170 K. cal    (2) +526 K.cal    (3) -170 K.cal    (4) -526 K.cal
- C-19.** Which of the following is the **correct** order of ionisation energy ?  
 (i) Be<sup>+</sup> > Be    (ii) Be > Be<sup>+</sup>    (iii) C > Be    (iv) B > Be  
 (1) ii, iii    (2) iii, iv    (3) i, iii    (4) None of these
- C-20.** In which of the following configuration, there will be large difference between second and third ionisation energies ?  
 (1) 1s<sup>2</sup>, 2s<sup>2</sup>, 2p<sup>6</sup>, 3s<sup>1</sup>    (2) 1s<sup>2</sup>, 2s<sup>2</sup>, 2p<sup>6</sup>, 3s<sup>2</sup>, 3p<sup>1</sup>  
 (3) 1s<sup>2</sup>, 2s<sup>2</sup>, 2p<sup>6</sup>, 3s<sup>2</sup>, 3p<sup>6</sup>    (4) 1s<sup>2</sup>, 2s<sup>2</sup>, 2p<sup>6</sup>, 3s<sup>2</sup>
- C-21.** The IP<sub>1</sub>, IP<sub>2</sub>, IP<sub>3</sub>, IP<sub>4</sub> and IP<sub>5</sub> of an element are 7.1, 14.3, 34.5, 46.8, 162.2 eV respectively. The element is likely to be  
 (1) Na    (2) Si    (3) F    (4) Ca
- C-22.** With reference to ionisation potential which one of the following sets is correct  
 (1) Li > K > B    (2) B > Li > K    (3) Cs > Li > K    (4) Cs < Li < K

## Section (D) : Electron gain enthalpy (Electron affinity)

- D-1.** Electron affinity is a :  
 (1) Relative strength to attract the shared electron pair  
 (2) Necessary energy required to remove the electron from the ultimate orbit  
 (3) Energy released when an electron is added to the outermost shell  
 (4) Energy released when an electron is added to the inner shell

- D-2.** Second electron affinity of an element is :  
 (1) Always exothermic (2) Endothermic for few elements  
 (3) Exothermic for few elements (4) Always endothermic
- D-3.** The order of first electron affinity of O, S and Se is :  
 (1)  $O > S > Se$  (2)  $S > Se > O$  (3)  $Se > O > S$  (4)  $S > O > Se$
- D-4.** For electron affinity of halogens which of the following is correct ?  
 (1)  $Br > F$  (2)  $F > Cl$  (3)  $Br < Cl$  (4)  $F^- > I$
- D-5.** In which case the energy released is minimum?  
 (1)  $Cl \rightarrow Cl^-$  (2)  $P \rightarrow P^-$  (3)  $N \rightarrow N^-$  (4)  $C \rightarrow C^-$
- D-6.** The correct order of electron affinity is :  
 (1)  $Be < B < C < N$  (2)  $Be < N < B < C$  (3)  $N < Be < C < B$  (4)  $N < C < B < Be$
- D-7.** The electron affinity values for the halogens shown the following trend :  
 (1)  $F < Cl > Br > I$  (2)  $F < Cl < Br < I$  (3)  $F > Cl < Br < I$  (4)  $F > Cl > Br > I$
- D-8.** In the process  $Cl(g) + e^- \xrightarrow{\Delta H} Cl^-(g)$ ,  $\Delta H$  is :  
 (1) positive (2) negative (3) zero (4) none of these
- D-9.** Which of the following is affected by the stable electron configuration of an atom ?  
 (i) Electronegativity (ii) Ionisation energy (iii) Electron affinity  
 Correct answer is :  
 (1) only electronegativity (2) only ionisation potential  
 (3) electron affinity and ionisation energy both (4) all of the above
- D-10.** Which is a true statement ?  
 (1) Larger is the value of ionisation energy easier is the formation of cation.  
 (2) Larger is the value of electron affinity easier is the formation of anion.  
 (3) Larger is the value of ionisation energy as well as electron affinity the smaller is the electronegativity of atom.  
 (4) Larger is the  $Z_{eff}$  larger is the size of atom.
- D-11.** In which of the following processes energy is liberated ?  
 (1)  $O^- + e^- \longrightarrow O^{2-}$  (2)  $Cl \longrightarrow Cl^+ + e^-$  (3)  $Cl + e^- \longrightarrow Cl^-$  (4)  $Ne + e^- \longrightarrow Ne^-$
- D-12.** Which of the following configuration will have least electron affinity  
 (1)  $ns^2np^5$  (2)  $ns^2np^2$  (3)  $ns^2np^3$  (4)  $ns^2np^4$
- D-13.** Which of the following will have the most negative electron gain enthalpy and which the least negative ?  
 (1) F, Cl (2) Cl, F (3) S, Cl (4) Cl, P
- D-14.** In the formation of a chloride ion, from as isolated gaseous chlorine atom, 3.8 eV energy is released, which would be equal to :  
 (1) Electron affinity of  $Cl^-$  (2) Ionisation potential of Cl  
 (3) Electronegativity of Cl (4) Ionisation potential of  $Cl^-$

### Section (E) : Electronegativity

- E-1.** Following the Mulliken scale, what parameters are required to evaluate electronegativity ?  
 (1) Only electronegativity (2) Only electron affinity  
 (3) Electron affinity and ionization energy (4) Ionic potential and electronegativity
- E-2.** The electronegativity values of C, N, O and F :  
 (1) increase from carbon to fluorine.  
 (2) decrease from carbon to fluorine.  
 (3) increase up to oxygen and is minimum at fluorine.  
 (4) is minimum at nitrogen and then increase continuously.
- E-3.** The electronegativity of the following elements increases in the order :  
 (1)  $C < N < Si < P$  (2)  $N < Si, < C < P$  (3)  $Si < P < C < N$  (4)  $P < Si < N < C$

- E-4.**  $\overset{1}{\text{H}}\text{C}\equiv\overset{2}{\text{C}}-\overset{3}{\text{C}}\text{H}=\overset{4}{\text{C}}\text{H}-\overset{5}{\text{C}}\text{H}_3$  Which carbon atom will show minimum electronegativity :  
 (1) Fifth (2) Third (3) First (4) Second  
**Hint :** Electronegativity  $\propto$  % s character
- E-5.** The correct set of decreasing order of electronegativity is :  
 (1) Li, H, Na (2) Na, H, Li (3) H, Li, Na (4) Li, Na, H
- E-6.** Increasing order of electronegativity is :  
 (1) Bi < P < S < Cl (2) P < Bi < S < Cl (3) C > F > N > O (4) F > O > N > C
- E-7.** Arrange F, C, N, O in the decreasing order of electronegativity :  
 (1) O > F > N > C (2) F > N > C > O (3) C > F > N > O (4) F > O > N > C
- E-8.** The outer most electronic configuration of the most electronegative atom is :  
 (1)  $ns^2 np^5$  (2)  $ns^2 np^6$  (3)  $ns^2 np^4$  (4)  $ns^2 np^3$

### Section (F) : Acidic and Basic Character

- F-1.** Amphoteric behaviour is shown by the oxides of :  
 (1) Al and Ca (2) Pb and Ba (3) Cr and Mg (4) Ge and Al
- F-2.** Which oxide is neutral :  
 (1)  $\text{CO}_2$  (2)  $\text{SO}_2$  (3) NO (4)  $\text{P}_2\text{O}_5$
- F-3.** Predict the nature of hydroxide of an element E–O–H (if difference between electronegativity of E & O is more than 1.4).  
 (1) Acidic (2) Basic (3) Neutral (4) Amphoteric
- F-4.** The order of basic character of given oxides is :  
 (1)  $\text{Na}_2\text{O} > \text{MgO} > \text{CuO} > \text{SiO}_2$  (2)  $\text{MgO} > \text{SiO}_2 > \text{CuO} > \text{Na}_2\text{O}$   
 (3)  $\text{SiO}_2 > \text{MgO} > \text{CuO} > \text{Na}_2\text{O}$  (4)  $\text{CuO} > \text{Na}_2\text{O} > \text{MgO} > \text{SiO}_2$
- F-5.** Amphoteric behaviour is shown by the oxides of :  
 (1) Al and Ca (2) Pb and Ba (3) Cr and Mg (4) Sn and Zn
- F-6.** An element X occurs in short period having configuration  $ns^2 np^1$ . The formula and nature of its oxide is :  
 (1)  $\text{XO}_3$ , basic (2)  $\text{XO}_3$ , acidic (3)  $\text{X}_2\text{O}_3$ , amphoteric (4)  $\text{X}_2\text{O}_3$ , basic

## Exercise # 2

### OBJECTIVE QUESTIONS

#### Single choice type

1. The number of d-electrons in  $\text{Fe}^{2+}$  is not equal to that of the :  
 (1) d-electrons in Fe (Atomic number = 26) (2) p-electrons in Ne (Atomic number = 10)  
 (3) p-electrons in  $\text{Cl}^-$  (Atomic number = 17) (4) d-electrons in  $\text{Co}^{3+}$  (Atomic number = 27)
2. Which represents alkali metals (i.e. 1 group metals) based on  $(\text{IE})_1$  and  $(\text{IE})_2$  values ?
- |     |   | $(\text{IE})_1$ | $(\text{IE})_2$ |     | $(\text{IE})_1$ | $(\text{IE})_2$ |     |
|-----|---|-----------------|-----------------|-----|-----------------|-----------------|-----|
| (1) | X | 100             | 110             | (2) | Y               | 95              | 120 |
| (3) | Z | 195             | 500             | (4) | M               | 200             | 250 |
3. The successive ionization energies for an unknown element are :  
 $\text{IE}_1 = 899 \text{ kJ/mol}$   
 $\text{IE}_2 = 1757 \text{ kJ/mol}$   
 $\text{IE}_3 = 14,847 \text{ kJ/mol}$   
 $\text{IE}_4 = 17,948 \text{ kJ/mol}$   
 To which family in the periodic table does the unknown element most likely belong ?  
 (1) Carbon family (2) Boron family  
 (3) Alkaline earth metal family (4) Nitrogen family

