NURTURE COURSE

QUANTUM NUMBER & ELECTRONIC CONFIGURATION
PERIODIC TABLE & PERIODIC PROPERTIES
CHEMICAL BONDING

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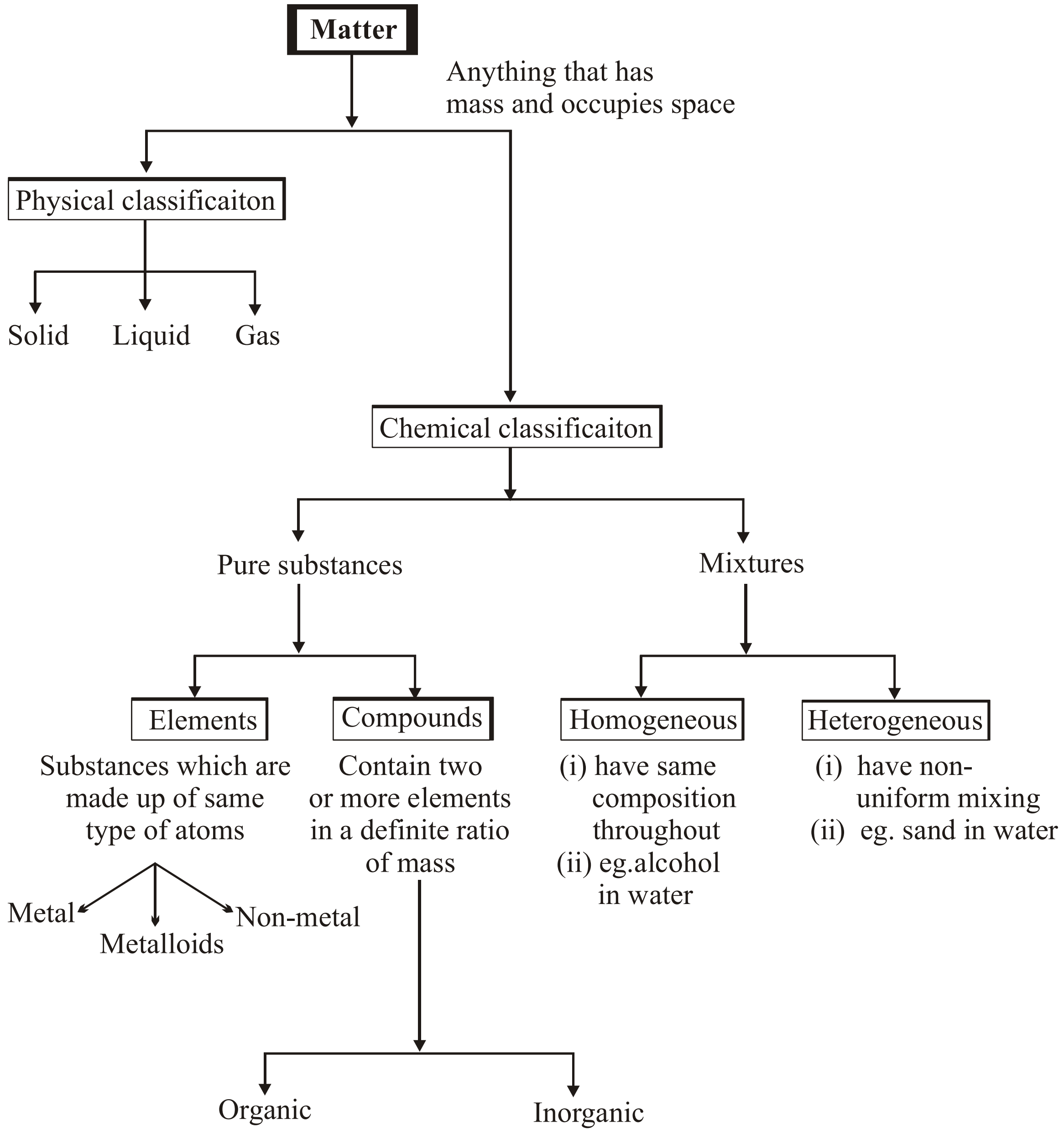
MATTER & ITS CLASSIFICATION



Example.1 Which of the following is homogeneous mixture :

(A) Oil + Water (B) Milk

(C) Salt dissolved in water (D) All of these

Example.2 Which of the following molecule is tetra-atomic :

(A) CH_2Cl_2 (B) NH_3 (C) H_2O (D) Both (B) and (C)

ATOM : An atom is the smallest particle of an element (made up of still smaller particle like electrons, protons, neutrons etc.) which can take part in a chemical reaction. It may or may not exist free in nature.

Name of particle	Mass	Nature of charge	Amount of charge	Presence in the atom
(i) Electron symbol = (e) Notation = $_{-1}e^{\circ}$ Discoverer J.J. Thomson (1897)	9.11×10^{-28} g $\frac{1}{1837}$th of H-atom	Negatively charged	-1.602×10^{-19} Coulomb or -4.8×10^{-10} e.s.u	Outside the nucleus
(ii) Proton symbol = (p) Notation = ($_1H^1$) Discoverer Rutherford (1911)	1.6725×10^{-24}g	Positively charged	$+ 1.602 \times 10^{-19}$ coulomb $+ 4.8 \times 10^{-10}$ e.s.u.	Inside the nucleus of an atom
(ii) Neutron symbol = (n) Notation = ($_0n^1$) Discoverer J. Chadwick (1932)	1.675×10^{-24} g	Neutral	0	Inside the nucleus of an atom

Representation of atom : $_ZX^A$

A → Mass number : (total number of protons + total number of neutrons present in an atom.)

Z → Atomic number : (total number of protons present in an atom.)

⇒ ***Isotope :*** Atoms of given element which have same atomic number but different mass number are called isotope : e.g. $_1H^1$, $_1H^2$, $_1H^3$ etc.

⇒ ***Isobar :*** Atoms of different elements with the same mass number but different atomic number .

e.g. $_{18}Ar^{40}$, $_{19}K^{40}$ and $_{20}Ca^{40}$

⇒ ***Iso-electronic species :*** Species (atom, molecules or ions) having same number of electrons are called iso-electronic e.g H^{-}, He, Li^{+} and Be^{2+} have 2 valence electrons each.

Note : Now a days this concept is extended to consider the same valence shell electron also.

⇒ ***Iso-sters :*** Species having same number of electrons & same number of atoms. eg. N_2O, CO_2

⇒ ***Iso-diaphers :*** Species having same difference in number of neutrons and protons. eg. $^{12}_{6}C$, $^{16}_{8}O$ and $^{14}_{7}N$.

⇒ ***Orbital :*** An orbital is defined as that zone in space where electron is most likely to be found .The orbitals are characterized by a set of 3 quantum numbers (n,l,m).

QUANTUM NUMBERS : Quantum numbers give complete information about an electron or orbital in an atom.

1. Principal Quantum number (n) :

(i) Permissible value of **n → 1 to ∞**

(ii) It represents shell number/energy level

(iii) The energy states corresponding to different principal quantum numbers are denoted by letters K,L,M, N etc.

n	:	1	2	3	4	5	6
Designation of shell	:	K	L	M	N	O	P

(iv) It indicates the distance of an electron from the nucleus.

(v) It also determines the energy of the electron. In general higher the value of 'n', higher is the energy of a electron.

(vi) It give an idea of total number of orbitals & electron (which may) present in a shell & that equal to $\mathbf{n^2}$ & $\mathbf{2n^2}$ respectively.

2. Azimuthal Quantum number (l) :

(i) The values of l depends upon the value of 'n' and possible values are '0' to (n –1) .

(ii) It gives the name of subshells associated with the energy level and number of subshells within an energy level.

(iii) The different value of 'l' indicates the shape of orbitals and designated as follows :

Value	**Notation**	**Name**	**Shape**
$l = 0$	s	Sharp	Spherical
$l = 1$	p	Principal	Dumbell
$l = 2$	d	Diffused	Double Dumbell
$l = 3$	f	Fundamental	Complex

(iv) It also determines the energy of orbital along with n.

For a particular energy level/shell energy of subshell is in the following order → **s < p < d < f**

(v) It gives the total number of orbitals in a subshell & that equals to **$(2l + 1)$** and number of electron in a subshell = **$2\,(2l + 1)$**

3. Magnetic Quantum number (m or m_l) :

(i) The value of m depends upon the value of l and it may have integral value $-l$ to $+l$ including zero.

(ii) It gives the number of orbitals in a given subshell and orientation of different orbitals in space.

e.g. for n = 4, l = 0 to 3.

l	0	1	2	3
m	0	+1, 0, – 1	+2, 1, 0, –1, –2	+3, +2, +1, 0, –1, –2, –3
Possible Orientation	1	3	5	7
Orbitals	s	p_x , p_y, p_z	$d_{z^2}, d_{x^2-y^2}$ d_{xy}, d_{yz}, d_{xz}	Not in syllabus

(iii) The orbitals having same value of n and l but different value of m, have same energy in absence of external electric & magnetic field. These orbitals having same energy of a particular subshell is known as Degenerate orbitals.

4. Spin Quantum number (s) & magnetic spin quantum number (m_s) :

(i) While moving around the nucleus, the electron always spins about its own axis either clockwise or anticlockwise. The magnetic spin quantum number represents the direction of electron spin (rotation) around its own axis (clockwise or anticlockwise).

(ii) There are two possible values of $\mathbf{m_s}$ are $+\frac{1}{2}$ & $-\frac{1}{2}$ and represented by the two arrows ↑ (spin up) and ↓ (spin down).

RULES FOR FILLING ELECTRONS :

1. Pauli's exclusion principle

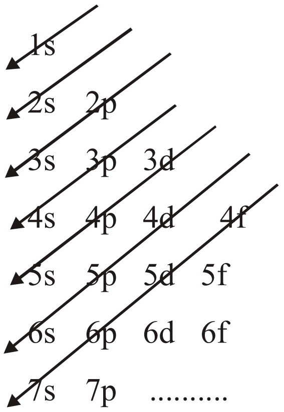
'No two electrons in an atom can have same values of all the four quantum numbers.

An orbital accommodates two electron with opposite spin. These two electrons have same values of principal, azimuthal and magnetic quantum number but the fourth, i.e. magnetic spin quantum number will be different. i.e.

For	K, shell (n = 1)
	$l = 0$, m = 0
For 1st Electron	$n = 1, l = 0, m = 0, m_s = +\frac{1}{2}$
For 2nd Electron	$n = 1, l = 0, m = 0, m_s = -\frac{1}{2}$

2. *Aufbau Principle (Means Building up) :*

The electrons are added progressively to the various orbitals in the order of increasing energies starting with the orbital of the lowest energy



$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p$

Alternatively, the order of increase of energy of orbitals can be calculated from $(n + l)$ rule.

(i) Lower the value of $(n + l)$ for an orbital, the lower will be its energy.

(ii) If two orbitals have the same $(n + l)$ value, then orbital with lower value of n has the lower energy.

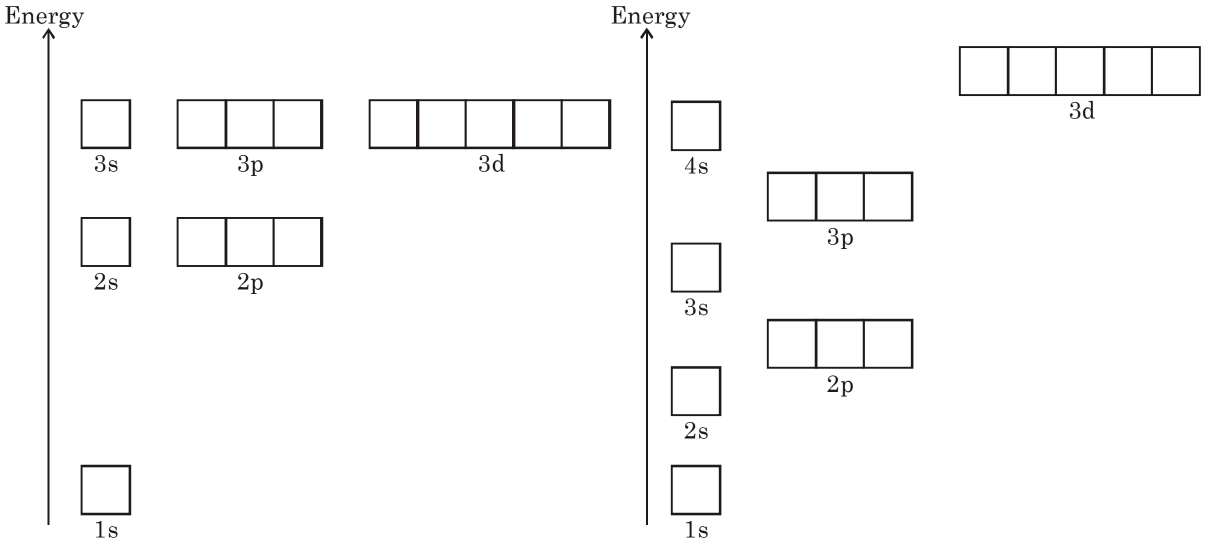
e.g. 2p & 3s

For 2p, $n = 2$, $l = 1$, $(n + l) = 2 + 1 = 3$

For 3s, $n = 3$, $l = 0$, $(n + l) = 3 + 0 = 3$

Then for 2p, n is lesser than for 3s, so 2p has lower energy than 3s.

(iii) $1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f$ .... energy order of different orbitals for single electron system like H, He^+, Li^{+2} etc.



(A) For single electron or hydrogenic atom (B) Multi electronic atoms

Energy level diagram for few electronic shells :

Example.4 Write the increasing order of energies of 4s, 3p, 4p and 3d.

Ans.
For 4s, $n = 4$, $l = 0$, $(n + l) = 4$
For 3p, $n = 3$, $l = 1$, $(n + l) = 4$
For 4p, $n = 4$, $l = 1$, $(n + l) = 5$
For 3d, $n = 3$, $l = 2$, $(n + l) = 5$
$\Rightarrow$ $3p < 4s < 3d < 4p$ increasing order

3. *Hund's rule of maximum multiplicity :*

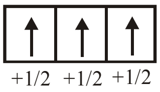
This rule deals with the filling of electrons into the orbitals belonging to the same subshell i.e. orbitals of equal energy, called degenerate orbitals.

"Electrons are distributed among the orbitals of a subshell in such a way as to give the maximum number of unpaired electron with parallel spins. "

"Pairing of electrons in the orbitals belonging to the same subshell (p, d, f) does not take place until each orbital belonging to that subshell has got one electron each i.e. singly occupied. Moreover, the singly occupied orbitals must have the electrons with the parallel spin multiplicity"

Multiplicity = $2|S| + 1$, where S = Total spin.

i.e.