4. Which one is not correct statement ?  
 (1) IE(I) of He is maximum among all elements.  
 (2)  $E_a(I)$  for noble gases is zero/positive.  
 (3) Electronegativity is maximum for fluorine.  
 (4) IE(I) for nitrogen is less than that of oxygen.
5. As one move down the group from top to bottom then which one among the following will not be observed?  
 (1) Ionisation energy increases (2) Electron affinity decreases  
 (3) Electronegativity decreases (4) Atomic radius increases
6. The first ionisation potential of Al is smaller than that of Mg because :  
 (1) the atomic size of Al > Mg. (2) the atomic size of Al < Mg.  
 (3) Al has one unpaired electron in 3p-orbital. (4) Mg has incompletely filled 3s-orbital.
7. The first ( $IE_1$ ) and second ( $IE_2$ ) ionization energies (kJ/mol) of a few elements designated by Roman numerals are given below. Which of these would be an alkali metal ?
- |         | $IE_1$ | $IE_2$ |        | $IE_1$ | $IE_2$ |
|---------|--------|--------|--------|--------|--------|
| (1) I   | 2372   | 5251   | (2) II | 520    | 7300   |
| (3) III | 900    | 1760   | (4) IV | 680    | 7800   |
8. The first Ionization energy of Li, Be, B & C are in the order :  
 (1) Li < Be > B < C (2) Li > Be > B > C  
 (3) Li < Be < B > C (4) Li > Be > B < C
9. Which of the following relation is correct with respect to first (I) and second (II) ionization energies of sodium and magnesium ?  
 (1)  $I_{Mg} = II_{Na}$  (2)  $I_{Na} > I_{Mg}$  (3)  $II_{Mg} > II_{Na}$  (4)  $II_{Na} > II_{Mg}$
10. The correct order of second ionisation potential for the given element is :  
 (1) C > N > O > F (2) O > N > F > C (3) O > F > N > C (4) F > O > N > C
11. The correct order of decreasing first ionization energy is :  
 (1) Si > Al > Mg > Na (2) Si > Mg > Al > Na  
 (3) Al > Si > Mg > Na (4) none of these
12. Element X, Y and Z have atomic numbers 19, 37 and 55 respectively. Which of the following statements is true :  
 (1) Their ionisation potential would increase with the increasing atomic number  
 (2) 'Y' would have an ionisation potential in between those of 'X' and 'Z'.  
 (3) 'Z' would have the highest ionisation potential  
 (4) 'Y' would have the highest ionisation potential
13. Fluorine has the highest electronegativity among the  $ns^2 np^5$  group on the Pauling scale, but the electron affinity of fluorine is less than that of chlorine because :  
 (1) the atomic number of fluorine is less than that of chlorine.  
 (2) fluorine being the first member of the family behaves in an unusual manner.  
 (3) chlorine can accommodate an electron better than fluorine by utilising its vacant 3d-orbital.  
 (4) small size, high electron density and an increased electron repulsion makes addition of an electron to fluorine less favourable than that in the case of chlorine.
14. The electron gain enthalpies of halogens are as given below.  
 $F = -332$ ,  $Cl = -349$ ,  $Br = -324$ ,  $I = -295$  kJ mol<sup>-1</sup>.  
 The less negative value for F as compared to that of Cl is due to :  
 (1) strong electron-electron repulsions in the compact 2-p sub shell of F.  
 (2) strong electron-electron repulsions in the bigger 3-p sub shell of Cl  
 (3) higher electronegativity value of Cl.  
 (4) higher effective nuclear charge of F.

15. The elements having very high ionization enthalpy but zero electron gain enthalpy is :  
 (1) H (2) F (3) He (4) Be
16. Which one of the following statements is incorrect ?  
 (1) Greater is the nuclear charge, greater is the negative electron gain enthalpy.  
 (2) Nitrogen has almost zero electron gain enthalpy.  
 (3) Electron gain enthalpy decreases from fluorine to iodine in the group.  
 (4) Chlorine has highest electron gain enthalpy.
17. If  $x$ ,  $y$  and  $z$  are electronegativity, ionisation potential and electron-affinity respectively. Then the electron affinity ( $z$ ) in the terms of electronegativity ( $x$ ) and ionisation potential ( $y$ ) will be :  
 (1)  $z = \frac{x+y}{z}$  (2)  $Z = \frac{x-y}{z}$  (3)  $Z = \frac{x^2 - y^2}{2}$  (4)  $z = 2x - y$
18. Which of the following statements are correct ?  
 (i) Generally the radius trend and the ionization energy trend across a period are exact opposites.  
 (ii) Electron affinity values of elements may be exothermic (negative) or endothermic (positive).  
 (iii) The first ionization energy of sulphur is higher than that of phosphorus.  
 (iv)  $\text{Te}^{2-} > \text{I}^- > \text{Cs}^+ > \text{Ba}^{2+}$  represents the correct decreasing order of ionic radii.  
 (1) (i), (iii) and (iv) (2) (ii), (iii) and (iv) (3) (i), (ii) and (iv) (4) (i), (ii) and (iii)
19. The incorrect statement among the following is :  
 (1) density increases across the period from left to right while decreases down the group.  
 (2) ionization energy depends upon the type of orbital (of same energy level) from which electron is being removed.  
 (3) generally electron affinity decreases down the group.  
 (4) moving diagonally, the charge to size ratio remains nearly same for 2 & 3<sup>rd</sup> period elements upto 14<sup>th</sup> group.
20. Which of the following orders is incorrect ?  
 (1)  $\text{F} > \text{N} > \text{C} > \text{Si} > \text{Ga}$  – non –metallic character.  
 (2)  $\text{F} > \text{Cl} > \text{O} > \text{N}$  – oxidising property.  
 (3)  $\text{C} < \text{Si} > \text{P} > \text{N}$  – electron affinity value.  
 (4) None of these.
21. Li resembles Mg due to diagonal relationship which is attributed to :  
 (1) nearly same atomic and ionic size (2) same value of electron affinity  
 (3) penetration of sub-shells (4) identical effective nuclear charge
22. In group moving top to down the atomic density  
 (1) Decreases (2) Remains the same (3) can't be said (4) Generally increases

## Exercise # 3

### AIIMS CORNER

#### ASSERTION / REASONING

**Directions :** Each of these questions contains an Assertion followed by reason. Read them carefully and answer the question on the basis of following options. You have to select the one that best describes the two statements.

- (1) If both assertion and reason are true and reason is the correct explanation of assertion.  
 (2) If both assertion and reason are true but reason is not the correct explanation of assertion.  
 (3) If Assertion is true but reason is false.  
 (4) If both assertion and reason are false.
1. **Assertion :** Caesium is the most electropositive element in the periodic table.  
**Reason :** Caesium has the lowest first ionisation energy among all the elements.
2. **Assertion :** Electron affinity values of the 3<sup>rd</sup> period elements on extreme right of the periodic table except noble gases are generally more than the 2<sup>nd</sup> period element of the same group.  
**Reason :** Due to smaller atomic size of the 2<sup>nd</sup> period element, its electron density increases which oppose the addition of electron.

3. **Assertion** : 'He' has highest ionisation energy among all the elements.  
**Reason** : Electron affinity values of metals are generally low while those of non-metals are high.
4. **Assertion** : In a period noble gas has smallest atomic radius.  
**Reason** : In case of noble gases atomic radii are taken as covalent radii.
5. **Assertion** : The 5<sup>th</sup> period of periodic table contains 18 elements not 32.  
**Reason** :  $n = 5$ ,  $\ell = 0, 1, 2, 3$ . The order in which the energy of available orbitals 4d, 5s and 5p increases is  $5s < 4d < 5p$  and the total number of orbitals available are 9 and thus 18 electrons can be accommodated.
6. **Assertion** : The 4f- and 5f- inner transition series of elements are placed between the d-block elements in the periodic table.  
**Reason** : The 4f- and 5f- inner transition series of elements are placed at the bottom of the periodic table.  
(i) This prevents the undue expansion of the periodic table i.e., maintains its structure.  
(ii) This preserve the principle of classification by keeping elements with similar properties in a single column.
7. **Assertion** : Nitrogen (atomic number 7) has less electron affinity than the oxygen (atomic number 8).  
**Reason** : The magnitude of an element's electron affinity only depends on the element's valence shell electrons configuration.
8. **Assertion** : In general,  $IE_1 < IE_2 < IE_3 \dots\dots$ .  
**Reason** : After the removal of each successive electron, remaining electrons are held more tightly by the nucleus. So removal of next electron becomes difficult.
9. **Assertion** : Second electron affinity of halogens is always positive.  
**Reason** : Fluorine has maximum value of electron affinity.
10. **Assertion** : Sc ( $Z = 21$ ) is placed in d-block elements.  
**Reason** : Last filling electron goes into 3d-subshell.
11. **Assertion** :  $Na^+$  and  $Al^{3+}$  are isoelectronic but the magnitude of ionic radius of  $Al^{3+}$  is less than of  $Na^+$ .  
**Reason** : The magnitude of effective nuclear charge of the outer most shell electrons in  $Al^{3+}$  is greater than that in  $Na^+$ .
12. **Assertion** : The atomic radii of the elements of the oxygen family are smaller than the atomic radii of the corresponding elements of the nitrogen family.  
**Reason** : The members of the oxygen family are more electronegative and thus have lower values of nuclear charge than those of the nitrogen family.
13. **Assertion** : Metallic character of first group metals increases down the group with the decreasing ionisation energy.  
**Reason** : The loss and gain of electron(s) can be used in explaining the reducing and oxidising behaviour of the elements respectively.
14. **Assertion** : Fluorine has a less negative electron affinity than chlorine.  
**Reason** : There is relative greater effectiveness of 2p-electrons in the small F atom to repel the additional electron entering the atom than to 3p-electrons in the larger Cl atom.
15. **Assertion** : Fluorine forms only one oxoacid, HOF.  
**Reason** : Fluorine has small size and high electronegativity.
16. **Assertion** : Li and Mg shows diagonal relationship.  
**Reason** : Li and Mg are diagonal to each other in the periodic table.
17. **Assertion** : The decreasing order of acidic character of  $CO_2$ ,  $N_2O_5$ ,  $SiO_2$  and  $SO_3$  is  $SO_3 > N_2O_5 > CO_2 > SiO_2$ .  
**Reason** : Oxides of the metals are generally basic and oxides of the nonmetals are acidic.

# Exercise # 4

## Level - 1

## Questions of Previous Year of AIPMT/AIIMS/AFMC/RPMT

- Which electronic configuration of an element has abnormally high difference between second and third ionization energy ? [AIPMT-93]  
(1)  $1s^2, 2s^2, 2p^6, 3s^1$  (2)  $1s^2, 2s^2 2p^6, 3s^2 3p^1$  (3)  $1s^2, 2s^2 2p^6, 3s^2 3p^2$  (4)  $1s^2, 2s^2 2p^6, 3s^2$
- One would expect protons to have very large [AIPMT-93]  
(1) charge (2) ionisation potential (3) hydration potential (4) radius
- $\text{Na}^+$ ,  $\text{Mg}^{2+}$ ,  $\text{Al}^{3+}$ , and  $\text{Si}^{4+}$  are isoelectronic. The order of their ionic size is [AIPMT-93]  
(1)  $\text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+} > \text{Si}^{4+}$  (2)  $\text{Na}^+ < \text{Mg}^{2+} > \text{Al}^{3+} > \text{Si}^{4+}$   
(3)  $\text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+} > \text{Si}^{4+}$  (4)  $\text{Na}^+ < \text{Mg}^{2+} > \text{Al}^{3+} < \text{Si}^{4+}$
- If the atomic number of an element is 33, it will be placed in the periodic table in the [AIPMT-93]  
(1) first group (2) third group (3) fifth group (4) seventh group
- In the periodic table from left to right in a period, the atomic volume [AIPMT-93]  
(1) decreases (2) increases (3) remains same (4) first decrease then increases
- One of the characteristic properties of non-metals is that they [AIPMT-93]  
(1) are reducing agents (2) form basic oxides  
(3) form cations by electron gain (4) are electronegative
- The electronic configuration of an element is  $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$ . What is the atomic number of the element, which is present just below the above element in the periodic table ? [AIPMT-95]  
(1) 33 (2) 34 (3) 36 (4) 49
- The element with the atomic number 118, will be [AIPMT-96]  
(1) alkali (2) noble gas (3) lanthanide (4) transition element
- Which one of the following ions will be smallest in size ? [AIPMT-96]  
(1)  $\text{Na}^+$  (2)  $\text{Mg}^{2+}$  (3)  $\text{F}^-$  (4)  $\text{O}^{2-}$
- Which of the following does not represent the correct order of the properties indicated ? [AIPMT-97]  
(1)  $\text{Ni}^{2+} > \text{Cr}^{2+} > \text{Fe}^{2+} > \text{Mn}^{2+}$  (size) (2)  $\text{Sc} > \text{Ti} > \text{Cr} > \text{Mn}$  (size)  
(3)  $\text{Mn}^{2+} > \text{Ni}^{2+} < \text{Co}^{2+} < \text{Fe}^{2+}$  (unpaired electron) (4)  $\text{Fe}^{2+} > \text{Co}^{2+} > \text{Ni}^{2+} > \text{Cu}^{2+}$  (unpaired electron)
- The correct decreasing order of basic strength is :  
(1)  $\text{AsH}_3 > \text{SbH}_3 > \text{PH}_3 > \text{NH}_3$  (2)  $\text{SbH}_3 > \text{AsH}_3 > \text{PH}_3 > \text{NH}_3$   
(3)  $\text{NH}_3 > \text{PH}_3 > \text{AsH}_3 > \text{SbH}_3$  (4)  $\text{PH}_3 > \text{AsH}_3 > \text{SbH}_3 > \text{NH}_3$
- Which of the following elements has maximum electron affinity ? [AIPMT-99]  
(1) Cl (2) Br (3) I (4) F
- Correct order of 1st ionisation potential (IP) among following elements Be, B, C, N, O is [AIPMT-2001]  
(1)  $\text{B} < \text{Be} < \text{C} < \text{O} < \text{N}$  (2)  $\text{B} < \text{Be} < \text{C} < \text{N} < \text{O}$   
(3)  $\text{Be} < \text{B} < \text{C} < \text{N} < \text{O}$  (4)  $\text{Be} < \text{B} < \text{C} < \text{O} < \text{N}$
- Which of the following order is wrong ? [AIPMT-2002]  
(1)  $\text{NH}_3 < \text{PH}_3 < \text{AsH}_3$  – acidic (2)  $\text{Li} < \text{Be} < \text{B} < \text{C}$  – 1st IP  
(3)  $\text{Al}_2\text{O}_3 < \text{MgO} < \text{Na}_2\text{O} < \text{K}_2\text{O}$  – basic (4)  $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Cs}^+$  – ionic radius

15. The ions  $O^{2-}$ ,  $F^-$ ,  $Na^+$ ,  $Mg^{2+}$  and  $Al^{3+}$  are isoelectronic, their ionic radii show : [AIPMT-2003]  
 (1) a significant increase from  $O^{2-}$  to  $Al^{3+}$       (2) a significant decrease from  $O^{2-}$  to  $Al^{3+}$   
 (3) an increase from  $O^{2-}$  to  $F^-$  and then decrease from  $Na^+$  to  $Al^{3+}$   
 (4) a decrease from  $O^{2-}$  to  $F^-$  and then increase from  $Na^+$  to  $Al^{3+}$
16. Ionic radii are [AIPMT-2004]  
 (1) inversely proportional to effective nuclear charge  
 (2) inversely proportional to square of effective nuclear charge  
 (3) directly proportional to effective nuclear charge  
 (4) directly CO
17. Which one of the following arrangements represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species ? [AIPMT-2005]  
 (1)  $Cl < F < S < O$       (2)  $O < S < F < Cl$       (3)  $S < O < Cl < F$       (4)  $F < Cl < O < S$
18. Four successive members of the first row transition elements are listed below with their atomic numbers. Which one of them is expected to have the highest third ionisation enthalpy ? [AIPMT-2005]  
 (1) Vanadium ( $Z = 23$ )    (2) Chromium ( $Z = 24$ )    (3) Iron ( $Z = 26$ )      (4) Manganese ( $Z = 25$ )
19. Which one of the following orders is not in accordance with the property stated against it ? [AIPMT-2006]  
 (1)  $F_2 > Cl_2 > Br_2 > I_2$  : Oxidising power      (2)  $HI > HBr > HCl > HF$  : Acidic property in water  
 (3)  $F_2 > Cl_2 > Br_2 > I_2$  : Electronegativity      (4)  $F_2 > Cl_2 > Br_2 > I_2$  : bond dissociation energy
20. The correct order regarding the electronegativity of hybrid orbitals of carbon is : [AIPMT-2006]  
 (1)  $sp < sp^2 > sp^3$       (2)  $sp < sp^2 < sp^3$       (3)  $sp > sp^2 < sp^3$       (4)  $sp > sp^2 > sp^3$
21. Identify the correct order of the size of the following [AIPMT-2007]  
 (1)  $Ca^{2+} < K^+ < Ar < S^{2-} < Cl^-$       (2)  $Ca^{2+} < K^+ < Ar < Cl^- < S^{2-}$   
 (3)  $Ar < Ca^{2+} < K^+ < Cl^- < S^{2-}$       (4)  $Ca^{2+} < Ar < K^+ < Cl^- < S^{2-}$
22. With which of the following electronic configuration an atom has the lowest ionisation enthalpy ? [AIPMT-2007]  
 (1)  $1s^2, 2s^2 2p^5$       (2)  $1s^2, 2s^2 2p^3$       (3)  $1s^2, 2s^2 2p^5 3s^1$       (4)  $1s^2, 2s^2 2p^6$
23. The correct order of decreasing second ionisation enthalpy of Ti(22), V(23), Cr(24) and Mn(25) is : [AIPMT-2008]  
 (1)  $Cr > Mn > V > Ti$       (2)  $V > Mn > Cr > Ti$       (3)  $Mn > Cr > Ti > V$       (4)  $Ti > V > Cr > Mn$
24. Which one of the elements with the following outer orbital configuration may exhibit the largest number of oxidation states ? [AIPMT-2008]  
 (1)  $3d^3, 4s^2$       (2)  $3d^5, 4s^1$       (3)  $3d^5, 4s^2$       (4)  $3d^2, 4s^2$
25. Amongst the elements with following electronic configurations, which one may have the highest ionisation energy ? [AIPMT-2009]  
 (1)  $[Ne] 3s^2 3p^3$       (2)  $[Ne] 3s^2 3p^2$       (3)  $[Ar] 3d^{10}, 4s^2 4p^3$       (4)  $[Ne] 3s^2 3p^1$
26. Among the elements Ca, Mg, P and Cl, the order of increasing atomic radii is : [AIPMT-2010]  
 (1)  $Mg < Ca < Cl < P$     (2)  $Cl < P < Mg < Ca$     (3)  $P < Cl < Ca < Mg$     (4)  $Ca < Mg < P < Cl$
27. Among the following which one has the highest cation to anion size ratio? [AIPMT-2010]  
 (1) CsI      (2) CsF      (3) LiF      (4) NaF
28. Which of the following ions will exhibit colour in aqueous solutions ? [AIPMT-2010]  
 (1)  $La^{3+}$  ( $Z = 57$ )      (2)  $Ti^{3+}$  ( $Z = 22$ )      (3)  $Lu^{3+}$  ( $Z = 71$ )      (4)  $Sc^{3+}$  ( $Z = 21$ )
29. Which of the following ions has electronic configuration  $[Ar]3d^6$  ? [AIPMT-2010]  
 (1)  $Ni^{3+}$       (2)  $Mn^{3+}$       (3)  $Fe^{3+}$       (4)  $Co^{3+}$