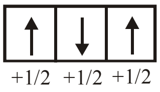


Find total spin & multiplicity

$$\text{Total spin } S = \frac{1}{2} + \frac{1}{2} + \frac{1}{2} = \frac{3}{2}$$

$$\text{Multiplicity} = 2 \times \frac{3}{2} + 1 = 4$$

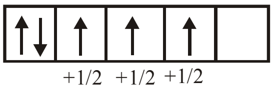
i.e.



$$\text{Total spin } S = \frac{1}{2} - \frac{1}{2} + \frac{1}{2} = \frac{1}{2}$$

$$\text{Multiplicity} = 2 \times \frac{1}{2} + 1 = 2$$

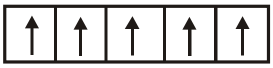
i.e.



$$\text{Total spin } S = \frac{1}{2} + \frac{1}{2} + \frac{1}{2} = \frac{3}{2}$$

$$\text{Multiplicity} = 2 \times \frac{3}{2} + 1 = 4$$

i.e.

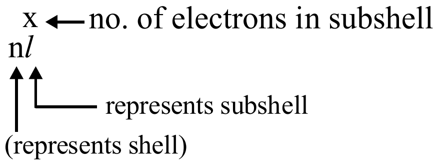


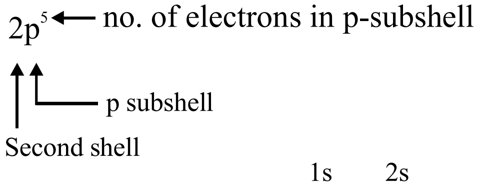
$$\text{Total spin} = 5 \times \frac{1}{2} = \frac{5}{2}$$

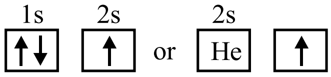
$$\text{Multiplicity} = 2 \times \frac{5}{2} + 1 = 6$$

ELECTRONIC CONFIGURATION OF ATOMS :

The distribution of electrons in various shells, subshells and orbitals, in an atom of an element, is called its electronic configuration.





$_3$Li ; $1s^2, 2s^1$ or [He] $2s^1$ or 

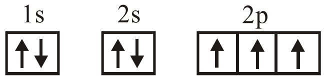
Electronic configuration :

Example.5 Nitrogen

Ans. $_7$N : $1s^2, 2s^2\ 2p^3$ [Orbital notation method]

$1s^2, 2s^2\ 2p_x^{\ 1}\, 2p_y^{\ 1} 2p_z^{\ 1}$

or



[Orbital diagram method]

or

[He] $2s^2\ 2p^3$ [Condensed form]

Extra stability of Half-filled and fully-filled orbitals.

The electronic configuration of most of the atoms follows the Aufbau's rule. However, in certain elements such as Cr, Cu etc. Where the two subshells (4s and3d) differ slightly in their energies (4s < 3d), an electron shifts from a subshell of lower energy (4s) to a subshell of higher energy (3d), provided such a shift results in all orbitals of the subshell of higher energy getting either completely filled or half-filled.

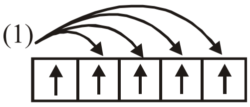
$_{24}\text{Cr} \rightarrow$ [Ar] $3d^5, 4s^1$ and not [Ar] $3d^4\ 4s^2$

$_{29}\text{Cu} \rightarrow$ [Ar] $3d^{10}, 4s^1$ and not [Ar] $3d^9\ 4s^2$

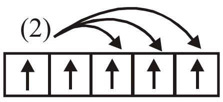
It has been found that there is extra stability associated with these electronic configuration. This stabilization is due to the following two factors.

(i) ***Symmetrical distribution of electron :*** It is well known that symmetry leads to stability. The completely filled or half-filled subshell have symmetrical distribution of electron in them and are therefore more stable. This effect is more dominant in d and f-orbitals. This means three or six electrons in p-subshell, 5 or 10 electrons in d-subshell and 7 or 14 in f-subshell forms a stable arrangement.

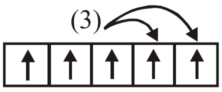
(ii) ***Exchange energy :*** This stabilizing effect arises whenever two or more electrons with the same spin are present in the degenerate orbitals of a subshell. these electrons tend to exchange their positions and the energy released due to this exchange is called exchange energy. The number of exchanges that can take place is maximum when the subshell is either half filled or fully filled. As result the exchange energy is maximum and so is the stability.



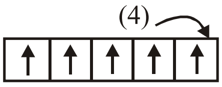
4-exchanges
by electron 1



3-exchanges
by electron 2



2-exchanges
by electron 3



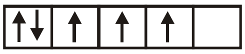
1-exchanges
by electron 4

Total exchange pairs = 10

$$\frac{n(n-1)}{2} \rightarrow \text{Number of exchange pairs}$$

n → Number of electron with parallel spins.

e.g.



Only 6 total exchange possible

Exceptional electronic configuration

S.No.	Element	Z	Configuration
1	Cr	24	$[Ar]4s^1 3d^5$
2.	Cu	29	$[Ar]4s^1 3d^{10}$
3.	Nb	41	$[Kr]5s^1 4d^4$
4.	Mo	42	$[Kr]5s^1 4d^5$
5.	Ru	44	$[Kr]5s^1 4d^7$
6.	Rh	45	$[Kr]5s^1 4d^7$
7.	Pd	46	$[Kr]4d^{10}$
8.	Ag	47	$[Kr]5s^1 4d^{10}$
9.	La	57	$[Xe]6s^2 5d^1$
10.	Ce	58	$[Xe]6s^2 4f^1 5d^1$
11.	Gd	64	$[Xe]6s^2 4f^7 5d^1$
12.	Pt	78	$[Xe]6s^1 4f^{14} 5d^9$
13.	Au	79	$[Xe]6s^1 4f^{14} 5d^{10}$
14.	Ac	89	$[Rn]7s^1 6d^1$
15.	Th	90	$[Rn]7s^2 6d^2$

MAGNETIC PROPERTIES :

❖ ***Paramagnetism :***

(i) The substances which are weakly attracted by magnetic field are paramagnetic and this phenomenon is known as paramagnetism.

(ii) Their magnetic character is retained till they are in magnetic field and lose their magnetism when removed from magnetic field.

❖ ***Diamagnetism :***

(i) The substances which are weakly repelled by magnetic field are diamagnetic and this phenomenon is known as diamagnetism.

(ii) Diamagnetic substances lack unpaired electrons and their spin magnetic moment is zero e.g., NaCl, N_2O_4 etc.

❖ ***Spin magnetic moment :***

The spin magnetic moment of electron (excluding orbit magnetic moment) is given by :

$\mu = \sqrt{[n(n+2)]}$ B.M.

Where n is number of unpaired electron in species.

The magnetic moment is expressed in Bohr magneton (B.M.)

Example.6 A compound of vanadium has magnetic moment of 1.73 BM. Work out the electronic configuration of vanadium ion in the compound.

Ans. Vanadium belongs to 3d series with Z = 23. The magnetic moment of 3d series metal is given by spin only formula.

$\mu = \sqrt{n(n+2)}$ BM (BM = Bohr's magneton)

$\because \quad \sqrt{1.73} = \sqrt{3}$

$\Rightarrow$ $n(n+2) = 3 \Rightarrow n = 1$

$\Rightarrow$ Magnetic moment correspond to one unpaired electron.

$\Rightarrow$ Electronic configuration of vanadium atom $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$.

For one unpaired electron 4 electron must be removed in which first 2 electron are lost from 4s orbital (outermost).

Electronic configuration of V^{+4}

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^1$

Nodal Planes of different orbitals :

Nodal plane is a plane at which the probability of finding an electron becomes zero.

eg.

Orbital	Nodal plane	Orbital	Nodal plane
s	None	d_{xy}	XZ & YZ planes
p_x	YZ plane	d_{yz}	XZ & XY planes
p_y	XZ plane	d_{xz}	XY & YZ planes
p_z	XY plane	$d_{x^2-y^2}$	Planes perpendicular to XY plane, passing through origin (nucleus) and inclined at 45° to X & Y axis.
		d_{z^2}	None (two nodal cones are available)

SCREENING EFFECT (s) AND EFFECTIVE NUCLEAR CHARGE (Z_{eff})

(a) Valence shell electron suffer force of attraction due to nucleus and force of repulsion due to inner shell electrons.

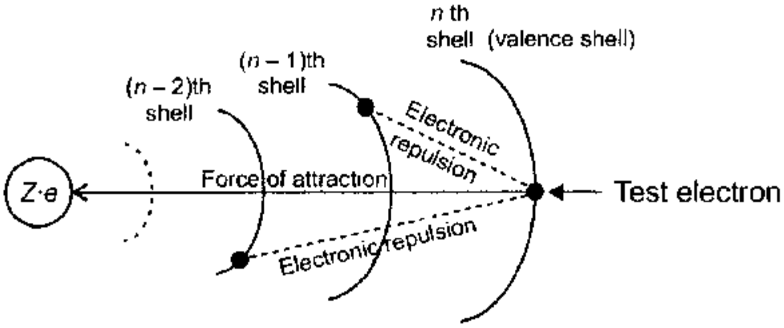
(b) The decrease in force of attraction on valence electron due to inner shell electron is called screening effect or shielding effect.(i.e. total repulsive force is called shielding effect.)

(c) Due to screening effect, valence shell electron experiences less force of attraction by nucleus.

(d) Due to screening effect, net attractive force felt by the electron is measured by effective nuclear charge, Z_{eff}

(e) If nuclear charge = Z, then effective nuclear charge = $Z - \sigma$ (Where σ 'sigma' is called screening constant/sheilding constant)

So, $Z_{eff} = Z - \sigma$



CALCULATION OF σ (using slater's rule)

To calculate the shielding constant (σ) :

(a) Write the electronic configuration of the element in the following order and groupings :

(1s), (2s, 2p), (3s, 3p), (3d), (4s, 4p), (4d), (4f), (5s, 5p), etc.

For s and p electrons :

(b) Electrons in any group to the right of the (ns, np) group contribute nothing to the shielding constant.

(n-shell no. of the electron for which σ is calculated)

(c) All of the other electrons in the (ns, np) group, shield the concern electron to an extent of 0.35 each. (Except for the 1s orbital for which value is 0.30).

(d) All electrons in the (n – 1) shell shield to an extent of 0.85 each.

(e) All electrons (n – 2) or lower group shield completely ; that is, their contribution is 1.00 each.

For d and f electrons :

(f) Electrons in any group to the right of the nd or nf group contribute nothing to the shielding constant.

(g) All the other electrons in the nd or nf group, shield the valence electron to an extent of 0.35 each.

(h) All electrons in groups lying to the left of the nd or nf group contribute 1.00.

(Effective Nuclear charge of elements of second period)

Element	Electronic Configaration	Z	σ of ns & np electron (a)	σ (n–1) orbital (b)	Total Screeing Constant (a + b)	Effective nuclear charge# Z*=Z – σ
$_3$Li	$1s^2\,2s^1$	3	–	0.85× 2=1.70	1.70	1.30
$_4$Be	$1s^2, 2s^2$	4	1× 0.35=0.35	0.85× 2=1.70	2.05	1.95
$_5$B	$1s^2,2s^2,2p^1$	5	2× 0.35=0.70	0.85× 2=1.70	2.40	2.60
$_6$C	$1s^2,2s^2,2p^2$	6	3× 0.35=1.05	0.85× 2=1.70	2.75	3.25
$_7$N	$1s^2,2s^2,2p^3$	7	4× 0.35=1.40	0.85× 2=1.70	3.10	3.90
$_8$O	$1s^2,2s^2,2p^4$	8	5× 0.35=1.75	0.85× 2=1.70	3.45	4.55
$_9$F	$1s^2,2s^2,2p^5$	9	6× 0.35=2.10	0.85× 2=1.70	3.80	5.20

\# Calculated for valence electron.

Key Points :

(a) From left to right in a period Z_{eff} increases

(i) For s and p-block elements, Z_{eff} in a period increases by 0.65 where atomic number increases by 1, and hence atomic size decreases considerably.

(ii) In transition series Z increase by + 1 but screening constant increases by 0.85 So Z_{eff} is increased by 0.15

(1– 0.85 = 0.15) [Because e^- enters in (n – 1) orbit which has value of $\sigma = 0.85$]

(b) From top to bottom in a group Z_{eff} remain constant for s-block elements, after Li and Be.

Element	Li	Na	K	Rb	Cs	Fr
Z_{eff}	1.30	2.20	2.20	2.20	2.20	2.20

Example-7 :

What is the effective nuclear charge at the periphery of nitrogen atom when an extra electron is added during the formation of an anion. Also find the value of Z_{eff} when the atom is ionized to N^+.