30. Which of the following pairs has the same size ? [AIPMT-2010]  
 (1)  $\text{Fe}^{2+}$ ,  $\text{Ni}^{2+}$                       (2)  $\text{Zr}^{4+}$ ,  $\text{Ti}^{4+}$                       (3)  $\text{Zr}^{4+}$ ,  $\text{Hf}^{4+}$                       (4)  $\text{Zn}^{4+}$ ,  $\text{Hf}^{4+}$
31. The correct order of the decreasing ionic radii among the following is electronic species are : [AIPMT-2010]  
 (1)  $\text{Ca}^{2+} > \text{K}^+ > \text{S}^{2-} > \text{Cl}^-$                       (2)  $\text{Cl}^- > \text{S}^{2-} > \text{Ca}^{2+} > \text{K}^+$   
 (3)  $\text{S}^{2-} > \text{Cl}^- > \text{K}^+ > \text{Ca}^{2+}$                       (4)  $\text{K}^+ > \text{Ca}^{2+} > \text{Cl}^- > \text{S}^{2-}$
32. Which of the following represents the correct order of increasing electron gain enthalpy with negative sign for the elements O, S, F and Cl ? [AIPMT-2010]  
 (1)  $\text{Cl} < \text{F} < \text{O} < \text{S}$                       (2)  $\text{O} < \text{S} < \text{F} < \text{Cl}$                       (3)  $\text{F} < \text{S} < \text{O} < \text{Cl}$                       (4)  $\text{S} < \text{O} < \text{Cl} < \text{F}$
33. Which of the following compounds has the lowest melting point ? [AIPMT-2011]  
 (1)  $\text{CaCl}_2$                       (2)  $\text{CaBr}_2$                       (3)  $\text{CaI}_2$                       (4)  $\text{CaF}_2$
34. What is the value of electron gain enthalpy of  $\text{Na}^+$  if  $\text{IE}_1$  of  $\text{Na} = 5.1 \text{ eV}$  ? [AIPMT-2011]  
 (1)  $-5.1 \text{ eV}$                       (2)  $-10.2 \text{ eV}$                       (3)  $+2.55 \text{ eV}$                       (4)  $+10.2 \text{ eV}$
35. Identify the **wrong** statement in the following : [AIPMT-2012]  
 (1) Amongst isoelectronic species, smaller the positive charge on the cation, smaller is the ionic radius.  
 (2) Amongst isoelectronic species, greater the negative charge on the anion, larger is the ionic radius.  
 (3) Atomic radius of the elements increases as one moves down the first group of the periodic table.  
 (4) Atomic radius of the elements decreases as one moves across from left to right in the 2<sup>nd</sup> period of the periodic table.
36. Hydrogen is produced by the reaction : [AIIMS 2002]  
 (1)  $\text{Na}_2\text{O}_2 + 2\text{HCl}$                       (2)  $\text{Mg} + \text{H}_2\text{O}$                       (3)  $\text{BaO}_2 + \text{HCl}$                       (4)  $\text{H}_2\text{S}_4\text{O}_8 + \text{H}_2\text{O}$
37. For electron affinity of halogens which of the following is correct ? [AIIMS 2004]  
 (1)  $\text{Br} > \text{F}$                       (2)  $\text{F} > \text{Cl}$                       (3)  $\text{Br} > \text{Cl}$                       (4)  $\text{F} > \text{I}$
38. The pair of amphoteric hydroxides is : [AIIMS 2005]  
 (1)  $\text{Al}(\text{OH})_3$ ,  $\text{LiOH}$                       (2)  $\text{Be}(\text{OH})_2$ ,  $\text{Mg}(\text{OH})_2$   
 (3)  $\text{B}(\text{OH})_3$ ,  $\text{Be}(\text{OH})_2$                       (4)  $\text{Be}(\text{OH})_2$ ,  $\text{Zn}(\text{OH})_2$
39. **Assertion** : First ionisation energy for nitrogen is lower than oxygen.  
**Reason** : Across a period effective nuclear charge decreases. [AIIMS 2005]  
 (1) If both assertion and reason are true and reason is the correct explanation of assertion.  
 (2) If both assertion and reason are true but reason is not the correct explanation of assertion.  
 (3) If Assertion is true but reason is false.  
 (4) If both assertion and reason are false.
40. **Assertion** : In the iodometric titration starch is used as an indicator.  
**Reason** : Starch is a polysaccharide. [AIIMS 2006]  
 (1) If both assertion and reason are true and reason is the correct explanation of assertion.  
 (2) If both assertion and reason are true but reason is not the correct explanation of assertion.  
 (3) If Assertion is true but reason is false.  
 (4) If both assertion and reason are false.
41. Consider the following statements : [AIIMS 2012]  
**I.** The radius of an anion is larger than that of the parent atom.  
**II.** The ionization energy generally increases with increasing atomic number in a period.  
**III.** The electronegativity of an element is the tendency of an isolated atom no attract an electron.  
 Which of the above statements is/are correct?  
 (1) I alone                      (2) II alone                      (3) I and II                      (4) II and III

42. The set representing the correct order for first ionisation potential is : [AFMC-2006]  
 (1)  $K > Na > Li$                       (2)  $Be > Mg > Ca$                       (3)  $B > C > N$                       (4)  $Ge > Si > C$
43. Which one of the following elements has the highest ionisation energy ? [AFMC-2007]  
 (1)  $[Ne]3s^23p^1$                       (2)  $[Ne]3s^2, 3p^2$                       (3)  $[Ne]3s^2 3p^3$                       (4)  $[Ar]3s^{10} 4s^2 4p^2$
44. In the following, the element with the highest ionisation energy is : [AFMC-2008]  
 (1)  $[Ne] 3s^2 3p^1$                       (2)  $[Ne] 3s^2 3p^3$                       (3)  $[Ne] 3s^2 3p^2$                       (4)  $[Ne] 3s^2 3p^4$
45. Strongest reducing agent is : [AFMC-2009]  
 (1) K                      (2) Mg                      (3) Al                      (4) Ba
46. The most polar bond is [RPMT, 2005]  
 (1) O–F                      (2) N–Cl                      (3) N–F                      (4) N–N
47. Which of the following electronic configuration represent noble gas ? [RPMT, 2005-06]  
 (1)  $ns^2np^6$                       (2)  $ns^2np^5$                       (3)  $ns^2np^4$                       (4)  $ns^2np^3$
48. Which of the following element shows maximum valency ? [RPMT,2005]  
 (1) Carbon                      (2) Barium                      (3) Nitrogen                      (4) Sulphur
49. The correct sequence of hybridisation of methane, ethene and acetylene is : [RPMT,2005-06]  
 (1)  $sp, sp^2, sp^3$                       (2)  $sp^2, sp^3, sp$                       (3)  $sp^3, sp^2, sp$                       (4)  $sp^3, sp, sp^2$
50. Which among the following elements have lowest value of  $IE^1$  ? [RPMT,2005-06]  
 (1) Pb                      (2) Sn                      (3) Si                      (4) C
51. Fluorine is the best oxidising agent because it has : [RPMT, 2006]  
 (1) highest electron affinity                      (2) highest  $E^{\circ}_{red}$   
 (3) highest  $E^{\circ}_{oxid}$                       (4) lowest electron affinity
52. Which has smallest size ? [RPMT, 2007]  
 (1)  $Na^+$                       (2)  $Mg^{2+}$                       (3) Ne highest                      (4)  $O^{2-}$
53. Element having maximum electron affinity is : [RPMT, 2007]  
 (1)  $F_2$                       (2)  $Cl_2$                       (3)  $Br_2$  highest                      (4)  $I_2$
54. Element having maximum ionization energy [RPMT, 2007]  
 (1) Na                      (2) Li                      (3) K                      (4) Rb
55. Element with atomic number 38, belongs to [RPMT, 2008]  
 (1) II A group and 5th period                      (2) II A group and 2nd period  
 (3) V A group and 2nd period                      (4) III A group and 5th period
56. Which of the following group of transition metals is called coinage metals ? [RPMT, 2008]  
 (1) Cu, Ag, Au                      (2) Ru, Rn, Pd                      (3) Fe, Co, Ni                      (4) Os, IR, Pt
57. The atomic numbers of vanadium (V), chromium (Cr), manganese (Mn) and iron (Fe) are respectively 23,24,25 and 26, Which one of these may be expected to have the highest second ionisation enthalpy ? [RPMT, 2008]  
 (1) Cu, Ag, Au                      (2) Ru, Rn, Pd                      (3) Fe, Co, Ni                      (4) Os, IR, Pt

58. Which one of the following substances has the highest proten affinity ? [RPMT, 2008]  
 (1) ZnO (2) Na<sub>2</sub>O (3) SO<sub>2</sub> (4) B<sub>2</sub>O<sub>3</sub>
59. Outemost configuration for Z = 25, is : [RPMT, 2009]  
 (1) 4s<sup>2</sup>, 3d<sup>5</sup> (2) 5s<sup>2</sup>, 4d<sup>5</sup> (3) 4s<sup>2</sup>, 3d<sup>3</sup>SO<sub>2</sub> (4) 4s<sup>2</sup>, 3d<sup>1</sup>
60. Which of the following has maximum ionization enthalpy. [RPMT, 2010]  
 (1) P (2) N (3) As3d<sup>3</sup>SO<sub>2</sub> (4) ,Sb
61. Atomic number of Ni and Cu are 28 and 29 respectively. Electronic configuration 1s<sup>2</sup>, 2s<sup>2</sup>, 2p<sup>6</sup>3, 3p<sup>6</sup>, 3d<sup>10</sup> shows : [RPMT, 2010]  
 (1) Ni (2) Ni<sup>2+</sup> (3) ,Cu<sup>2+</sup> 3d<sup>3</sup>SO<sub>2</sub> (4) ,Cu<sup>+</sup>

## Level -2

## AIEEE Problems of Previous Years

1. Arrange Ce<sup>+3</sup>, La<sup>+3</sup>, Pm<sup>+3</sup> and Yb<sup>+3</sup> in increasing order of their ionic radii. [AIEEE-2002]  
 (1) Yb<sup>+3</sup> < Pm<sup>+3</sup> < Ce<sup>+3</sup> < La<sup>+3</sup> (2) Ce<sup>+3</sup> < Yb<sup>+3</sup> < Pm<sup>+3</sup> < La<sup>+3</sup>  
 (3) Yb<sup>+3</sup> < Pm<sup>+3</sup> < La<sup>+3</sup> < Ce<sup>+3</sup> (4) Pm<sup>+3</sup> < La<sup>+3</sup> < Ce<sup>+3</sup> < Yb<sup>+3</sup>.
2. According to the periodic law of elements, the variation in properties of elements is related to their : [AIEEE-2003]  
 (1) atomic masses (2) nuclear masses  
 (3) atomic numbers (4) nuclear neutron-proton number
3. Which one of the following is an amphoteric oxide? [AIEEE-2003]  
 (1) ZnO (2) Na<sub>2</sub>O (3) SO<sub>2</sub> (4) B<sub>2</sub>O<sub>3</sub>.
4. Which one of the following groupings represents a collection of isoelectronic species? [AIEEE-2003]  
 (At. nos. : Cs-55, Br-35)  
 (1) Na<sup>+</sup>, Ca<sup>2+</sup>, Mg<sup>2+</sup> (2) N<sup>3-</sup>, F<sup>-</sup>, Na<sup>+</sup> (3) Be, Al<sup>3+</sup>, Cl<sup>-</sup> (4) Ca<sup>2+</sup>, Cs<sup>+</sup>, Br.
5. Which one of the following substances has the highest proton affinity? [AIEEE-2003]  
 (1) H<sub>2</sub>O (2) H<sub>2</sub>S (3) NH<sub>3</sub> (4) PH<sub>3</sub>.
6. Which one of the following ions has the highest value of ionic radius ? [AIEEE-2004, 3/225]  
 (1) Li<sup>+</sup> (2) B<sup>3+</sup> (3) O<sup>2-</sup> (4) F<sup>-</sup>
7. The formation of the oxide ion O<sup>2-</sup><sub>(g)</sub> requires first an exothermic and then an endothermic step as shown below :  