Ans. Ground state electron configuration of N(Z = 7) = $1s^2\,2s^2\,2p^3$

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

Shielding constant for the last 2p electron,

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

So $Z_{eff} = Z - \sigma = 7 - 3.45 = 3.55$

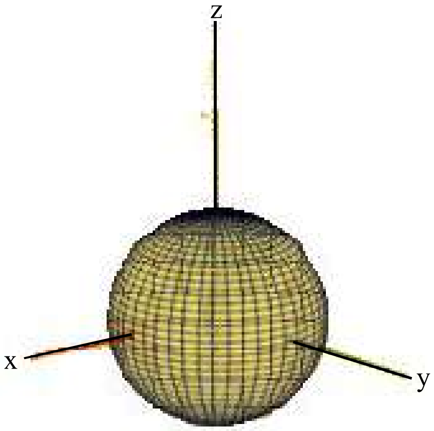
Electron configuration of $N^+ = (1s^2)\,(2s^2\,2p^2)$

Shielding constant for the last 2p electron,

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

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

SHAPES OF ATOMIC ORBITALS



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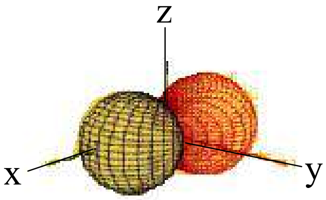
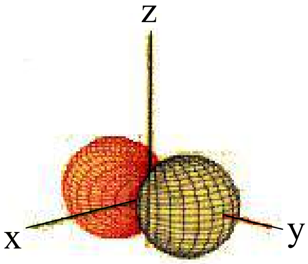
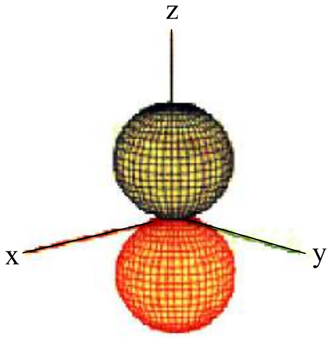
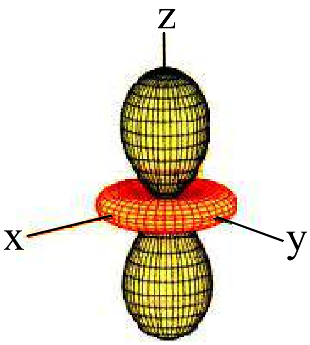
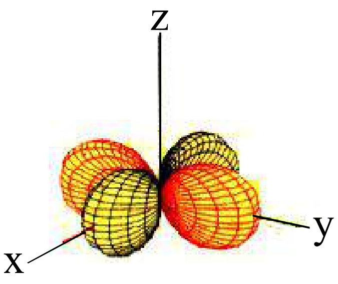


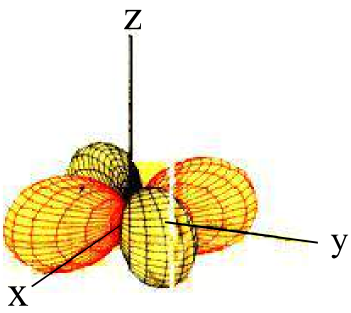




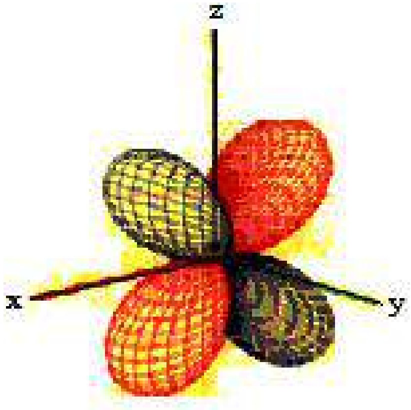
p_X p_y p_z



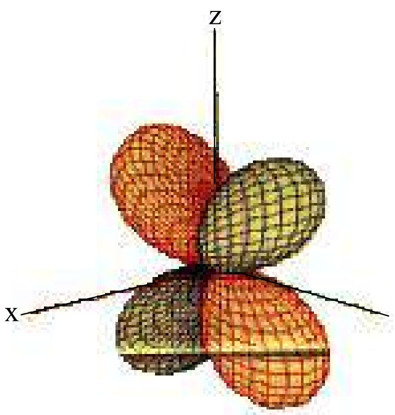




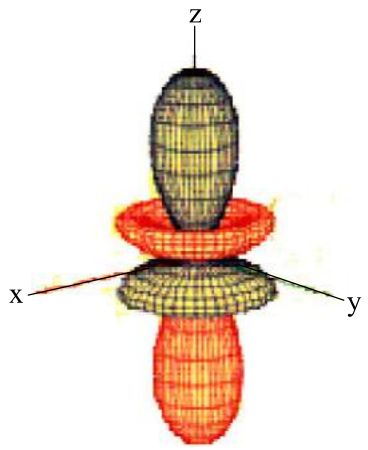
d_{z^2} $d_{x^2-y^2}$ d_{xy}



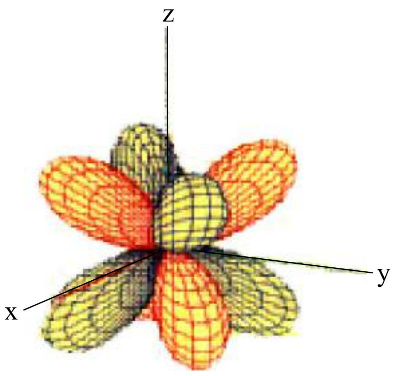
d_{xz}



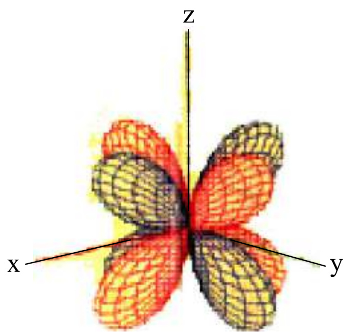
d_{yz}



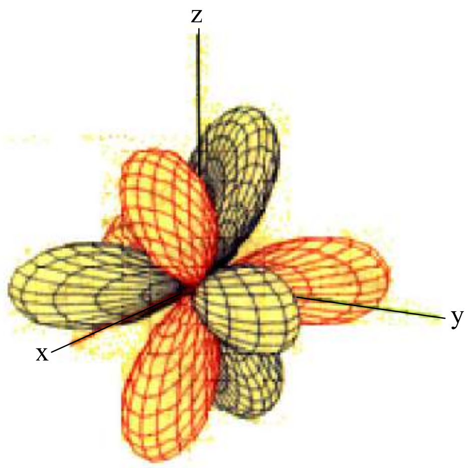
f_{z^3}



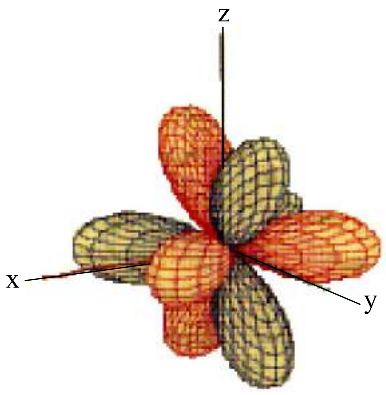
f_{xyz}



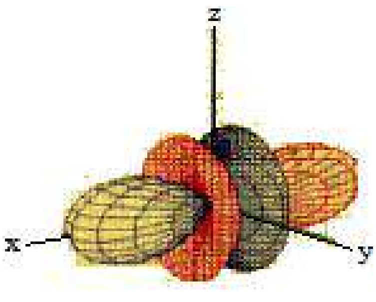
$f_{z(x^2-y^2)}$



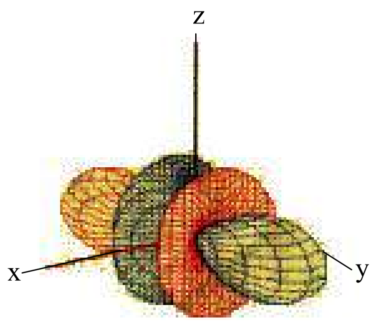
$f_{x(y^2-z^2)}$



$f_{y(z^2-x^2)}$



f_{x^3}



f_{y^3}

EXERCISE # O-1

General Introduction :

1. The total number of neutrons in dipositive zinc ion with mass number 70 is

(A) 34 (B) 40 (C) 36 (D) 38

2. It is known that atom contain protons, neutrons and electrons. If the mass of neutron is assumed to half of its original value where as that of proton is assumed to be twice of its original value then the atomic mass of ${}^{14}_{6}C$ will be -

(A) same (B) 25% more (C) 14.28 % more (D) 28.5% less

3. Two monoatomic cations $x^{\oplus}$ and y^{2+} are isoelectronic then select the correct statement : (Both elements are consecutive)

(A) Both element x and y have same number of electrons

(B) Total number of valence electrons are more in element x, than element 'y'

(C) Total number of valence electrons are more in element y, than element 'x'

(D) Both (A) and (B) are correct

4. Which of the following sets contain only isoelectronic ions?

(A) Zn^{2+}, Ca^{2+}, Ga^{3+} , Al^{3+} (B) K^{+}, Ca^{2+}, Sc^{3+}, Cl^{-}

(C) P^{3-}, S^{2-}, Cl^{-}, Zn^{+2} (D) Ti^{4+}, Ar, Cr^{3+}, V^{5+}

Quantum number

5. Which quantum number will determine the shape of the orbital

(A) Principal quantum number (B) Azimuthal quantum number

(C) Magnetic quantum number (D) Spin quantum number

6. In Palladium (Atomic no.–46), number of electron having ($\ell = 2$) will be -

(A) 20 (B) 18 (C) 16 (D) 22

7. For an electron present in which of the following orbital for which $(n + l + m + s)$ value is maximum. Consider maximum possible value for 'm' and minimum possible value of m_s (where ever applicable).

(A) 3p (B) 5p (C) 4d (D) 5s

8. Choose the correct option for the quantum numbers of the last electron of K^{+}.

(A) 4, 0, 0, +1/2 (B) 3, 1, –1, –1/2 (C) 4, 1, 0, $-\frac{1}{2}$ (D) 3, 0, 1, $\frac{1}{2}$

9. Find the sum of maximum number of electrons having +1 and –1 value of 'm' in Ti (Atomic number = 22)

(A) 6 (B) 8 (C) 10 (D) 12

10. The number of electrons in Ca having minimum value of $\left|\frac{n}{\ell \times m_{\ell}}\right|$ is.

(Consider only non-zero values of ℓ and m)

(A) 6 (B) 3 (C) 4 (D) None of these

Electronic Configuration

11. A neutral atom of an element has two K, eight L, nine M and two N electrons then electronic configuration of the element is

(A) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^1$
(B) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$
(C) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3d^2\ 3p^6\ 4s^1$
(D) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2$

12. The explanation for the presence of three unpaired electrons in the nitrogen atom can be given by

(A) Pauli's exclusion principle
(B) Hund's rule
(C) Aufbau's principle
(D) Uncertainity principle

13. If the nitrogen atom had electronic configuration $1s^7$, it would have energy lower than that of normal ground state configuration $1s^2\ 2s^2\ 2p^3$, because the electrons would be closer to the nucleus. Yet $1s^7$ is not observed because it violates :–

(A) Uncertainity principle
(B) Hund's rule
(C) Pauli's exclusion principle
(D) Bohr postulate of stationary orbits

Effective Nuclear charge (Z_{eff}) :

14. The Z_{eff} for (as Slater's rule)

3d electron of Cr

4s electron of Cr

3d electron of Cr^{3+}

3s electron of Cr^{3+} are in the order respectively

(A) 4.6, 2.95, 4.95, 8.05
(B) 4.95, 2.95, 4.6, 8.05
(C) 4.6, 2.95, 5.3, 12.75
(D) none of these

15. Total number of possible shells in uranium atom (atomic no. z = 92)

(A) 7
(B) 1
(C) 6
(D) None of these

16. Which of the following has the maximum number of unpaired electrons ?

(A) Mg^{2+}
(B) Ti^{3+}
(C) V^{3+}
(D) Fe^{2+}

17. Gaseous state electronic configuration of nitrogen atom can be represented as :

(A) ↑↓ ↑↓ ↑ ↑ ↑
(B) ↑↓ ↑↓ ↑ ↓ ↑
(C) ↑↓ ↑↓ ↑ ↓ ↓
(D) ↑↓ ↑↓ ↓ ↓ ↑

EXERCISE : O-2

General Introduction :

1. Isotones are :

(A) The atoms of different elements

(B) Have same number of neutrons

(C) Have same number of (neutrons + protons)

(D) Have same difference of mass number and atomic number

Quantum number

2. For an electron present in which of the following orbital for which $(n + l)$ value is maximum.

(A) 3p (B) 5p (C) 4d (D) 5s

3. Correct set of four quantum numbers for valence electron of rubidium(Z = 37) is

(A) $5, 0, 0, +\frac{1}{2}$ (B) $5, 0, 0, -\frac{1}{2}$ (C) $5, 1, 1, +\frac{1}{2}$ (D) $6, 0, 0, +\frac{1}{2}$

4. The correct set of quantum numbers for the unpaired electron of chlorine atom is

	n	**ℓ**	**m**		**n**	**ℓ**	**m**
(A)	2	1	0	(B)	2	1	1
(C)	3	1	1	(D)	3	1	0

5. Which of the following sets of quantum numbers represent an impossible arrangement ?

	n	**ℓ**	**m**	**m_s**		**n**	**ℓ**	**m**	**m_s**
(A)	3	3	–2	$\frac{1}{2}$	(B)	4	0	0	$\frac{1}{2}$
(C)	3	2	–3	$\frac{1}{2}$	(D)	5	3	0	$\frac{1}{2}$

6. The quantum numbers for the 19th electron of Cr (Z = 24) are

(A) $n = 3, \ell = 0, m = 0, m_s = +\frac{1}{2}$ (B) $n = 4, \ell = 0, m = 0, m_s = +\frac{1}{2}$

(C) $n = 3, \ell = 2, m = 2, m_s = +\frac{1}{2}$ (D) $n = 4, \ell = 0, m = 0, m_s = -\frac{1}{2}$

7. The maximum number of electron having $n \times \ell \times m = 0$ in Zn^{2+} is equal to the -

(A) Atomic number of Mg

(B) 12

(C) Total number of electron in Zn which have $n + \ell = 0$

(D) 'p' electrons in Ar

Electronic Configuration

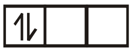
8. The species which have same number of electrons in outer most and penultimate shell -

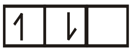
(A) Ca (B) Ar (C) V^{+3} (D) Sc^{3+}

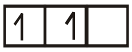
9. Which sub-shell fill completely before the 4f?

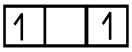
(A) 6s (B) 5p (C) 5d (D) 4d

10. The electronic configuration of a carbon atom is $1s^2$, $2s^2$, $2p^2$ and consider the following four arrangements of the 2p electrons. Which arrangement have lowest energy ?

(A) 

(B) 

(C) 

(D) 

11. Hund's rule is applicable for :-

(A) d-subshell (B) p-subshell (C) s-subshell (D) f-subshell

12. Which of the following has maximum number of unpaired electron.

(A) Fe (B) Fe (II) (C) Fe (III) (D) Mn (II)

13. Mn (Z = 25) = $1s^3\ 1p^9\ 2s^3\ 2p^9\ 2d^1$

Which of the following change is required so that Mn have above ground state electronic configuration :

(A) Change in the value of ℓ (azimuthal quantum number) for any subshell

(B) Change in the possible values of ℓ (azimuthal quantum number)

(C) Change in the Pauli rule

(D) Change in the $(n + \ell)$ rule

14. The number of d- electrons in Mn^{2+} is equal to that of

(A) p-electrons in N (B) s-electron in Na

(C) d-electrons in Fe^{+3} (D) p-electrons in O^{-2}

15. Select incorrect statement(s) :

(A) d_{z^2} orbital has different shape from rest of all d-orbitals

(B) For the formation of cation electrons are always removed from 4s.

(C) Zinc is a p-block element.

(D) Principal quantum number depend upon the value of azimuthal quantum number

EXERCISE # S-1

Integer Answer Type (0 to 9) :

1. Find total no. of orbitals in nickel which have $|m| \leq 1$ and at least one electron is present, where 'm' is magnetic quantum number.

(Given your ans. as sum of digits for example. If your ans is 57 then 5 + 7 = 12 and 1 + 2 =3)

2. Minimum number of electrons having $m_s = \left(-\frac{1}{2}\right)$ in Cr is "_____".

3. How many elements are possible for the Ist period of periodic table if azimuthal quantum number can have integral values from 0 to (n + 1). [n = shell number & other rules are remaning same to form periodic table.

4. Find number of unpaired electrons when Fe does not follow $(n + \ell)$ rule and filling of electron takes palce shell after shell and Hund's rule is also not obeyed.

5. Find the maximum number of electrons having same Z_{eff} value for sulphur atom

6. Find the sum of maximum unpaired e^- present in one 5g & one 6g orbital.

7. Find out the maximum number of electrons that can involve in the shielding of an electron, having quantum numbers : $n = 2, \ell = 1, m = 0, m_s = +\frac{1}{2}$, in an atom.

8. Find the sum of minimum and maximum possible value of x in Fe^{+x} ion, if magnetic moment of Fe^{+x} = 4.89 B.M.

EXERCISE # S-2

Paragraph for Question 1 to 3

The general electronic configuration of outer most and penultimate shell is given as $(n-1)s^2$ $(n-1)p^6$ $(n-1)d^x ns^2$. Then for an element with $n = 4$ and $x = 6$.

1. The number of protons present in the divalent cation of the element of above configuration is :-

(A) 24 (B) 25 (C) 26 (D) 27

2. The element is :

(A) Mn (B) Fe (C) Co (D) Li

3. The number of unpaired electrons in the divalent cation of the given element in isolated gaseous state is :-

(A) 0 (B) 3 (C) 4 (D) 1

Paragraph for Question 4 & 5

Consider a hypothetical atom where p_x, p_y, p_z, d_{xy}, d_{xz}, d_{yz} and $d_{x^2-y^2}$ orbitals are present for principal quantum number $n = 4$.

4. Find the number of other orbital which lobes are fully present in the nodal plane of p_x orbital :-

(A) 2 (B) 4 (C) 3 (D) 5

5. Which of the following orbitals lobe is not present at all either in the nodal plane of p_x orbital or in the nodal plane of p_y orbital.

(A) d_{xy} (B) d_{yz} (C) $d_{x^2-y^2}$ (D) p_z

Paragraph for Question 6 & 7

Isotopes, Isobars and Iso-diaphers are some basic definitions related to the atom, which are based upon the variation in no. of electron, proton or neutrons.

6. Which of the following pair represents the Isobar species :-

(A) ${}_{18}Ar^{40}$, ${}_{6}C^{12}$ (B) ${}_{8}O^{16}$, ${}_{7}N^{14}$ (C) N_2O, CO_2 (D) None of these

7. Which of the following pair is correct for isosters :-

(A) N_2O, CH_4 (B) N_2O, CO_2 (C) CO_2, SO_2 (D) N_2O, HOCl

8.

Column-I **Element**	**Column-II** **Max. value of n and ℓ respectively ; consider filled subshell only**
(P) P	(1) 6, 2
(Q) N	(2) 3, 1
(R) Pb	(3) 6, 3
(S) Cs	(4) 2, 1

Code :

	P	**Q**	**R**	**S**		**P**	**Q**	**R**	**S**
(A)	4	1	2	3	(B)	2	4	1	3
(C)	3	1	2	4	(D)	2	4	3	1

9. Match the following:

Column–I	Column–II
(P) Same number of unpaired electrons are present in (excluding zero)	(1) Na^+, Mg^{2+}, F^-
(Q) Same number of electrons in s & p subshells.	(2) F^-, Mg, O^{2-}
(R) Same number of electrons with the $l = 1$	(3) Mg, Ne, O^{2-}
(S) Same number of total electrons	(4) Li, Na, K

Code :

	P	Q	R	S
(A)	1, 2	3, 4	1	4
(B)	4, 2	3, 1	2, 3	4, 1
(C)	4	1	1, 2, 3	1
(D)	3	1, 2	3, 4	1, 3, 4

10. Match the following :

Column–I (Orbital)	Column–II (Property)
(P) s	(1) Have electron density at all three axes
(Q) p_x	(2) YZ plane is nodal plane
(R) d_{xy}	(3) dumbell shape
(S) $d_{x^2-y^2}$	(4) have azimuthal quantum no. $\ell = 2$

Code :

	P	Q	R	S
(A)	1, 3	2, 3	3	3, 4
(B)	1	2, 3	2, 4	4
(C)	1	2, 3, 4	3, 4	1, 4
(D)	2	3, 4	2, 3	1, 4

Match The Column :

11. Match the following :

Column–I (e^- configuration)	Column–II (Property)
(A) d^8	(P) Symmetrical distribution
(B) d^{10}	(Q) Unsymmetrical distribution
(C) d^6	(R) No of exchange pair are maximum among these
(D) d^5	(S) two electrons must be present in $d_{x^2-y^2}$ orbital
	(T) at least one electron is present in orbital having $m = -1$

12.

Column-I	Column-II
(A) Zn^{2+}	(P) Diamagnetic
(B) Ga^+	(Q) Spin magnetic moment = $\sqrt{35}$ BM
(C) Fe^{3+}	(R) 18 e^- in outer most shell
(D) Br^-	(S) 3d subshell is fully filled
	(T) All the orbital of outer most shell are fully filled

MATCHING LIST TYPE 1 × 3 Q. (THREE LIST TYPE Q.)

Column - I	Column - II	Column - III
(A) Paramagnetic set	(i) Na^+, Mg^{+2}, F^-	(P) same value of principal quantum number for last electron
(B) Isoelectronic set	(ii) Li, Na, K	(Q) The non zero number(s) of e^-(s) for $n = 3$ and $\ell \geq 1$ is
(C) The set for which value(s) of spin multiplicity is ≥ 1	(iii) Fe^{+3}, Co^{2+}, Ni^{+2}	(R) The value of "m_s" must be $+\frac{1}{2}$ for last electron
(D) The set of elements belongs to same period in periodic table	(iv) S^{-2}, Cl^-, P^{-3}	(S) Set for which the value of $m = \pm 2$ is possible for electron(s)

13. Which one of the following options is the **CORRECT** combination?

(A) (A, i, P) (B) (B, iv, S) (C) (D, iii, S) (D) (C, iii, R)

14. Which one of the following options is the **INCORRECT** combination?

(A) (A, iii, P) (B) (C, ii, P) (C) (B, i, P) (D) (B, iv, Q)

15. Which one of the following options is the **CORRECT** set of species with number of nodal planes for filled/partially filled orbitals is $\leq$ 1 for all given species in set?

(A) (B, ii, R) (B) (B, iv, P) (C) (A, i, Q) (D) (D, iii, S)

EXERCISE # JEE-MAIN

1. Which of the following groupings represent a collection of isoelectronic species ? (At. no. Cs = 55, Br = 35) **[AIEEE-2003]**

(1) N^{3-}, F^-, Na^+
(2) Be, Al^{3+}, Cl^-
(3) Ca^{2+}, Cs^+, Br
(4) Na^+, Ca^{2+}, Mg^{2+}

2. The number of d-electrons retained in Fe^{2+} (At. no. of Fe = 26) ion is : **[AIEEE-2003]**

(1) 6
(2) 3
(3) 4
(4) 5

3. Which of the following sets of quantum number is correct for an electron in 4f orbital ? **[AIEEE-2004]**

(1) $n = 3, \ell = 2, m = -2, s = +\frac{1}{2}$
(2) $n = 4, \ell = 4, m = -4, s = -\frac{1}{2}$
(3) $n = 4, \ell = 3, m = +1, s = +\frac{1}{2}$
(4) $n = 4, \ell = 3, m = +4, s = +\frac{1}{2}$

4. Consider the ground state of Cr atom (Z = 24). The numbers of electrons with the azimuthal quantum numbers, $l = 1$ and 2 are, respectively. **[AIEEE-2004]**

(1) 16 and 5
(2) 12 and 5
(3) 16 and 4
(4) 12 and 4

5. Which one of the following sets of ions represents the collection of isoelectronic species? **[AIEEE-2004]**

(1) Na^+, Mg^{2+}, Al^{3+}, Cl^-
(2) Na^+, Ca^{2+}, Sc^{3+}, F^-
(3) K^+, Cl^-, Mg^{2+}, Sc^{3+}
(4) K^+, Ca^{2+}, Sc^{3+}, Cl^-

6. In a multi-electron atom, which of the following orbitals described by the three quantum members will have the same energy in the absence of magnetic and electric fields ? **[AIEEE-2005]**

(A) $n = 1, l = 0, m = 0$
(B) $n = 2, l = 0, m = 0$
(C) $n = 2, l = 1, m = 1$
(D) $n = 3, l = 2, m = 1$
(E) $n = 3, l = 2, m = 0$

(1) (D) and (E)
(2) (C) and (D)
(3) (B) and (C)
(4) (A) and (B)

7. Of the following sets which one does not contain isoelectronic species ? **[AIEEE-2005]**

(1) BO_3^{3-}, CO_3^{2-}, NO_3^-
(2) SO_3^{2-}, CO_3^{2-}, NO_3^-
(3) CN^-, N_2, C_2^{2-}
(4) PO_4^{3-}, SO_4^{2-}, ClO_4^-

8. Which of the following statements in relation to the hydrogen atom is correct ? **[AIEEE-2005]**

(1) 3s, 3p and 3d orbitals all have the same energy
(2) 3s and 3p orbitals are of lower energy than 3d orbitals
(3) 3p orbital is lower in energy than 3d orbital
(4) 3s orbitals is lower in energy than 3p orbital

9. Which one of the following sets of ions represents a collection of isoelectronic species ? **[AIEEE-2006]**

(1) N^{3-}, O^{2-}, F^-, S^{2-}
(2) Li^+, Na^+, Mg^{+2}, Ca^{+2}
(3) K^+, Cl^-, Ca^{+2}, Sc^{+3}
(4) Ba^{+2}, Sr^{+2}, K^{+2}, Ca^{+2}

10. Which of the following sets of quantum numbers represents the highest energy of an atom ? **[AIEEE-2007]**

(1) $n = 3, \ell = 1, m = \ell, s = +½$
(2) $n = 3, \ell = 2, m = \ell, s = +½$
(3) $n = 4, \ell = 0, m = 0, s = +½$
(4) $n = 3, \ell = 0, m = 0, s = +½$

11. Which one of the following constitutes a group of the isoelectronic species? **[AIEEE-2008]**

(1) C_2^{2-}, O_2^-, CO, NO (2) $NO^+, C_2^{2-}, CN^-, N_2$

(3) $CN^-, N_2, O_2^{2-}, C_2^{2-}$ (4) N_2, O_2^-, NO^+, CO

12. The electrons identified by quantum numbers n and ℓ :-

[JEE-1999, AIEEE-2012, JEE-MAIN, (ONLINE)-2012]

(a) $n = 4$, $\ell = 1$ (b) $n = 4$, $\ell = 0$ (c) $n = 3$, $\ell = 2$ (d) $n = 3$, $\ell = 1$

Can be placed in order of increasing energy as

(1) (a) < (c) < (b) < (d) (2) (c) < (d) < (b) < (a)

(3) (d) < (b) < (c) < (a) (4) (b) < (d) < (a) < (c)

13. Which of the following paramagnetic ions would exhibit a magnetic moment (spin only) of the order of 5 BM ? **[JEE-MAIN, (ONLINE)-2012]**

(At. No : Mn = 25, Cr = 24, V = 23, Ti = 22)

(1) V^{2+} (2) Ti^{2+} (3) Mn^{2+} (4) Cr^{2+}

14. In an atom how many orbital (s) will have the quantum numbers; $n = 3$, $l = 2$ and $m_l = +2$? **[JEE-MAIN, (ONLINE)-2013]**

(1) 1 (2) 5 (3) 3 (4) 7

15. The numbers of protons, electrons and neutrons in a molecule of heavy water are respectively

(1) 10, 10, 10 (2) 8, 10, 11 **[JEE-MAIN, (ONLINE)-2013]**

(3) 10, 11, 10 (4) 11, 10, 10

16. Given **[JEE-MAIN, (ONLINE)-2013]**

(a) $n=5$, $m_l = +1$ (b) $n = 2$, $l = 1$, $m_l = -1$, $m_s = -1/2$

The maximum number of electron(s) in an atom that can have the quantum numbers as given in (a) and (b) are respectively :

(1) 8 and 1 (2) 25 and 1 (3) 2 and 4 (4) 4 and 1

17. The correct set of four quantum numbers for the valence electrons of rubidium atom (Z = 37) is:

[JEE(Main)-2014]

(1) $5,1,1,+\frac{1}{2}$ (2) $5,0,1,+\frac{1}{2}$ (3) $5,0,0,+\frac{1}{2}$ (4) $5,1,0,+\frac{1}{2}$

18. If the principal quantum number n = 6, the correct sequence of filling of electrons will be:-

[JEE-MAIN, (ONLINE)-2015]

(1) ns→(n–1)d→(n–2)*f*→np (2) ns→np→(n–1)d→(n–2)*f*

(3) ns→(n–2)*f*→(n–1)d→np (4) ns→(n–2)*f*→np→(n–1)d

19. The total number of orbitals associated with the principal quantum number 5 is :

[JEE-MAIN, (ONLINE)-2016]

(1) 25 (2) 5 (3) 20 (4) 10

EXERCISE # JEE-ADVANCED

1. The electronic configuration of an element is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$. This represents its : **[JEE 2000]**

(A) excited state (B) ground state (C) cationic form (D) none

2. The maximum number of electrons that can have principal quantum number, $n = 3$, and spin quantum number, $m_s = -1/2$, is **[JEE 2011]**

3. In an atom, the total number of electrons having quantum numbers $n=4$, $|m_\ell| = 1$ and $m_s = -\frac{1}{2}$ is:

[JEE Advanced 2014]

4. Not considering the electronic spin the degeneracy of the second excited state ($n = 3$) of H-atom is 9, where the degeneracy of the second excited state of H^- is **[JEE Advanced 2015]**

ANSEWR KEY

EXERCISE # (O–1)

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
Ans.	B	C	C	B	B	A	C	B	C	C
Que.	**11**	**12**	**13**	**14**	**15**	**16**	**17**			
Ans.	B	B	C	C	D	D	A			

EXERCISE # (O–2)

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
Ans.	A, B, D	B, C	A, B	C, D	A, C	B, D	A, B, D	B, D	A, B, D	C, D
Que.	**11**	**12**	**13**	**14**	**15**					
Ans.	A, B, D	C, D	B, C, D	B, C	B, C, D					

EXERCISE # (S–1)

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**
Ans.	4	9	8	0	8	2	9	6

EXERCISE # (S–2)

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
Ans.	C	B	C	C	A	D	B	D	C	B
Que.	**11**					**12**				
Ans.	(A)-Q,T (B)-P,R,S,T (C)-Q,T (D)-P,T					(A)-P,R,S,T (B)-P,S (C)-Q (D)-P,S				
Que.	**13**	**14**	**15**							
Ans.	C	B	B							

EXERCISE # JEE-MAIN

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
Ans.	1	1	3	2	4	1	2	1	3	2
Que.	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	
Ans.	2	3	3	1	1	1	3	3	1	

EXERCISE # JEE-ADVANCED

Que.	**1**	**2**	**3**	**4**
Ans.	B	9	6	3

PART-A (PERIODIC TABLE)

INTRODUCTION :

The arrangement of all the known elements according to their properties in such a way that the elements of similar properties are grouped together in a tabular form is called periodic table.