$$\text{O}_{(g)} + e^- = \text{O}^-_{(g)} ; \Delta H^\circ = -142 \text{ kJmol}^{-1}$$

$$\text{O}^-_{(g)} + e^- = \text{O}^{2-}_{(g)} ; \Delta H^\circ = 844 \text{ kJmol}^{-1}$$
 This is because : [AIEEE-2004, 3/225]  
 (1) oxygen is more electronegative.  
 (2) oxygen has high electron affinity.  
 (3) O<sup>-</sup> ion will tend to resist the addition of another electron.  
 (4) O<sup>-</sup> ion has comparatively larger size than oxygen atom.
8. Among Al<sub>2</sub>O<sub>3</sub>, SiO<sub>2</sub>, P<sub>2</sub>O<sub>3</sub> and SO<sub>2</sub> the correct order of acid strength is : [AIEEE-2004, 3/225]  
 (1) SO<sub>2</sub> < P<sub>2</sub>O<sub>3</sub> < SiO<sub>2</sub> < Al<sub>2</sub>O<sub>3</sub> (2) SiO<sub>2</sub> < SO<sub>2</sub> < Al<sub>2</sub>O<sub>3</sub> < P<sub>2</sub>O<sub>3</sub>  
 (3) Al<sub>2</sub>O<sub>3</sub> < SiO<sub>2</sub> < SO<sub>2</sub> < P<sub>2</sub>O<sub>3</sub> (4) Al<sub>2</sub>O<sub>3</sub> < SiO<sub>2</sub> < P<sub>2</sub>O<sub>3</sub> < SO<sub>2</sub>
9. Which of the following oxides is amphoteric in nature ? [AIEEE-2005, 1 1/2/225]  
 (1) CaO (2) CO<sub>2</sub> (3) SiO<sub>2</sub> (4) SnO<sub>2</sub>
10. In which of the following arrangements the order is NOT according to the property indicated against it ? [AIEEE-2005, 3/225]  
 (1) Al<sup>3+</sup> < Mg<sup>2+</sup> < Na<sup>+</sup> < F<sup>-</sup> – increasing ionic size  
 (2) B < C < N < O – increasing first ionisation enthalpy  
 (3) I < Br < F < Cl – increasing electron gain enthalpy (with negative sign)  
 (4) Li < Na < K < Rb – increasing metallic radius

11. Which of the following factors may be regarded as the main cause of lanthanide contraction ?  
 (1) Greater shielding of 5d electrons by 4f electrons. [AIEEE 2005, 4<sup>1/2</sup>/225]  
 (2) Poorer shielding of 5d electron by 4f electrons.  
 (3) Effective shielding of one of 4f electrons by another in the sub-shell.  
 (4) Poor shielding of one of 4f electron by another in the sub-shell.
12. The lanthanide contraction is responsible for the fact that : [AIEEE-2005, 3/225]  
 (1) Zr and Y have about the same radius (2) Zr and Nb have similar oxidation state  
 (3) Zr and Hf have about the same radius (4) Zr and Zn have same oxidation state.
13. The increasing order of the first ionization enthalpies of the elements B, P, S and F (lowest first) is :  
 [AIEEE-2006, 4/220]  
 (1)  $F < S < P < B$  (2)  $P < S < B < F$  (3)  $B < P < S < F$  (4)  $B < S < P < F$
14. Which of the following statements is true ? [AIEEE-2006, 4/220]  
 (1)  $H_3PO_3$  is a stronger acid than  $H_2SO_3$ . (2) In aqueous medium, HF is a stronger acid than HCl.  
 (3)  $HClO_4$  is a weaker acid than  $HClO_3$ . (4)  $HNO_3$  is a stronger acid than  $HNO_2$ .
15. Lanthanoid contraction is caused due to : [AIEEE-2006, 4/220]  
 (1) the appreciable shielding on outer electrons by 4f electrons from the nuclear charge  
 (2) the appreciable shielding on outer electrons by 5f electrons from the nuclear charge  
 (3) the same effective nuclear charge from Ce to Lu  
 (4) the imperfect shielding on outer electrons by 4f electrons from the nuclear charge
16. The stability of dihalides of Si, Ge, Sn and Pb increases steadily in the sequence. [AIEEE-2007, 3/120]  
 (1)  $SiX_2 \ll GeX_2 \ll SnX_2 \ll PbX_2$  (2)  $PbX_2 \ll SnX_2 \ll GeX_2 \ll SiX_2$   
 (3)  $GeX_2 \ll SiX_2 \ll SnX_2 \ll PbX_2$  (4)  $SiX_2 \ll GeX_2 \ll PbX_2 \ll SnX_2$
17. The set representing the correct order of ionic radius is : [AIEEE-2009, 4/144]  
 (1)  $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$  (2)  $Li^+ > Na^+ > Mg^{2+} > Be^{2+}$   
 (3)  $Mg^{2+} > Be^{2+} > Li^+ > Na^+$  (4)  $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$
18. In which of the following arrangements, the sequence is not strictly according to the property written against it ? [AIEEE-2009, 4/144]  
 (1)  $HF < HCl < HBr < HI$  : increasing acid strength  
 (2)  $NH_3 < PH_3 < AsH_3 < SbH_3$  : increasing basic strength  
 (3)  $B < C < O < N$  : increasing first ionization enthalpy  
 (4)  $CO_2 < SiO_2 < SnO_2 < PbO_2$  : increasing oxidising power
19. The correct sequence which shows decreasing order of the ionic radii of the elements is : [AIEEE-2010, 4/144]  
 (1)  $Al^{3+} > Mg^{2+} > Na^+ > F^- > O^{2-}$  (2)  $Na^+ > Mg^{2+} > Al^{3+} > O^{2-} > F^-$   
 (3)  $Na^+ > F^- > Mg^{2+} > O^{2-} > Al^{3+}$  (4)  $O^{2-} > F^- > Na^+ > Mg^{2+} > Al^{3+}$
20. The outer electron configuration of Gd (Atomic No : 64) is : [AIEEE 2011 (Cancelled), 4/120]  
 (1)  $4f^3 5d^5 6s^2$  (2)  $4f^8 5d^0 6s^2$  (3)  $4f^4 5d^4 6s^2$  (4)  $4f^7 5d^1 6s^2$
21. Which one of the following orders presents the correct sequence of the increasing basic nature of the given oxides ? [AIEEE 2011 (Cancelled), 4/120]  
 (1)  $Al_2O_3 < MgO < Na_2O < K_2O$  (2)  $MgO < K_2O < Al_2O_3 < Na_2O$   
 (3)  $Na_2O < K_2O < MgO < Al_2O_3$  (4)  $K_2O < Na_2O < Al_2O_3 < MgO$
22. The correct order of electron gain enthalpy with negative sign of F, Cl, Br and I, having atomic number 9, 17, 35 and 53 respectively, is:  
 (1)  $F > Cl > Br > I$  (2)  $Cl > F > Br > I$  (3)  $Br > Cl > I > F$  (4)  $I > Br > Cl > F$
23. The increasing order of the ionic radii of the given isoelectronic species is : [AIEEE-2012, 4/144]  
 (1)  $Cl^-, Ca^{2+}, K^+, S^{2-}$  (2)  $S^{2-}, Cl^-, Ca^{2+}, K^+$   
 (3)  $Ca^{2+}, K^+, Cl^-, S^{2-}$  (4)  $K^+, S^{2-}, Ca^{2+}, Cl^-$

24. Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar ? [JEE Mains-2013, 4/120]
- (1) Ca < S < Ba < Se < Ar (2) S < Se < Ca < Ba < Ar  
 (3) Ba < Ca < Se < S < Ar (4) Ca < Ba < S < Se < Ar
25. The first ionisation potential of Na is 5.1 eV. The value of electron gain enthalpy of Na<sup>+</sup> will be : [JEE Mains-2013, 4/120]
- (1) -2.55 eV (2) -5.1 eV  
 (3) -10.2 eV (4) +2.55 eV

## Level -3

## Objective Questions of Previous Years of IIT-JEE

1. Which one of the following oxides is neutral ? [JEE-1996, 2/200]
- (A) CO (B) SnO<sub>2</sub>  
 (C) ZnO (D) SiO<sub>2</sub>
2. The incorrect statement among the following is : [JEE- 1997(Cancelled), 2/200]
- (A) the first ionization energy of Al is less than first ionization energy of Mg.  
 (B) the second ionization energy of Mg is greater than second ionization energy of Na.  
 (C) the first ionization energy of Na is less than first ionization energy of Mg.  
 (D) the third ionization energy of Mg is greater than third ionization energy of Al.
3. Property of the alkaline earth metals that increases with their atomic number is : [JEE-1997(Cancelled), 2/200]
- (A) ionisation energy (B) solubility of their hydroxides  
 (C) solubility of their sulphates (D) electronegativity
4. Arrange the following ions in order of their increasing size : Li<sup>+</sup>, Mg<sup>2+</sup>, K<sup>+</sup>, Al<sup>3+</sup>. [JEE-1997, 1/100]
5. Fill in the blanks :  
 Compounds that formally contain Pb<sup>4+</sup> are easily reduced to Pb<sup>2+</sup>. The stability of the lower oxidation state is due to .....
6. Assertion : F atom has a less negative electron affinity than Cl atom. [JEE-1998, 2/200]  
 Reason : Additional electrons are repelled more effectively by 3p electrons in Cl atom than by 2p electrons in F atom.
- (A) Both Assertion and Reason are true, and Reason is the correct explanation of Assertion.  
 (B) Both Assertion and Reason are true, but Reason is not correct explanation of Assertion.  
 (C) Assertion is true but Reason is false.  
 (D) Assertion is false but Reason is true.
- 7.\* Ionic radii of : [JEE-1999, 3/200]
- (A) Ti<sup>4+</sup> < Mn<sup>7+</sup> (B) <sup>35</sup>Cl<sup>-</sup> < <sup>37</sup>Cl<sup>-</sup> (C) K<sup>+</sup> > Cl<sup>-</sup> (D) P<sup>3+</sup> > P<sup>5+</sup>
8. Give reason for the following in one or two sentences only. [JEE-1999, 2/200]  
 CrO<sub>3</sub> is an acid anhydride.
9. The correct order of radii is : [JEE-2000, 1/35]
- (A) N < Be < B (B) F<sup>-</sup> < O<sup>2-</sup> < N<sup>3-</sup>  
 (C) Na < Li < K (D) Fe<sup>3+</sup> < Fe<sup>2+</sup> < Fe<sup>4+</sup>