Development of periodic table :

(A) LAVOISIER CLASSIFICATION OF ELEMENTS

At first Lavoisier classified the elements into two categories.

(i) Metal (ii) Non-metal

Note : This classification cannot categories the metalloid variety.

(B) DOBEREINER'S TRIAD RULE [1817]

(i) He made groups of three elements having similar chemical properties called TRIAD.

(ii) In Dobereiner triad, atomic weight of middle element is nearly equal to the average atomic weight of first and third element.

e.g.

Cl	Br	I	
35.5	$\boxed{80.0}$	127	$\frac{35.5+127}{2}=81.25$
Ca	Sr	Ba	
40	$\boxed{87.5}$	137	$\frac{40+137}{2}=88.5$
Li	Na	K	
7	$\boxed{23}$	39	At. wt of $\text{Na}=\frac{7+39}{2}=23$

(iii) Other examples – (K, Rb, Cs), (P, As, Sb), (S, Se, Te), (H, F, Cl), (Sc, Y, La)

(C) NEWLAND'S OCTAVE RULE [1865]

(i) He arranged the elements in the increasing order of their atomic mass and observed that properties of every 8th element was similar to the 1st one, like in the case of musical vowels notation.

Sa	Re	Ga	Ma	Pa	Dha	Ni	Sa
1	2	3	4	5	6	7	8

(ii) At that time inert gases were not known.

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

(iii) The properties of Li are similar to 8th element i.e. Na, Be are similar to Mg and so on.

Drawbacks or Limitations :

(a) This rule is valid only upto Ca. because after Ca due to filling of d-orbitals there is difference of 18 elements instead of 8 element.

(b) After the discovery of inert gas this law had to be dropped out.

(c) He failed in the case of heavier metals as Fe has been placed along with O and S.

(D) LOTHER MEYER'S CURVE [1869]

(i) He plotted a curve between atomic weight and atomic volume of different elements.

(ii) The following observations can be made from the curve –

(a) Most electropositive elements i.e. alkali metals (Li, Na, K, Rb, Cs etc.) occupy the peak portions of the curve.

(b) Less electropositive i.e. alkaline earth metals (Be, Mg, Ca, Sr, Ba) occupy the descending portions of the curve.

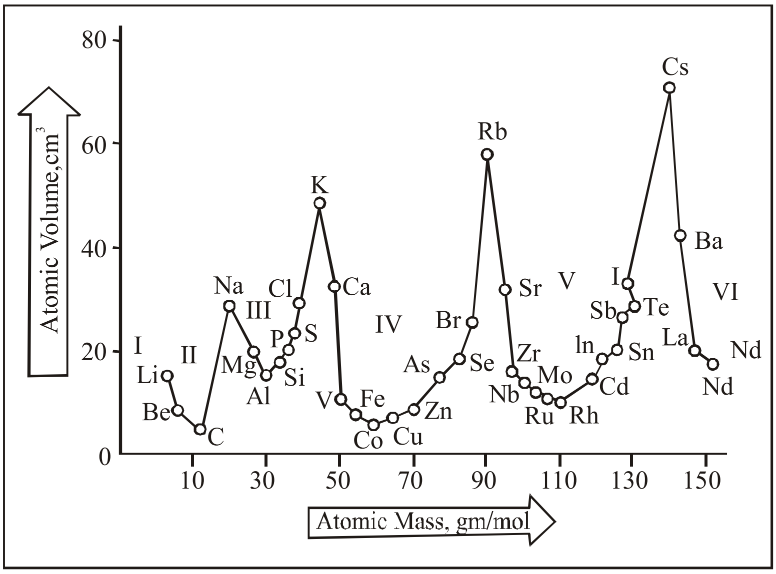
(c) Metalloids (B, Si, As, Te, At etc.) and transition metals occupy bottom part of the curve.

(d) Most electronegative i.e. halogens (F, Cl, Br, I) occupy the ascending portion of the curve.

Note : Elements having similar properties occupy similar portions of the curve.

Conclusion :

On the basis of this curve, Lother Meyer proposed that the physical properties of the elements are periodic function of their atomic weight and this became the basis of Mendeleev's periodic table.



(E) MENDELEEV'S PERIODIC TABLE [1869]

(i) ***Mendeleev's periodic law –*** The physical and chemical properties of elements are the periodic function of their atomic weight

(ii) ***Characteristic of Mendeleev's periodic table –***

(a) It is based on atomic weight

(b) 63 elements were known, noble gases were not discovered.

(c) He was the first scientist to classify the elements in a systamatic manner i.e. in horizontal rows and in vertical columns.

(d) Horizontal rows are called periods and there were 7 periods in Mendeleev's Periodic table.

(e) Vertical columns are called groups and there were 8 groups in Mendeleev's Periodic table.

(f) Each group upto VIIth is divided into A & B subgroups.'A' sub group elements are called normal elements and 'B' sub group elements are called transition elements.

(g) The VIIIth group consists of 9 elements in three rows.

(h) The elements belonging to same group exhibit similar properties.

(iii) Merits or advantages of Mendeleev's periodic table –

(a) Study of elements – For the first time, all known elements were classified in groups according to their similar properties. So study of the properties of elements became easier.

(b) Prediction of new elements – It gave encouragement to the discovery of new elements as some gaps were left in it.

Sc (Scandium), Ga (Gallium), Ge (Germanium), Tc (Technetium)

were the elements whose position and properties were well defined by Mendeleev even before their discoveries and he left the blank spaces for them in his table.

e.g. - Blank space at atomic weight 72 in silicon group was called Eka silicon (means properties like silicon) and element *(discovered later)* was named Germanium .

Similarly other elements discovered after mendeleev periodic table were :

Eka aluminium – Gallium(Ga)
Eka Boron – Scandium (Sc)
Eka Silicon – Germanium (Ge)
Eka Manganese – Technetium (Tc)

(c) Correction of doubtful atomic weights–Corrections were done in atomic weight of some elements.

$$\boxed{\text{Atomic Weight} = \text{Valency} \times \text{Equivalent weight.}}$$

Initially, it was found that equivalent weight of Be is 4.5 and it is trivalent (V = 3), so the weight of Be was 13.5 and there is no space in Mendeleev's table for this element. So, after correction, it was found that Be is actually divalent (V = 2). So, the weight of Be became 2 × 4.5 = 9 and there was a space between Li and B for this element in Mendeleev's table.

Corrections were done in atomic weight of elements are – U, Be, In, Au, Pt.

(iv) Demerits of Mendeleev's periodic table –

(a) Position of hydrogen – Hydrogen resembles both, the alkali metals (IA) and the halogens (VIIA) in properties so Mendeleev could not decide where to place it.

(b) Position of isotopes – As atomic weight of isotopes differs, they should have been placed in different position in Mendeleev's periodic table. But there was no such place for isotopes in Mendeleev's table.

(c) Anomalous pairs of elements – There were some pair of elements which did not follow the increasing order of atomic wts.

eg : Ar and Co were placed before K and Ni respectively in the periodic table, but having higher atomic weights.

$$\underset{\text{in modified mendeleev P.T.}}{\begin{pmatrix} \text{Ar} & \text{K} \\ 39.9 & 39.1 \end{pmatrix}} \qquad \begin{pmatrix} \text{Te} & \text{I} \\ 127.5 & 127 \end{pmatrix}$$

$$\begin{pmatrix} \text{Co} & \text{Ni} \\ 58.9 & 58.6 \end{pmatrix} \qquad \begin{pmatrix} \text{Th} & \text{Pa} \\ 232 & 231 \end{pmatrix}$$

(d) Like elements were placed in different groups.

There were some elements like Platinum (Pt) and Gold (Au) which have similar properties but were placed in different groups in Mendeleev's table.

Pt	Au
VIII	IB

(e) Unlike elements were placed in same group.

Ist group

	IA	IB	
	Li		
	Na		
More reactive	K	Cu	Less reactive
Alkali metal	Rb	Ag	Coinage metal
Normal elements	Cs	Au	Transition element
	Fr		

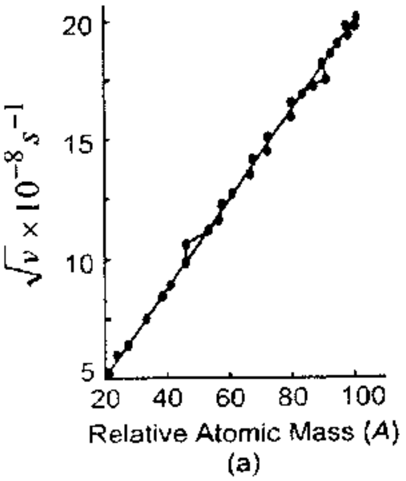
Cu, Ag and Au placed in Ist group along with Na, K etc. While they differ in their properties (Only similar in having ns^1 electronic configuration)

(F) MODERN PERIODIC TABLE (MODIFIED MENDELEEV PERIODIC TABLE)

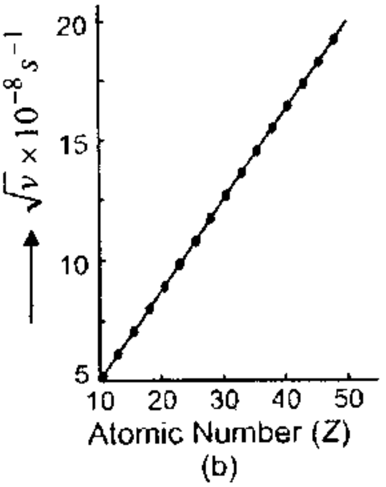
(i) It was proposed by **Moseley (1913)**.

(ii) Modern periodic table is based on atomic number.

(iii) Moseley did an experiment in which he bombarded high speed electron on different metal surfaces and obtained X-rays.



(a)



(b)

He found out that $\sqrt{\nu} \propto Z$ (where ν = frequency) of X-rays from this experiment, Moseley concluded that the physical and chemical properties of the elements are periodic function of their atomic number. It means that when the elements are arranged in the increasing order of their atomic number, elements having similar properties gets repeated after a regular interval. This is also known as **'Modern Periodic Law'**.

(iv) Modern periodic law – The physical & chemical properties of elements are a periodic function of their atomic number.

(v) Characteristics of modern periodic table –

(a) 9 vertical columns called groups.

(b) IA to VIIA, IB to VIIB, VIII and 0

(c) Inert gases were introduced in periodic table by Ramsay.

(d) 7 horizontal series called periods.

(G) LONG FORM / PRESENT FORM OF MODERN PERIODIC TABLE

(It is also called as 'Bohr-Burry & **Rang**, Werner Periodic Table.)

(i) It is based on the Bohr-Burry electronic configuration concept and atomic number.

(ii) This model was proposed by Rang & Werner

(iii) It consists of 7 horizontal periods and 18 vertical columns (groups)

(iv) According to I. U. P. A. C. 18 vertical columns are named as 1st to 18th group.

(v) The co-relation between the groups in long form of periodic table and in modern form of periodic table are given below.

IA,	IIA,	IIIB,	IVB,	VB,	VIB,	VIIB,	VIII,	IB,	IIB,	IIIA,	IVA,	VA,	VIA,	VIIA,	0
1	2	3	4	5	6	7	8, 9, 10	11	12	13	14	15	16	17	18

(vi) Elements belonging to same group have same no. of electrons in the outermost shell so their properties are similar.

Description of periods :

Period	n	Sub shell	No. of elements	Element	Name of Period
1.	1	1s	2	$_{1}$H – $_{2}$He	Shortest
2.	2	2s, 2p	8	$_{3}$Li – $_{10}$Ne	Short
3.	3	3s, 3p	8	$_{11}$Na – $_{18}$Ar	Short
4.	4	4s, 3d, 4p	18	$_{19}$K – $_{36}$Kr	Long
5.	5	5s, 4d, 5p	18	$_{37}$Rb – $_{58}$Xe	Long
6.	6	6s, 4f, 5d, 6p	32	$_{55}$Cs – $_{86}$Rn	Longest
7.	7	7s, 5f, 6d,	26	$_{87}$Fr – $_{112}$Uub	Incomplete

CLASSIFICATION OF ELEMENTS INTO s , p , d & f-BLOCK ELEMENTS :

s – block :

(i) Configuration ns^{1-2} **(ii)** Last e^- enters in s orbital

(iii) Two groups I A or 1 ; II A or 2

p – block :

(i) Configuration $ns^2\ np^{1-6}$ **(ii)** Last e^- enters in p orbital

(iii) Six groups III A , IV A , V A , VI A , VII A , zero or 13 , 14 , 15 , 16 , 17 , 18

d – block : [Transition Elements]

(i) Configuration $(n-1)\ d^{1-10}\ ns^{0-2}$ **(ii)** Last e^- enters in d orbital

(iii) Their two outermost shell are incomplete

(iv) 10 groups III B , IV B , V B , VI B , VII B , VIII (Triad) , I B , II B or 3 , 4 , 5 , 6 , 7 , (8 , 9 , 10) , 11 , 12 .

(v) Four series 3 d , 4 d , 5 d , 6 d belong to 4th, 5th, 6th & 7th period respectively in long form of Periodic table.

f – block : [Inner Transition]

(i) configuration $(n-2)\,f^{0-14}\,(n-1)\,d^{0-2}\,ns^2$

(ii) last e^- enters in f orbital

(iii) two series 4 f Lanthanides & 5 f Actinides belong to 6th & 7th period respectively in long form of Periodic table.

Neil Bohr's classification of elements :

Using electronic configuration as the criteria, elements are of four types. The classification of the elements into these groups is dependent on the extent to which the s, p, d and f orbitals are filled.

Inert Gases :

(a) s – and p-orbitals of the outer most shell of these elements are completely filled. The outermost electronic configuration is ns^2np^6.

(b) Helium is also inert gas but its electronic configuration is $1s^2$

Representative or Normal Elements :

(a) Outermost shell of these elements is incomplete. The number of electrons in the outermost shell is less than eight.

(b) s-and p-block elements except inert gases are called normal or representative elements.

Transition Elements :

(a) These metals were placed between s-block metals and p-block elements so, are named transition metals.

(b) Their outermost electronic configuration is similar to d-block elements i.e. $(n-1)\,d^{1\text{-}10}\,ns^{1\text{-}2}$.

(c) Last two shells of these elements namely outermost and penultimate shells are incomplete.

(d) The last shell contains one or two electrons and the penultimate shell may contain more than eight and up to eighteen electrons.

(e) According to definition of transition elements, those elements which have partly filled d-orbitals in neutral state or in any stable oxidation state are called transition elements. According to this definition Zn, Cd and Hg (IIB group) are d-block elements but not transition elements because these elements have d^{10} configuration in neutral as well as in stable +2 oxidation state.

Inner Transition Elements :

(a) In these elements last three shells i.e. last, penultimate and prepenultimate shells are incomplete.

(b) These are related to IIIB i.e. group 3.

(c) The last shell contains two electrons. Penultimate shell may contain eight or nine electrons and pre-penultimate shell contains more than 18, up to 32 electrons.

(d) Their outemost electronic configuration is similar to f-block element

i.e. $(n-2)\,f^{0-14}\,(n-1)\,d^{0-2}\,ns^2$

METALS, NON-METALS & METALLOIDS

Apart from classifying elements into s, p, d and *f*-blocks, there is yet another broad classification of elements based on their properties. The elements can be broadly classified into

(a) Metals:

Majority of the elements in periodic table are metals and appears on the left side of the periodic table.

Properties:

(i) These are usually solid at room temperature [exception - mercury]

(ii) They have high melting and boiling point [exception Gallium & Cesium have very low melting point (303 K and 302 K respectively)]

(iii) They are good conductor of heat and electricity.

(iv) They are malleable (can be flattened into thin sheets by hammering) and ductile (can be drawn into wires)

(b) Non-Metals:

These are placed at the top right hand side of periodic table. As we move horizontally along a period, the property of elements changes from metallic (on left) to non-metallic (on the right).

Properties :

(i) These are usually solids or gases at room temperature.

(ii) They have low melting point and boiling point (exception : Boron, Carbon).

(iii) Most Non-metallic solids are brittle and are neither malleable nor ductile.

(c) Metalloids (Semi-metals):

Properties of these elements show the characteristics of both metals and non-metals. Silicon (Si), Germanium(Ge), Arsenic(As), Antimony(Sb) and Tellurium(Te) are metalloids.

ESTIMATING POSITION OF AN ELEMENT FROM ITS ELECTRONIC CONFIGURATION

The last electron enters in which subshell gives idea of its block.

[☺ ***Think*** :$1s^1$ and $1s^2$ belongs to which block]

Period number = Principal quantum number of valence shell electron in ground state electronic configuration.

Group number for s block = number of valence shell electrons

Group number for p block = 10 + number of valence shell electrons

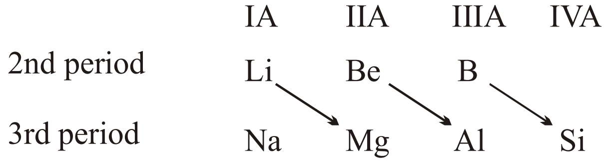
Group number for d block = number of [ns + (n–1) d] electrons

Group number for f-block = 3

[☺ Use these carefully while locating the position.]

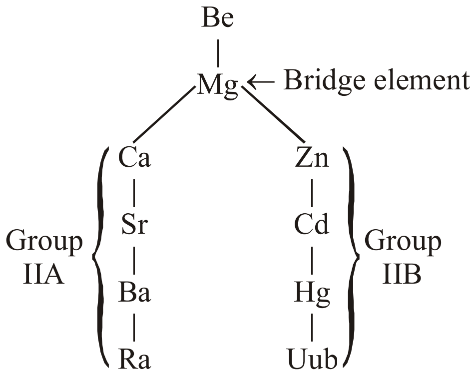
SOME COMMONLY USED TERMS

1. ***Noble Gases :*** Element of group 18 are called noble gases. These are also called as inert gases because their outermost ns and np orbitals are completely filled (except He and $1s^2$) and these gases are non-reactive in nature under ordinary conditions.
2. ***Typical elements :*** Elements of second and third period are known as typical elements.
3. ***Diagonal relationship :*** Properties of elements of second period resemble with the element of third period. This resemblance between properties of 2^{nd} & 3^{rd} period is called diagonal relationship.



4. ***Bridge elements :***

The typical elements of third period are also called bridge elements as the division between two subgroups A and B starts from these elements. In second group Mg acts as a bridge element. The properties of bridge element are some what mixed of the elements of two subgroups as magnesium shows similarities with alkaline earth metals (IIA) on one hand and with zinc metals (IIB) on the other.



IUPAC NOMENCLATURE OF THE ELEMENT :

The names are derived by using roots for the three digits in the atomic number of the element and adding the ending –ium. The roots for the number are

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

Thus element with atomic number 109 will be named as **une** (**u** for 1, **n** for 0 and **e** for 9). Table summarises the names of the elements with atomic number above 100.

PART-B (PERIODIC PROPERTIES)

PERIODICITY :

(a) The regular gradation in properties from top to bottom in a group and from left to right in a period is called periodicity in properties.

(b) In a period, the ultimate shell remain same, but the number of electrons gradually increases.

(c) In a group, the number of electrons in the ultimate shell remains same, but the values of n increases.

Cause of periodicity

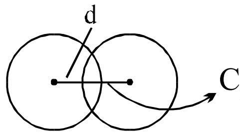
(a) The cause of periodicity in properties is due to the same outermost shell electronic configuration repeating over regular intervals.

(b) In the periodic table, elements with similar properties occur at intervals of 2, 8, 8, 18, 18 and 32. These numbers are called magic numbers.

ATOMIC RADIUS :

Since there is a problem in calculating actual size of atom, three types of radii can be defined :

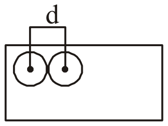
Type-I ***Covalent radius :***



$C_r = \frac{d}{2}$

[Used for H_2, Cl_2 and such molecules]

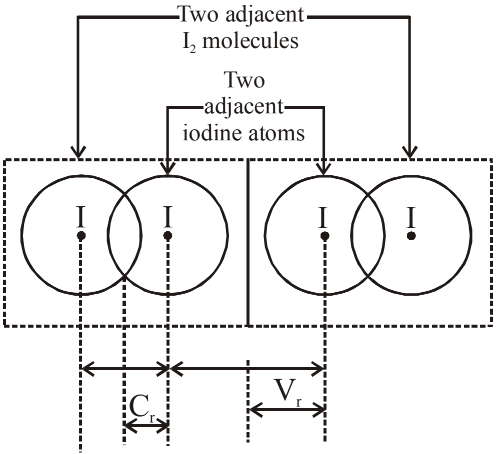
Type-II ***Metallic Radius :***



$M_r = \frac{d}{2}$

[Used for metals]

Type-III **VanderWaal's Radius or Collision radius**



VanderWaal's radius = $\frac{1}{2}$ × Internuclear distance between nuclei of two neighbouring atoms belonging to nearest molecules.

VanderWaal's radius > Metallic radius > Covalent radius

The VanderWaal's radius and covalent radius of chlorine atom are 1.80 Å and 0.99 Å respectively.

Ionic Radius

A neutral atom changes to a cation by the loss of one or more electrons and to an anion by the gain of one or more electrons. The magnitude of charge on cation and anion is equal to the number of electrons lost or gained respectively. The ionic radii of the ions present in an ionic crystal may be calculated from the internuclear distance between the two ions.

(a) ***Radius of Cation***

Radius of cation is smaller than that of corresponding atom. Since due to removal of electron(s), Z_{eff} increases.

(b) ***Radius of an Anion***

Radius of an anion is invariably larger than that of the corresponding atom, since due to addition of electron(s) Z_{eff} decreases.

Factors affecting atomic radius:

(a)	Z_{eff} increases, atomic radius decreases	$Li > Be > B > C > N > O > F$
(b)	Number of shell(n) increases, atomic radius increases	$Li < Na < K < Rb < Cs$
(c)	Screening effect increases, atomic radius increases.	
(d)	Magnitude of –ve charge increases, atomic radius increases	$O < O^{-} < O^{-2}$
(e)	Magnitude of +ve charge increases, atomic radius decreases	$Mn > Mn^{+2} > Mn^{+3} > Mn^{+4}$
(f)	Bond order increases, atomic radius decreases	$>N-N< \; > \; -N=N- \; > \; N\equiv N$

Periodic Trend :

(a) ***For normal elements:***

(i) ***Across a period :*** It decreases from left to right in a period as $Z_{eff.}$ increases.

Ex. $Ne > Li > Be > B > C > N > O > F$

(ii) ***In a group :*** It increases from top to bottom in a group as number of shells increases.

Ex. $Li < Na < K < Rb < Cs$

☞ ***Note :*** In III A group size of Al and Ga is nearly same (transition contraction)

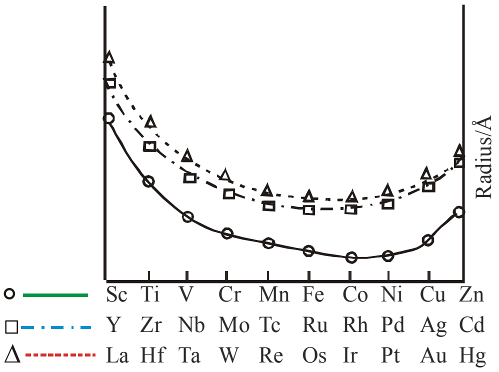
(b) ***For inert gases:***

In respective period generally, the atomic radius of inert gas is largest, because for inert gas VanderWaal's radius is defined. The VanderWaal's radius of inert gases also increases on moving from top to bottom in the group.

(c) ***For transition elements:***

From left to right in a period:

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



Trends in atomic radii of transition elements

In a group :

(i) The atomic radius of elements increases on moving down the first transition series (3d) to second transition series (4d). This is due to the increases in number of shells with the increase in atomic number.

(ii) The atomic radii of second (4d) and third (5d) transition series in a group is almost same except Y(39) and La(57)

(d) ***For inner transition elements:***

As we move along the lanthanide series, there is a decrease in atomic as well as ionic radius. The decrease in size is regular in ions but not so regular in atoms. This is called lanthanide contraction*.

Exceptions :

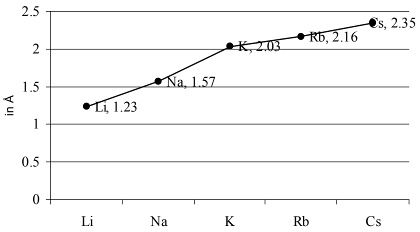
(1) Noble gases have largest atomic sizes [Vander waal radii]. However, their covalent radii are smaller *e.g.* Xe.

(2) Size of Ga and Al are same, [Z_{eff} increasing]

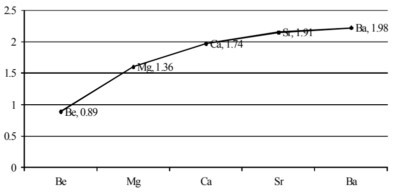
(3) Size of Hf & Zr are same (lanthanide contraction)

Graphical representation of atomic radius :

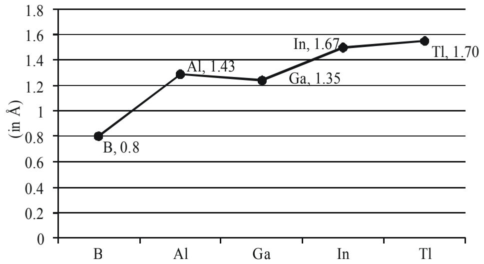
ALKALI METALS



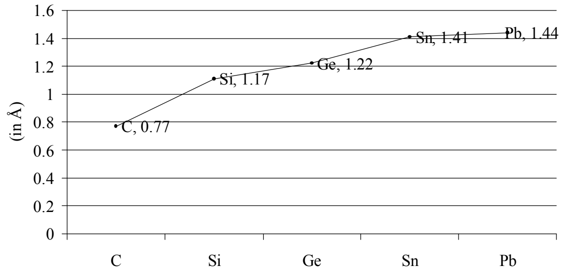
ALKALINE EARTH METALS



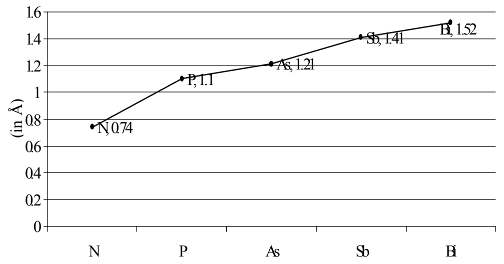
BORON FAMILY



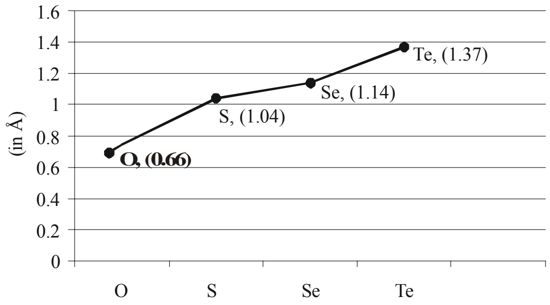
CARBON FAMILY



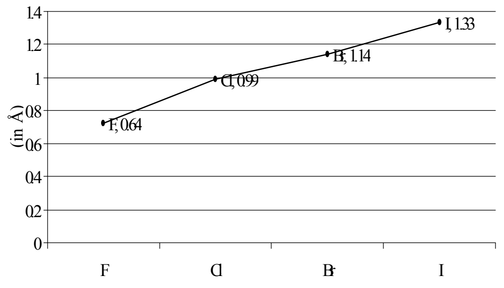
NITROGEN FAMILY (PNICOGENS)



CHALCOGENS



HALOGENS



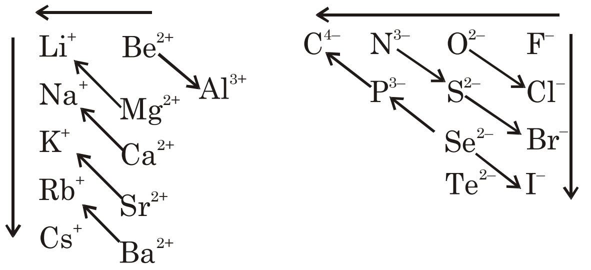
[☺ what can you predict or say about the increment in size along a group and decrement along a period]

ISOELECTRONIC MONOATOMIC SPECIES [Size ∝ 1/Z] :

(i) S^{2-}, Cl^{-}, K^{+}, Ca^{2+}, Sc^{3+} **(ii)** H^{-}, He, Li^{+}

(iii) O^{-2}, F^{-}, Na^{+}, Mg^{+2}, Al^{+3} **(iv)** $\frac{\text{radius of cation}}{\text{radius of anion}} = \frac{Z_{eff}\text{ of Anion}}{Z_{eff}\text{ of Cation}}$

Note:- In the direction of arrow (→) ionic size increases.



IONISATION ENERGY :

Amount of energy required to remove the most loosely bonded electron from an isolated gaseous atom from its ground state electronic configuration.

Units : kJ mol^{-1}, k cal mol^{-1}, eV per atom.