10. The correct order of acidic strength is : [JEE-2000, 1/35]  
 (A)  $\text{Cl}_2\text{O}_7 > \text{SO}_3 > \text{P}_4\text{O}_{10}$  (B)  $\text{CO}_2 > \text{N}_2\text{O}_5 > \text{SO}_3$   
 (C)  $\text{Na}_2\text{O} > \text{MgO} > \text{Al}_2\text{O}_3$  (D)  $\text{K}_2\text{O} > \text{CaO} > \text{MgO}$
11. **Assertion** : The first ionization energy of Be is greater than that of B. [JEE-2000, 1/35]  
**Reason** : 2p orbital is lower in energy than 2s.  
 (A) Both Assertion and Reason are true and Reason is the correct explanation of Assertion.  
 (B) Both Assertion and Reason are true but Reason is not correct explanation of Assertion.  
 (C) Assertion is true but Reason is false.  
 (D) Assertion is false but Reason is true.
12. The set representing the correct order of first ionization potential is : [JEE-2001, 1/35]  
 (A)  $\text{K} > \text{Na} > \text{Li}$  (B)  $\text{Be} > \text{Mg} > \text{Ca}$   
 (C)  $\text{B} > \text{C} > \text{N}$  (D)  $\text{Ge} > \text{Si} > \text{C}$
13. The set with correct order of acidity is : [JEE-2001, 1/35]  
 (A)  $\text{HClO} < \text{HClO}_2 < \text{HClO}_3 < \text{HClO}_4$  (B)  $\text{HClO}_4 < \text{HClO}_3 < \text{HClO}_2 < \text{HClO}$   
 (C)  $\text{HClO} < \text{HClO}_4 < \text{HClO}_3 < \text{HClO}_2$  (D)  $\text{HClO}_4 < \text{HClO}_2 < \text{HClO}_3 < \text{HClO}$
14. Identify the correct order of acidic strengths of  $\text{CO}_2$ ,  $\text{CuO}$ ,  $\text{CaO}$ ,  $\text{H}_2\text{O}$  : [JEE-2002, 3/90]  
 (A)  $\text{CaO} < \text{CuO} < \text{H}_2\text{O} < \text{CO}_2$  (B)  $\text{H}_2\text{O} < \text{CuO} < \text{CaO} < \text{CO}_2$   
 (C)  $\text{CaO} < \text{H}_2\text{O} < \text{CuO} < \text{CO}_2$  (D)  $\text{H}_2\text{O} < \text{CO}_2 < \text{CaO} < \text{CuO}$
15. Identify the least stable ion amongst the following : [JEE-2002, 3/90]  
 (A)  $\text{Li}^-$  (B)  $\text{Be}^-$  (C)  $\text{B}^-$  (D)  $\text{C}^-$
16. Arrange the following oxides in the increasing order of Bronsted basicity : [JEE-2004, 2/60]  
**Hint** : Consider Bronsted basicity to be simply basic strength.  
 $\text{Cl}_2\text{O}_7$ ,  $\text{BaO}$ ,  $\text{SO}_3$ ,  $\text{CO}_2$ ,  $\text{B}_2\text{O}_3$ .
17. **Statement-1** :  $\text{Pb}^{4+}$  compounds are stronger oxidizing agents than  $\text{Sn}^{4+}$  compounds. [JEE-2008, 3/163]  
**Statement-2** : The higher oxidation states for the group 14 elements are more stable for the heavier members of the group due to 'inert pair effect'.  
 (A) Statement-1 is True, Statement-2 is True; Statement-2 is a correct explanation for Statement-1.  
 (B) Statement-1 is True, Statement-2 is True; Statement-2 is NOT a correct explanation for Statement-1  
 (C) Statement-1 is True, Statement-2 is False  
 (D) Statement-1 is False, Statement-2 is True

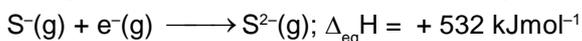
# Answers

## BOARD LEVEL EXERCISE : HINT & SOLUTIONS

- The basic theme of organisation of elements in the periodic table is **to simplify and systematize the study of the properties of all elements and millions of their compounds.**
- Mendeleev Periodic Law** states that the **physical and chemical properties of the elements are a periodic function of their atomic weights** while **Modern Periodic Law** states that **physical and chemical properties of elements are a periodic function of their atomic numbers**. Thus, the basic difference in approach between Mendeleev's Periodic Law and Modern Periodic Law is the **change in basis of classification of elements from atomic weight to atomic number.**
- In the **fourth period**, the **filling up of only 4s, 3d and 4p-orbitals** occur. Therefore, in this period, there are only **two s block, ten d-block and six p-block elements.** (½ Mark) Since fourth period starts with  $Z = 19$  and ends at  $Z = 36$ , therefore, elements with  $Z = 19$  and  $Z = 20$  are s-block elements. The next ten elements with  $Z = 21$  to  $30$  are d-block elements and belong to groups 3, 4, 5, 6, 7, 8, 9, 10, 11 & 12. Therefore, the element which will lie in eight group will have  $Z = 20 + 6 = 26$ . (½ Mark)
- The roots for 1, 0 and 9 are un, nil and enn respectively. Hence, the **symbol** for the element is **Une** (½ Mark) and **IUPAC name** for element is **Unnilennium.** (½ Mark)
- (i) Lawrencium ( $Z = 103$ ) and Berkelium ( $Z = 97$ ) (½ Mark)  
(ii) Seaborgium ( $Z = 106$ ) (½ Mark)
- Elements in the same period** have the same outermost shell but **different outermost electronic configuration** and hence have different physical and chemical properties. (½ Mark)  
However, **elements in the same group** have **similar outermost electronic configuration** and hence have similar physical and chemical properties. (½ Mark)
- No, the **oxidation state** of Be is **+2** [ $\because 1 \times x + 3 \times 0 + 1 \times (-1) = +1$ ] (½ Mark), and the **covalency is 4** (since four bonds are formed by Be). (½ Mark)
- Metallic character** (tendency to lose electron) **increases down a group and decreases along a period**, as we move from left to right. Hence the order of increasing metallic character is : **P < Si < Mg < Ca < K.**
- Electronic configuration :  $[\text{Rn}] 7s^2 5f^{14} 6d^{10} 7p^2$   
Period = 7th (maximum value of principal quantum number involved in electronic configuration)  
Block : p-block (orbital in which last electron enters)  
Group : 14 (for p-block elements, Group number = Number of ns electrons + Number of np electrons)
- s-Block elements :  $ns^{1-2}$  (½ Mark)  
p-Block elements :  $ns^2 np^{1-6}$  (½ Mark)  
d-Block elements :  $(n-1) d^{1-10} ns^{0-2}$  (½ Mark)  
f-Block elements :  $(n-2) f^{1-14} (n-1) d^{0-1} ns^2$ . (½ Mark)
- (i) Electronic configuration must be  $ns^2 np^4$ . So, **16<sup>th</sup> group elements. Eg. Oxygen** (½ Mark)  
(ii) Electronic configuration must be  $ns^1$ . So, **1<sup>st</sup> group elements. Eg. Sodium** (½ Mark)  
(iii) Electronic configuration must be  $ns^2 np^3$ . So, **15<sup>th</sup> group elements. Eg. Nitrogen** (½ Mark)  
(iv) **1<sup>st</sup> group** consists of H (gas & non metal), Li (solid & metal), Na (solid & metal), K (solid & metal), Rb (solid & metal) & Cs (liquid & metal). (½ Mark)

12. The ionic radius of a cation is always smaller than the parent atom because the **loss of one or more electrons increases the effective nuclear charge ( $Z_{\text{eff}}$ )**. As a result, the **force of attraction of nucleus for the remaining electrons increases and hence the electron cloud contracts** and ionic radii decreases. **(1 Mark)**  
 In contrast, the ionic radius of an anion is always larger than its parent atom because the **addition of one or more electrons decreases the effective nuclear charge ( $Z_{\text{eff}}$ )**. As a result, the **force of attraction of the nucleus for the remaining electrons decreases and hence electron cloud expands** and the ionic radii increases. **(1 Mark)**
13. (i) Each one of these ions **contains 18 electrons** and hence all are **isoelectronic ions**. **(1 Mark)**  
 (ii) The **ionic radii of isoelectronic ions decreases with the increase in the magnitude of the nuclear charge**. For example, consider the isoelectronic ions :  $\text{P}^{3-}$ ,  $\text{S}^{2-}$ ,  $\text{Cl}^-$ ,  $\text{K}^+$ ,  $\text{Ca}^{2+}$  and  $\text{Sc}^{3+}$ . All these ions have 18 electrons but their nuclear charges increase in the order :  $\text{P}^{3-}$ (+15),  $\text{S}^{2-}$  (+16),  $\text{Cl}^-$ (+17),  $\text{K}^+$  (+19),  $\text{Ca}^{2+}$  (+20) and  $\text{Sc}^{3+}$ (+21). Therefore, their **ionic radii decrease in the order :  $\text{P}^{3-} > \text{S}^{2-} > \text{Cl}^- > \text{K}^+ > \text{Ca}^{2+} > \text{Sc}^{3+}$** . **(1 Mark)**
14. **Atomic radii decrease across a period. Cations are smaller than their parent atoms. Among isoelectronic species, the one with larger positive nuclear charge will have a smaller radius (1 Mark)**. Hence the **largest species is K ; the smallest one is  $\text{Ca}^{2+}$** . **(1 Mark)**
15. Ionization energy of atomic hydrogen is the minimum amount of energy required to remove the electron from the ground state to infinity. Now, energy of the electron in the ground state =  $-5.19 \times 10^{-19}$  Cal. Energy of the electron at infinity = 0.  
 The **energy required to remove an electron in the ground state of hydrogen atom**  
 $= 0 - (\text{its energy in the ground state}) = -(-5.19 \times 10^{-19} \text{ Cal}) = 5.19 \times 10^{-19} \text{ Cal. (1 Mark)}$   
 $\therefore$  Ionization enthalpy per mole of hydrogen atoms =  $\frac{5.19 \times 10^{-19} \times 6.02 \times 10^{23}}{1000} \text{ KCal.}$   
**= 312.4 KCal/mol. (1 Mark)**
16. Within the main group elements, the ionization enthalpy decreases regularly as we move down the group due to the following two factors :  
 (i) **Atomic size**. On moving down the group, the **atomic size increases gradually due to the addition of one new principal energy shell at each succeeding element**. As a result, the distance of the valence electrons from the nucleus increases. Consequently, the **force of attraction of the nucleus for the valence electron decreases and hence the ionization enthalpy decreases. (1 Mark)**  
 (ii) **Screening effect**. **With the addition of new shells, the number of inner electron shells, which shield the valence electrons, increase**. In other words, the shielding effect or the screening effect increases. As a result, the **force of attraction of the nucleus for the valence electrons further decreased and hence, the ionisation enthalpy decreases. (1 Mark)**
17. It will be **more close to  $690 \text{ kJmol}^{-1}$  ( $\frac{1}{2}$  Mark)**. The value for B should be lower than that of Be ( $\frac{1}{2}$  Mark) because of **effective shielding of 2p-electrons from the nucleus by 2s-electrons**. This is expected also because of **fully filled s-orbital configuration of Be. (1 Mark)**
18. Ionization enthalpy depends upon the electronic configuration (number of electrons), and nuclear charge (number of protons). Since the isotopes of an element have the **same electronic configuration and same nuclear charge(1 Mark)**, they are expected to have **same ionization enthalpy. (1 Mark)**
19. The **second electron gain enthalpy of S is positive (1 Mark)** as explained below :  
 When an electron is added to S atom to form  $\text{S}^-$  ion, energy is released. Thus, first electron gain enthalpy of S is negative.  
 $\text{S(g)} + \text{e}^-(\text{g}) \longrightarrow \text{S}^-(\text{g}); \Delta_{\text{eg}}\text{H} = -200.4 \text{ kJmol}^{-1}$