Ionisation is endothermic (endoergic) i.e. requires energy. Hence $\Delta H_{ionsation}$ is +ve.

$$M + \text{Energy } (IE_1) \longrightarrow M^+ + e^-$$

$$\left.\begin{array}{ll} M \longrightarrow M^+ + e^- & \Delta H = IE_1 \\ M^+ \longrightarrow M^{+2} + e^- & \Delta H = IE_2 \\ M^{+2} \longrightarrow M^{+3} + e^- & \Delta H = IE_3 \end{array}\right] \text{Successive ionisation energy}$$

$IE_3 > IE_2 > IE_1$ (always)

FACTORS AFFECTING IONISATION ENERGY :

(1) Atomic size : Varies inversely

(2) Screening effect : Varies inversely

(3) Nuclear charge (Z) : Varies directly

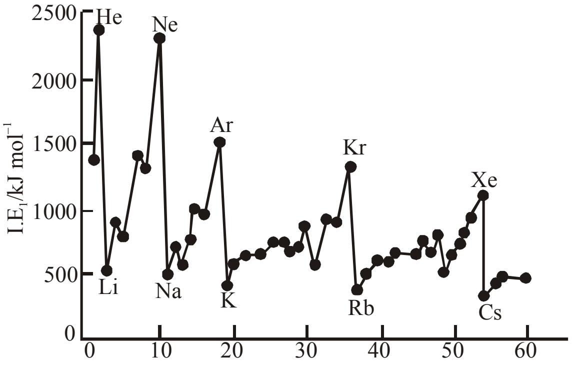
(4) Special electronic configuration of outermost electron (half filled / fully filled)

(5) Type of orbitals involved in Ionisation :s > p > d > f.

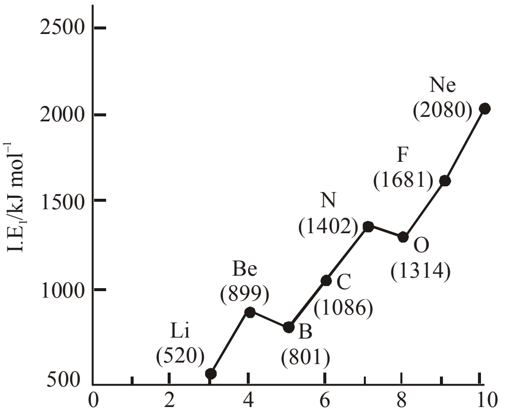
Note : Half filled and full filled inner orbitals, affects d-block and *f*-block trends.

General Trend: Along period I.E. increases [with some exception] [$Z_{eff}\uparrow$]

Along a group I.E. decrease [Z_{eff} constant, n ↑]



Variation of first ionization energy (I.E$_1$) with atomic number for elements with Z = 1 to 60



First ionization energy (I.E$_1$) of elements of the second period as a function of atomic number(Z)

Exception :

(1) Along a period, half filled and fully filled configurations have higher I.E. e.g. Be > B and N > O.

(2) Along a group, Ga ≃ Al

PROPERTIES AFFECTED BY IONISATION ENERGY :

(1) Metallic character (Varies inversely)

(2) Reducing power (Varies inversely)

(3) Tendency to stay in which state A^{+1}. A^{+2} or A^{+3}

☞ ***Note :***

(a) Helium (He) has the highest ionisation energy (IE_1) among all the elements and Caesium (Cs) has the least (IE_1) value.

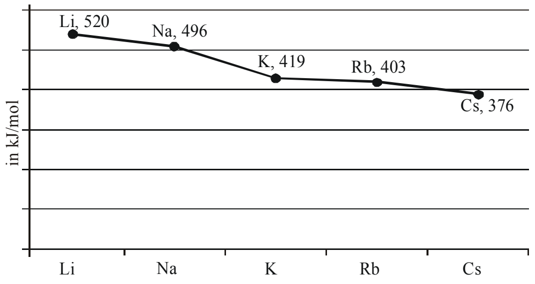
(b) Ionisation potential of inert gases is very high due to most stable s^2p^6 electronic configuration.

Element	He	Ne	Ar	Kr	Xe	Rn
IE_1(eV)	24.5	21.6	15.8	14.0	12.1	10.7

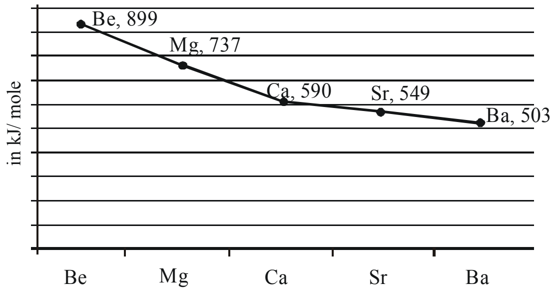
(c) For isoelectronic species I.E. increases with positive charge and decreases with negative charge. e.g. $Al^{+3} > Mg^{+2} > Na^{+} > F^{-} > O^{-2} > N^{-3}$

Graphical representation of ionisation energy :

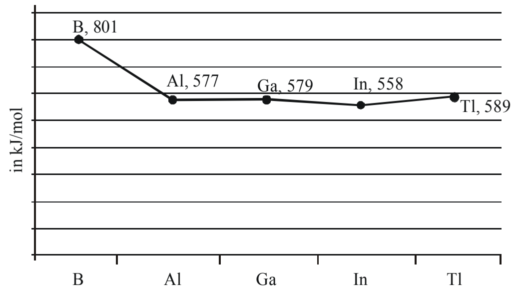
ALKALI METALS



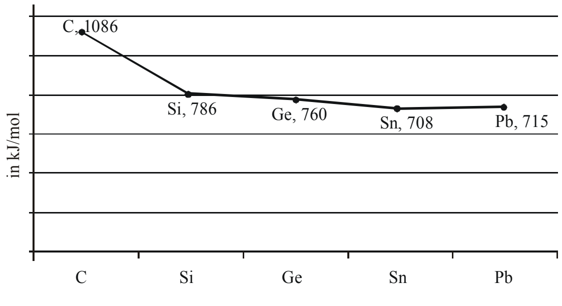
ALKALINE EARTH METALS



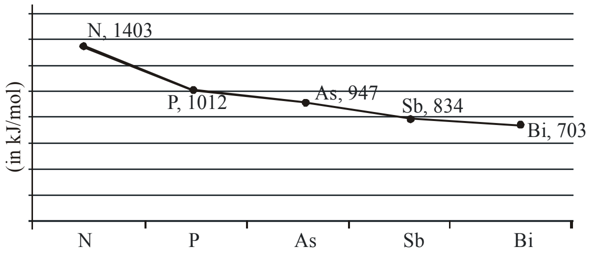
BORON FAMILY



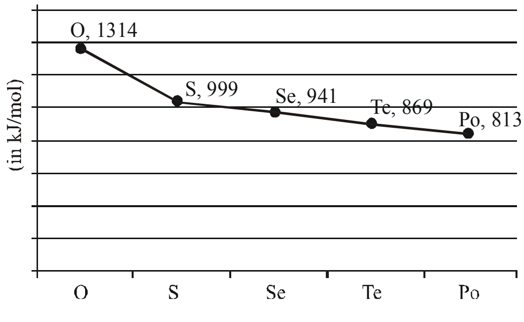
CARBON FAMILY

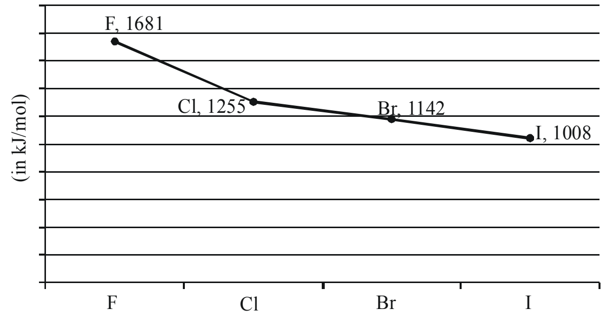


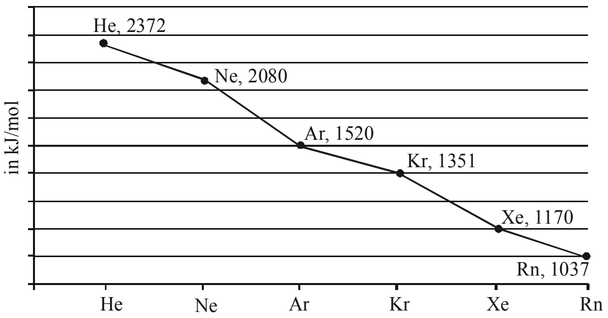
NITROGEN FAMILY (PNICOGENS)



CHALCOGENS




HALOGENS


NOBLE GASES

Ionisation energy of d-block elements :

1st, 2nd, 3rd IE's are increasing from left to right for 1st Transition series, but not regularly.

For 2nd IE Cr > Fe > Mn and Cu > Zn

For 3rd IE Mn > Cr > Fe and Zn has highest.

Element		Sc	Ti	v	Cr	Mn	Fe	Co	Ni	Cu	Zn
Atomic number		21	22	23	24	25	26	27	28	29	30
Electronic configuration											
	M	$3d^1 4s^2$	$3d^2 4s^2$	$3d^3 4s^2$	$3d^5 4s^1$	$3d^5 4s^2$	$3d^6 4s^2$	$3d^7 4s^2$	$3d^{10} 4s^2$	$3d^{10} 4s^1$	$3d^{10} 4s^2$
	M^{2+}	$3d^1$	$3d^2$	$3d^3$	$3d^4$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$3d^9$	$3d^{10}$
	M^{3+}	[Ar]	$3d^1$	$3d^2$	$3d^3$	$3d^4$	$3d^5$	$3d^6$	$3d^7$	–	–
Enthalpy of atomisation, $\Delta_a H^\ominus$/kJ mol^{-1}		326	473	515	397	281	416	425	430	339	126
Ionisation Enthalpy, $\Delta_i H^\ominus$/kJ mol^{-1}											
$\Delta_i H^\ominus$	I	631	656	650	653	717	762	758	736	745	906
	II	1235	1309	1414	1592	1509	1561	1644	1752	1958	1734
	III	2393	2657	2833	2990	3260	2962	3243	3402	3556	3829

ELECTRON AFFINITY & ELECTRON GAIN ENTHALPY (EGE) :

Electron Affinity : Amount of energy released when an electron is added to an isolated gaseous atom.

Units : k J mol^{-1} , k Cal mol^{-1} and eV per atom.

Addition of electron results in release of energy in most of the cases but for addition of second electron energy is always required. The sum of EA_1 & EA_2 is +ve (energy required)

$$E\,A \;\alpha\; \frac{1}{\text{atomic size}} \;\alpha\; Z_{eff}$$ (Cl has the highest E.A.)

Electron gain Enthalpy (EGE) : When expressed in terms of enthalpy change (ΔH_{eg}) then it is termed as EGE Remember that ΔH = – ve for exothermic change.

For EA_1, energy is released ∴ ΔH_{eg1} = – ve (Not always)

For EA_2, EA_3 ....energy is required ∴ ΔH_{eg2} = + ve (always)

In general $EA_1 + EA_2$, energy is required. ∴ $(\Delta H_{eg1} + \Delta H_{eg2}) > 0$

Note : $\Delta H_{e.g.} \simeq - EA$

FACTORS AFFECTING ELECTRON AFFINITY :

(1) ***Atomic size :*** Varies inversely

(2) ***Nuclear charge :*** Varies directly

(3) For stable electronic configuration i.e. half filled and fully filled shells EA decreases.

General Trend : Along a period, electron affinity increases [with a few exceptions] as $Z_{eff} \uparrow$. Along a group, electron affinity decreases after 3rd period. Between 2nd and 3rd period in p block electron affinity of 2nd period is lesser.

Exception :

(1) A fully filled and half filled configuration have low values of EA or even sometimes energy is required rather than getting released.

(2) 2nd period has lower value than 3rd owing to repulsion between electrons in small sized 2nd period elements.

Electron affinity of the main-group elements (in electron volts)

1							18
H							**He**
+0.754	2	13	14	15	16	17	–0.5
Li	**Be**	**B**	**C**	**N**	**O**	**F**	**Ne**
+0.618	≤ 0	$\leq +0.277$	+1.263	–0.07	+1.461	+3.399	–1.2
Na	Mg	Al	Si	P	S	Cl	Ar
+0.548	≤ 0	+0.441	+1.385	+0.747	+2.077	+3.617	–1.0
K	**Ca**	**Ga**	**Ge**	**As**	**Se**	**Br**	**Kr**
+0.502	-	+0.03	+1.2	+0.81	+2.021	+3.365	–1.0
Rb	**Sr**	**In**	**Sn**	**Sb**	**Te**	**I**	**Xe**
+0.486	-	+0.3	+1.2	+1.07	+1.971	+3.059	–0.8

ELECTRO NEGATIVITY :

Property of an atom in a molecule

F has highest electronegativity in the periodic table.

Decreasing order → $F > O > Cl \simeq N > Br > S \simeq C > I > H$.

Pauling Scale : $X_A \sim X_B = 0.208\sqrt{\Delta_{A-B}}$ Δ in kcal/mol

$X_A \sim X_B = 0.102\sqrt{\Delta_{A-B}}$ Δ in kJ/mol

$$\Delta_{A-B} = E_{A-B} - (E_{A-A} \times E_{B-B})^{1/2}$$

Where, E_{A-B} = Bond energy of A–B; E_{A-A} = Bond energy of A–A; E_{B-B} = Bond energy of B–B

Mulliken's Scale : $X_M = \frac{IP + EA}{2}$ (IP & EA are expressed in eV)

Mulliken's values of E N are about 2.8 times higher than values on Pauling's scale.

i.e. $X_P = \frac{X_M}{2.8}$

Allred–Rochow's Scale : $X_{AR} = \frac{0.359 Z_{eff}}{r^2}$; $X_P = X_{AR} + 0.744$

r = covalent radius of atom (in Å)

Z_{eff} = Effective nuclear charge on periphery

FACTORS AFFECTING ELECTRO NEGATIVITY :

(1) ***Nuclear attraction :*** Varies directly

(2) ***Atomic radius :*** Varies inversely

(3) ***Charge on ions :*** More positive charge, more electronegativity and more –ve charge, less electronegativity.

(4) ***Hybridisation :*** To be discussed later in chemical bonding chapter.

General Trends : Along a period, electronegativity increases

Along a group, electronegativity decreases

Exceptions : Not Noteworthy.

☞ **Note: Electronegativity of F > Cl but Electron affinity of Cl > F**

Electronegativity of Some Elements (on Paulling's Scale)

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9					

In **Pauling's Scale,** elements having almost same electronegativity are-

C ≃ S ≃ I ≃ 2.5
N = Cl = 3.0
P = H = 2.1
Cs= Fr = 0.7
Be= Al = 1.5

PROPERTIES DEPENDENT ON ELECTRO NEGATIVITY :

(1) % ionic character

(2) Strength of bond

(3) Bond Length

(4) Nature of hydrides

(5) Nature of hydroxide.