But when another electron is added to  $S^-$  to form  $S^{2-}$  ion, energy is absorbed to overcome the strong electrostatic repulsion between the negatively charged  $S^-$  ion and the second electron being added. Thus, the second electron gain enthalpy of sulphur is positive. (1 Mark)



20. **Electron gain enthalpy** refers to the **enthalpy change when an isolated gaseous atom accepts an additional electron** to form a negative ion. It has unit of energy. (1 Mark)

**Electronegativity** refers to the **tendency of the atom** of an element to attract the shared pair of electrons towards it in a covalent bond. It has no unit. (1 Mark)

21. **Boron** is group 13 element with a **valence of 3**; **Oxygen** belongs to group 16 (chalcogen family) with a **valence of 2**. Hence, the **formula of the compound** formed would be  $B_2O_3$  (by criss-cross method). (1 Mark)

**Calcium** belongs to group 2 (alkaline earth metals) with a **valence of 2**; **Fluorine** belongs to group 17 (halogen family) with a **valence of 1**. Hence, the **formula of the compound** formed would be  $CaF_2$  (by criss-cross method). (1 Mark)

22. **BaO with water forms a strong base, whereas  $N_2O_5$  forms strong acid.**



Their basic or acidic nature can be qualitatively tested with litmus paper.

23. **Mendeleev used atomic weight as the basis of classification of elements in the periodic table.** (1 Mark) He arranged the then known elements in order of increasing atomic weights grouping together elements with similar properties. Mendeleev relied on the similarities in the properties of the compounds formed by the elements. **He realized that some of the elements did not fit in with his scheme of classification if the order of atomic weight was strictly followed. He ignored the order of atomic weights and placed the elements with similar properties together.** For example, iodine with lower atomic weight than tellurium (Group 16) was placed in group 17 along with fluorine, chlorine, etc., because of similarities in properties. Thus, Mendeleev did not stick strictly to arrangement of elements in the increasing order of atomic weight. (2 Marks)

24. In the modern periodic table, **each period starts with the filling of a new principal energy level.** Thus, the fourth period begins with the filling of principal quantum number,  $n = 4$ . **When  $n = 4$ ,  $\ell = 0, 1, 2, 3$ .** (1 Mark) But according to Aufbau principle, the electrons are added to different orbitals in order of their increasing energies. Now, the **energy of 4d-subshell is higher than that of 5s-subshell but energies of 3d-subshell is lower than that of 4p-subshell.** Therefore, in 4th period, electrons can be filled in only 4s, 3d and 4p-subshells, whose energies increase in the order:  $4s < 3d < 4p$ . (1 Mark) Now, s-subshell has one, p-subshell has three and d-subshell has five orbitals. Hence, in all, there are **9 (1 + 3 + 5) orbitals that can be filled in this period.** Since according to Pauli's exclusion principle, each orbital, at the maximum, can accommodate two electrons. Therefore, **9 orbitals, at the maximum, can have 18 electrons** and hence, fourth period has 18 elements. (1 Mark)

25. **Atomic radius** literally means **size of the atom.** It can be measured either by X-ray diffraction or by spectroscopic methods. **In case of non-metals, atomic radius is called covalent radius.** It is defined as **one-half the distance between the nuclei of two covalently bonded atoms of the same element in a molecule.** For example, the internuclear distance between two chlorine atoms in chlorine molecule is 198 pm. Therefore, the covalent radius of chlorine atom is  $198/2 = 99$  pm (0.99 Å). **In case of metals, atomic radius is called metallic radius.** It is defined as **one-half the distance between the two adjacent atoms in the crystal lattice.** For example, the distance between two adjacent copper atoms in solid copper is 256 pm. Therefore, the metallic radius of copper is  $256/2 = 128$  pm (1.28 Å). Another type of atomic radius is also defined called **Vander waal's radius, mainly used for noble gases.** It may be defined as **half of the internuclear distance between two adjacent atoms of the same element**

belonging to two nearest neighbouring molecules of the same substance in solid state. (2 Marks)  
**Ionic radius** means **size of the ions**. An ion can be a cation or an anion. **The size of a cation is always smaller than that of the parent atom** because loss of one or more electrons increases the effective nuclear charge. As a result, force of attraction on remaining electrons increases and hence the ionic size decreases. On the other hand, **the size of the anion is larger than the parent atom** because the addition of one or more electrons decreases the effective nuclear charge. As a result, the force of attraction decreases and hence the ionic size increases. For example, the ionic radius of  $\text{Na}^+$  is 95 pm while the atomic radius of sodium is 186 pm. Ionic radius of  $\text{F}^-$  is 136 pm while the atomic radius of fluorine atom is only 72 pm. (1 Mark)

26. Within a group, the **atomic radius increases down the group**. ( $\frac{1}{2}$  Mark) This is because a **new energy shell is added (i.e. principal quantum number increases by unity) at each succeeding element** while the number of electrons in the valence shell remains the same. In other words, the electrons in the valence shell of each succeeding element lie farther and farther away from the nucleus. As a result, the **force of attraction of the nucleus for the valence electrons decreases** and hence the **atomic size increases**. (1 Mark)

In contrast, the **atomic size decreases as we move from left to right in a period**. ( $\frac{1}{2}$  Mark) This is because within a period, the **outermost electrons remain in the same shell but the nuclear charge increases by one unit at each succeeding element**. Due to this increased charge, the **attraction of the nucleus for the outermost electrons increases** and hence the **atomic size decreases**. (1 Mark)

27. (i) The ionization enthalpy depends upon the type of electron to be removed from the same principal shell. In case of **Mg ( $1s^2 2s^2 2p^6 3s^2$ )**, the **outermost electron is present in 3s-orbital**, while in **Al ( $1s^2 2s^2 2p^6 3s^2 3p^1$ )**, it is present in **3p-orbital**. ( $\frac{1}{2}$  Mark) Since **3s-electrons are more strongly attracted by the nucleus than 3p-electrons**, therefore, **lesser amount of energy is required to knock out a 3p-electron than 3s-electron**. Consequently,  $\Delta_i H$  of Mg is greater than that of Al. This is expected also because of **fully filled s-orbital configuration of Mg**. (1 Mark)

(ii) The electronic configuration of **P ( $1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1 3p_z^1$ )**, in which **3p-orbitals are exactly half-filled, is more stable than the electronic configuration of S ( $1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^1 3p_z^1$ )** in which the 3p-orbitals are neither exactly half-filled nor completely filled. Therefore, it is **difficult to remove an electron from P than from S**. As a result,  $\Delta_i H$  of P is higher than that of S. (1 Mark)

Further, the electronic configuration of Cl is  $1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^2 3p_z^1$ . **Because of higher nuclear charge (+17), the first ionization enthalpy of Cl is higher than that of S**. ( $\frac{1}{2}$  Mark)

28. The electronic configurations of Li and Be are :

Li :  $1s^2 2s^1$                       Be :  $1s^2 2s^2$

Thus, the first electron in both the cases has to be removed from the 2s-orbital. But the **nuclear charge of Li (+3) is lower than that of Be (+4)**. Therefore, the first ionization energy of Lithium is lower than that of Beryllium. (1 Mark)

After the loss of first electron, the **electronic configuration of  $\text{Li}^+$  is  $1s^2$** . Here, the **electron is to be removed from inert (helium) gas configuration**, which is very stable and hence **removal of second electron from Lithium is very difficult**. However, in case of Beryllium, after the loss of first electron, the **electronic configuration of  $\text{Be}^+$  is  $1s^2 2s^1$** . Here, the **electron is to be removed from a 2s orbital which is much easier than to remove an electron from inert gas configuration**. Therefore, the second ionization enthalpy of Lithium is higher than that of Beryllium. (2 Marks)

29. On moving down the group 13 from B to Al, the ionization enthalpy decreases as expected due to an increase in atomic size and screening effect, which outweighs the effect of increased nuclear charge. However,  $\Delta_i H$  of Ga is only slightly greater ( $2 \text{ kJ mol}^{-1}$ ) than that of Al, while that of Tl is much higher than In. These deviations can be explained as follows :

Al follows immediately after s-block elements, while Ga and In follow after d-block elements and Tl after d- and f-block elements. These extra d- and f-electrons do not shield (or screen) the outermost shell electrons from the nucleus very effectively. (1 Mark) As a result, the valence electrons remain more tightly held by the nucleus and hence a larger amount of energy is needed for their removal. (1 Mark) This explains why Ga has higher ionization enthalpy than Al. Further, on moving down the group from Ga to In, the increased shielding effect (due to the presence of additional 4d-electrons) outweighs the effect of increased nuclear charge ( $49 - 31 = 18$  units) and hence the  $\Delta_i H$  of In is lower than that of Ga. Thereafter, the effect of increased nuclear charge ( $81 - 49 = 32$  units) outweighs the shielding effect due to the presence of additional 4f and 5d electrons and hence the  $\Delta_i H$  of Tl is higher than that of In. (1 Mark)

30. (i) Ionization enthalpy is the minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom in ground state ( $\frac{1}{2}$  Mark), so as to convert it into a gaseous cation. The force with which an electron is attracted by the nucleus of an atom is appreciably affected by presence of other atoms within its molecule or in the neighbourhood. Therefore, for the purpose of determination of ionization enthalpy, it is essential that these interatomic forces of attraction should be minimum. Since in the gaseous state, the atoms are widely separated, ( $\frac{1}{2}$  Mark), therefore, these interatomic forces are minimum. Further, since it is not possible to isolate a single atom for the purpose of determination of its ionization enthalpy, therefore, the interatomic distances are further increased by carrying out the measurement of the gaseous atom at a low pressure. It is because of these reasons, that the term isolated gaseous atom has been included in the definition of ionization enthalpy. (1 Mark)

(ii) Electron gain enthalpy is the enthalpy change when an isolated gaseous atom in the ground state accepts an extra electron ( $\frac{1}{2}$  Mark) to form gaseous negative ion.

The term 'isolated gaseous atom' has already been explained above. The term 'ground state' here means that the atom must be present in the most stable electronic configuration i.e., the ground state. The reason being that when the isolated gaseous atom is in the excited state, lesser amount of energy will be released when it gets converted into gaseous anion after accepting an electron and lesser amount of energy will be required when it gets converted into gaseous cation. Therefore, for comparison purpose, the ionisation enthalpies and electron gain enthalpies of gaseous atoms must be determined in their respective most stable state, i.e., ground state. (1 Mark)

31. Electron gain enthalpy generally becomes more negative across a period as we move from left to right. Within a group, electron gain enthalpy becomes less negative down a group. (1 Mark) However, adding an electron to the 2p-orbital leads to greater repulsion than adding an electron to the larger 3p-orbital. Hence the element with most negative electron gain enthalpy is chlorine; the one with the least negative electron gain enthalpy is oxygen. (1 Mark)

32. The electronegativity of any given atom is not constant. Therefore, the statement that the electronegativity of C on Pauling scale is 2.5 in all carbon compounds is wrong. (1 Mark) Actually, electronegativity varies with the state of hybridization and the oxidation state of the element. The electronegativity increases with increase in % s-character in hybridisation and also with increase in oxidation state of element. (1 Mark) For example, the electronegativity of C increases as we move from  $sp^3 \rightarrow sp^2 \rightarrow sp$  hybridised C atom. Similarly, the electronegativity of C in +4 oxidation state is higher than where the oxidation state of C is -4. (1 Mark)

33. (i)  $n = 2$  indicates that the element belongs to second period. ( $\frac{1}{2}$  Mark) Since the last electron enters the p-orbital, therefore, the given element is a p-block element. ( $\frac{1}{2}$  Mark) For p-block elements, group number = 10 + number of electrons in the valence shell.

$\therefore$  Group number of the element =  $10 + 4 = 14$ . ( $\frac{1}{2}$  Mark)

(ii)  $n = 4$  indicates that the element lies in the 4th period. ( $\frac{1}{2}$  Mark)

Since the d-orbital is incomplete, therefore, it is d-block element. ( $\frac{1}{2}$  Mark)

The **group number of the d-block element** = number of  $(n - 1)d$ -electrons + number of  $(n)s$ -electrons =  $5 + 1 = 6$ . (1 Mark)

(iii)  $n = 6$  means that the **element lies in the sixth period**. (½ Mark) Since the last electron goes to the  $f$ -orbital, therefore, the **element is a f-block element**. (½ Mark) **All f-block elements lie in 3rd group**. (½ Mark).

34. (a) **Sodium** is an element of 1st group. It has only one electron in the valence shell, therefore, **its valence is 1**. **Sulphur** is a 16th group element with a **valence of  $8 - 6 = 2$** . Therefore, **formula of the compound formed would be  $\text{Na}_2\text{S}$** . (½ Mark)
- (b) **Calcium** is an alkaline earth metal (Group-2) and hence has a **valence of 2**. **Phosphorus** is a 15th group element with a **valence of  $8 - 5 = 3$** . Thus, the **formula of the compound formed would be  $\text{Ca}_3\text{P}_2$** . (1 Mark)
- (c) **Boron** is a 13th group element with a **valence of 3**, while **chlorine** is a halogen (17th group) with a **valence of 1**. Therefore, the **formula of the compound formed would be  $\text{BCl}_3$** . (½ Mark)
- (d) **Carbon** is a 14th group element with a **valence of 4**, while **sulphur** is a 16th group element with a **valence of 2**. Hence, the **formula of the compound formed is  $\text{CS}_2$** . (1 Mark)
- (e) **Nitrogen** is a 15th group element with a **valence of 3**, while **bromine** is a 17th group element with a **valence of 1**. Hence, the **formula of the compound formed would be  $\text{NBr}_3$** . (½ Mark)
- (f) **Element with atomic number 71 is a lanthanoid called Lutetium (Lu)**. **Its common valence is 3**. **Fluorine** is a 17th group element with a **valence of 1**. Therefore, the **formula of the compound would be  $\text{LuF}_3$** . (1½ Marks)

## EXERCISE - 1

A-1.	(3)	A-2.	(3)	A-3.	(4)	A-4.	(3)	A-5.	(3)
A-6.	(1)	A-7.	(4)	A-8.	(1)	A-9.	(1)	A-10.	(4)
A-11.	(2)	A-12.	(2)	A-13.	(3)	A-14.	(4)	A-15.	(2)
A-16.	(2)	A-17.	(4)	A-18.	(2)	A-19.	(4)	A-20.	(4)
A-21.	(4)	A-22.	(4)	B-1.	(3)	B-2.	(2)	B-3.	(4)
B-4.	(4)	B-5.	(2)	B-6.	(4)	B-7.	(2)	B-8.	(2)
B-9.	(1)	B-10.	(3)	B-11.	(3)	B-12.	(3)	B-13.	(1)
B-14.	(1)	B-15.	(2)	B-16.	(1)	B-17.	(2)	B-18.	(3)
C-1.	(1)	C-2.	(1)	C-3.	(1)	C-4.	(2)	C-5.	(2)
C-6.	(1)	C-7.	(2)	C-8.	(1)	C-9.	(2)	C-10.	(4)
C-11.	(3)	C-12.	(2)	C-13.	(3)	C-14.	(3)	C-15.	(2)
C-16.	(3)	C-17.	(2)	C-18.	(2)	C-19.	(3)	C-20.	(4)
C-21.	(2)	C-22.	(2)	D-1.	(3)	D-2.	(4)	D-3.	(2)
D-4.	(3)	D-5.	(3)	D-6.	(2)	D-7.	(1)	D-8.	(2)
D-9.	(3)	D-10.	(2)	D-11.	(3)	D-12.	(3)	D-13.	(4)
D-14.	(4)	E-1.	(3)	E-2.	(1)	E-3.	(3)	E-4.	(1)
E-5.	(3)	E-6.	(1)	E-7.	(4)	E-8.	(1)	F-1.	(4)
F-2.	(3)	F-3.	(2)	F-4.	(1)	F-5.	(4)	F-6.	(3)

---

**EXERCISE - 2**

---

1.	(3)	2.	(3)	3.	(3)	4.	(4)	5.	(1)
6.	(3)	7.	(2)	8.	(1)	9.	(4)	10.	(3)
11.	(2)	12.	(2)	13.	(4)	14.	(1)	15.	(3)
16.	(3)	17.	(4)	18.	(3)	19.	(1)	20.	(4)
21.	(1)	22.	(4)						

---

**EXERCISE - 3**

---

1.	(1)	2.	(1)	3.	(2)	4.	(4)	5.	(1)
6.	(4)	7.	(3)	8.	(1)	9.	(3)	10.	(1)
11.	(1)	12.	(3)	13.	(2)	14.	(1)	15.	(1)
16.	(2)	17.	(2)						

---

**EXERCISE - 4**

---

**Level - 1**

1.	(4)	2.	(3)	3.	(3)	4.	(3)	5.	(4)	6.	(4)	7.	(1)
8.	(2)	9.	(2)	10.	(1)	11.	(3)	12.	(1)	13.	(1)	14.	(2)
15.	(2)	16.	(1)	17.	(2)	18.	(4)	19.	(4)	20.	(4)	21.	(2)
22.	(3)	23.	(1)	24.	(3)	25.	(1)	26.	(2)	27.	(2)	28.	(2)
29.	(4)	30.	(3)	31.	(3)	32.	(2)	33.	(3)	34.	(1)	35.	(1)
36.	(2)	37.	(4)	38.	(4)	39.	(4)	40.	(2)	41.	(3)	42.	(2)
43.	(3)	44.	(2)	45.	(1)	46.	(3)	47.	(1)	48.	(4)	49.	(3)
50.	(2)	51.	(2)	52.	(2)	53.	(2)	54.	(2)	55.	(1)	56.	(1)
57.	(1)	58.	(1)	59.	(1)	60.	(2)	61.	(4)				

**Level - 2**

1.	(1)	2.	(3)	3.	(1)	4.	(2)	5.	(3)
6.	(3)	7.	(3)	8.	(4)	9.	(4)	10.	(2)
11.	(4)	12.	(3)	13.	(4)	14.	(4)	15.	(4)
16.	(1)	17.	(1)	18.	(2)	19.	(4)	20.	(4)
21.	(1)	22.	(2)	23.	(3)	24.	(3)	25.	(2)

**Level - 3**

1.	(A)	2.	(B)	3.	(B)	4.	$Al^{3+} < Mg^{2+} < Li^+ < K^+$		
5.	Inert Pair Effect			6.	(C)	7.*	(D)		
8.	CrO <sub>3</sub> on reaction with water produces Chromic acid (H <sub>2</sub> CrO <sub>4</sub> ). So, CrO <sub>3</sub> is an acid anhydride. $CrO_3 + H_2O \longrightarrow H_2CrO_4$								
9.	(B)	10.	(A)	11.	(C)	12.	(B)	13.	(A)
14.	(A)	15.	(B)	16.	$Cl_2O_7 < SO_3 < CO_2 < B_2O_3 < BaO$		17.	(C)	