Graphical representation of Electronegativity :

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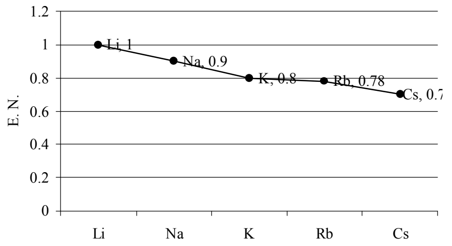
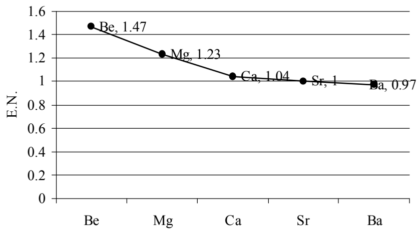
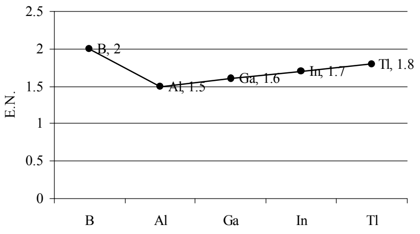


ALKALINE EARTH METALS



BORON FAMILY



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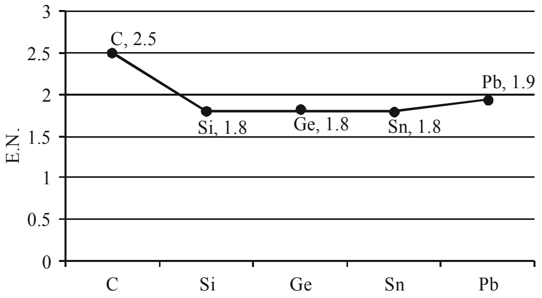

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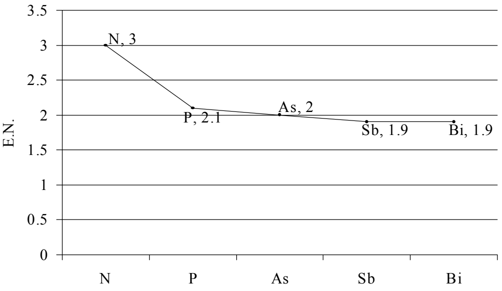
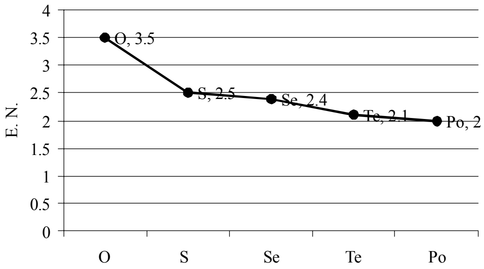


CHALCOGENS



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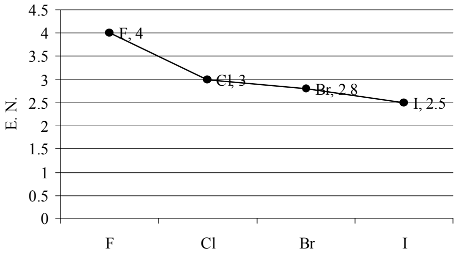


LATTICE ENERGY :

The amount of energy released during the formation of 1 mole crystal lattice from consituent gaseous ions.

$Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$ Lattice energy = -788 kJ mol^{-1}

OR

The lattice energy of an ionic compound is the energy required to separate 1 mole of solid ionic substance completely into gaseous ions.

$NaCl(s) \rightarrow Na^+(g) + Cl^-(g)$ Lattice energy = $+788$ kJ mol^{-1}

Lattice energies are large and positive because of attraction between positive and negative ions. The potential energy of two interacting charged particles is given by:

$E = K\frac{Q_1Q_2}{r}$ where, $r = r^+ + r^-$; $r^+ \rightarrow$ radius of cation, $r^- \rightarrow$ radius of anion

Where Q_1 and Q_2 are the charge on the particles in coulombs, and r is the distance between their centres in meters. The constant K has the value 9.0×10^9 J – m/C^2.

Factors affecting lattice energy :

(a) The lattice energy increases as the charge on the ions increases and as their radii decreases.

(b) The magnitude of lattice energies however depends primarily on the ionic charges because ionic radii do not vary over a wide range.

For Example :

The lattice energy order for some compounds is given below.

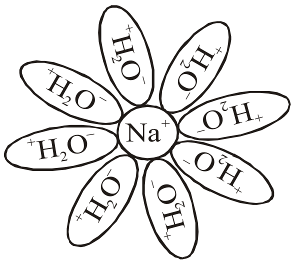
LiI < LiBr < LiCl < LiF (Due to small size of halide ion F^- alongwith small Li^+)

HYDRATION ENERGY :

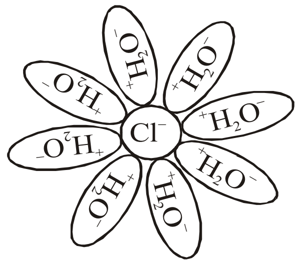
It is the energy released when 1 mol of gaseous ions are hydrated in water. It is directly proportional to nuclear charge and inversely proportional to size. It always decreases down the group.

(a) As the charge density of ion increases hydrated size (or aqueous radius) increases.

Size : $Li^+(aq) > Na^+(aq) > K^+(aq) > Rb^+(aq) > Cs^+(aq)$



Hydration of Na^+



Hydration of Cl^-

(b) As the hydrated size of ion increases ionic mobility decreases, which thus, decreases conductivity of ions.

Mobility : $Li^+(aq) < Na^+(aq) < K^+(aq) < Rb^+(aq) < Cs^+(aq)$

Conductivity : $Li^+(aq) < Na^+(aq) < K^+(aq) < Rb^+(aq) < Cs^+(aq)$

(c) Hydration energy also affects the solubility of ionic compounds. If hydration energy is greater than lattice energy then ionic compound will be soluble in water. More is the hydration energy, greater is the solubility, whereas, if lattice energy decreases, solubility of ionic compound increases.

MISCELLANEOUS CHEMICAL PROPERTIES :

1. ***Periodicity of hydra acids :***

(a) Acidic character of hydra acid increases from left to right in a period.

(b) Acidic character of hydra acid increases from top to bottom in a group.

2. ***Periodicity of oxy acids :***

(a) Acidic character of oxy acid increases from left to right in a period.

(b) Acidic character of oxy acid decreases from top to bottom in a group.

3. ***Periodicity of nature of oxide :***

(a) On moving from left to right in a period acidic nature of oxide generally increases.

e.g. $CO_2 < P_2O_5 < SO_3 < Cl_2O_7$

(b) On moving from top to bottom in a group acidic nature of oxide generally decreases.

4. ***General trends***

(a) ***Hydration energy*** decreases along a group.

(b) ***Lattice energy*** decreases along a group.

:: *Some points to Remember* ::

(without considering radioactive elements)

1. Second most electronegative element is Oxygen

2. Hydrogen is the lightest element and Lithium is lightest metal.

3. Helium has the highest value of I.P.

4. In periodic table metalloids are only in p-block.

5. Total gaseous elements are 11 (He, Ne, Ar, Kr, Xe, Rn, H_2, N_2, O_2, Cl_2, F_2)

6. Liquid metal at room temperature is – Hg.

7. Diamond is hardest natural substance.

8. Halogens have highest electron affinity and amongst them, Cl has the highest amongst them.

9. The largest cation of the periodic table = Cs^+.

10. The smallest cation of the periodic table = H^+.

11. The smallest anion of the periodic table = F^-.

12. The biggest element of periodic table = Fr.

13. The smallest element of periodic table = H.

14. Br is liquid non-metal element at room temperature.

15. Osmium is the heaviest element known.

16. Fluorine is the most electronegative element.

Extended or Long Form of the Periodic Table

☐ Metals

🟩 Non metals

■ Metalloids

s-Block Elements

d-Block Elements

p-Block Elements

Group → Period ↓	IA (1)	IIA (2)	IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIII (8)	VIII (9)	VIII (10)	IB (11)	IIB (12)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	0 (18)
1	1 **H** 1.0079 Hydrogen																	2 **He** 4.0026 Helium
2	3 **Li** 6.940 Lithium	4 **Be** 9.0122 Beryllium											5 **B** 10.811 Boron	6 **C** 12.011 Carbon	7 **N** 14.007 Nitrogen	8 **O** 15.999 Oxygen	9 **F** 18.998 Fluonne	10 **Ne** 20.180 Neon
3	11 **Na** 22.990 Sodium	12 **Mg** 24.305 Magnesium											13 **Al** 26.982 Aluminium	14 **Si** 28.086 Silicon	15 **P** 30.974 Phosporus	16 **S** 32.066 Sulphur	17 **Cl** 35.453 Chlorine	18 **Ar** 39.948 Argon
4	19 **K** 39.098 Potassium	20 **Ca** 40.078 Calcium	21 **Sc** 44.956 Scandium	22 **Ti** 47.867 Titanium	23 **V** 50.941 Vanadium	24 **Cr** 51.996 Chrominum	25 **Mn** 54.938 Manganese	26 **Fe** 55.847 Iron	27 **Co** 58.933 Cobolt	28 **Ni** 58.693 Nickel	29 **Cu** 63.546 Copper	30 **Zn** 65.39 Zinc	31 **Ga** 62.723 Gallium	32 **Ge** 72.61 Germanium	33 **As** 74.922 Arsenic	34 **Se** 78.96 Selenium	35 **Br** 79.904 Bromine	36 **Kr** 83.80 Krypton
5	37 **Rb** 85.468 Rubidium	38 **Sr** 87.62 Strontium	39 **Y** 88.906 Yttrium	40 **Zr** 91.224 Zirconium	41 **Nb** 92.906 Niobium	42 **Mo** 95.94 Molybdenum	43 **Tc** 98 Technetium	44 **Ru** 101.07 Ruthenium	45 **Rh** 102.91 Rhodium	46 **Pd** 106.42 Palladium	47 **Ag** 107.87 Silver	48 **Cd** 112.41 Cadmium	49 **In** 114.82 Indium	50 **Sn** 118.71 Tin	51 **Sb** 121.76 Antimony	52 **Te** 127.60 Tellurium	53 **I** 126.90 Iodine	54 **Xe** 131.29 Xenon
6	55 **Cs** 132.91 Cesium	56 **Ba** 137.33 Barium	57 **La*** 138.91 Lanthanum	72 **Hf** 178.49 Hafnium	73 **Ta** 180.95 Tantalum	74 **W** 183.84 Tungsten	75 **Re** 186.21 Rhenium	76 **Os** 190.23 Osmium	77 **Ir** 192.22 Iridium	78 **Pt** 195.08 Platinum	79 **Au** 196.97 Gold	80 **Hg** 200.59 Mercury	81 **Tl** 204.38 Thallium	82 **Pb** 207.2 Lead	83 **Bi** 208.98 Bismuth	84 **Po** 210 Polonium	85 **At** 210 Astatine	86 **Rn** 222 Radon
7	87 **Fr** 223 Francium	88 **Ra** 226 Radium	89 **Ac**** 227 Actinium	104 **Unq** 261 Unnilquadium	105 **Unp** 262 Unnilpentium	106 **Unh** 266 Unnilhexium	107 **Uns** 264 Unnilseptium	108 **Uno** 269 Unniloctium	109 **Une** 268 Unnilennium	110 **Uun** 269 Ununnilium	111 **Uuu** 272 Unununium	112 **Uub** 277 Ununbium						

The symbols for elements 104-109 used in this table are those proposed by the American Chemical Society and 110-112 proposed by IUPAC

f-Block Elements

6	***Lanthanide Series**				58 **Ce** 140.12 Cerium	59 **Pr** 140.91 Praseodymium	60 **Nd** 144.24 Neodymium	61 **Pm** 145 Promethium	62 **Sm** 150.36 Samarium	63 **Eu** 151.96 Europium	64 **Gd** 157.25 Gadolinium	65 **Tb** 158.93 Terbium	66 **Dy** 162.50 Dyspro-	67 **Ho** 164.93 Holmium	68 **Er** 167.26 Erbium	69 **Tm** 168.93 Thulium	70 **Yb** 173.07 Ytterbium	71 **Lu** 174.97 Lutetium
7	****Actinide Series**				90 **Th** 232.04 Thorium	91 **Pa** 231.04 Protactinium	92 **U** 238.03 Uranium	93 **Np** 237 Neptunium	94 **Pu** 244 Plutonium	95 **Am** 243 Americium	96 **Cm** 247 Curium	97 **Bk** 247 Berkelium	98 **Cf** 251 Californium	99 **Es** 252 Einsteinium	100 **Fm** 257 Fermium	101 **Md** 258 Mendelevium	102 **No** 259 Nobelium	103 **Lr** 262 Lawrencium

IUPAC designations of groups of elements are given in brackets

GENERAL TRENDS OF DIFFERENT PROPERTIES IN THE PERIOD AND GROUPS

	PERIODS →	GROUPS →
Atomic radius	Decreases	Increases
Ionisation potential	Increases	Decreases
Electronegativity	Increases	Decreases
Electron affinity	Increases	Decreases
Covalent character of halides	Increases	Decreases
Metallic character	Decreases	Increases
Oxidising nature	Increases	Decreases
Reducing nature	Decreases	Increases
Screening effect	Increases	Increases
Effective nuclear charge (Z_{eff})	Increases	Constant
Valency w.r.t Oxygen	Increases	Constant
Basic character of hydroxides	Decreases	(Metals)Increases
Basic character of hydrides	Decreases	(Non metal) Decreases (Metals) Increases
Basic character of Oxides	Decreases	Increases
Strength of oxy acids	Increases	Decreases
Thermal stability of sulphate	Decreases	Increases
Thermal stability of carbonates	Decrease	Increases
Thermal stability of nitrates	Decrease	Increases
Therma stability of hydroxide	Decrease	Increases
Density	First increases then decreases	Increases
Electro positivity	Decrease	Increases

EXERCISE # O-1

Periodic Table

1. Which is not anomalous pair of elements in the Mendeleev's periodic table:-
(A) Ar and K (B) Co and Ni (C) Te and I (D) Al and Si

2. Representative elements belong to :
(A) s-and p-block (B) d-block (C) d and f-block (D) f-block

3. True statement is :-
(A) All the transuranic elements are synthetic elements
(B) Elements of third period are called transition elements
(C) Element of [Ar] $3d^{10}4s^2$ configuration is placed in IIA group
(D) Electronic configuration of elements of a group is same

4. Which of the following match is correct :-
(A) Last natural element – Uub
(B) General electronic configuration of IA group –ns^2
(C) Inert gas elements lies between 2nd – 6th period
(D) Typical elements – 3rd period elements

5. The electronic configuration of elements X and Z are $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ and $1s^2\ 2s^2\ 2p^5$ respectively. What is the position of element X with respect to position of Z in the periodic table -
(A) Just below element Z (B) Just above Z
(C) Left to the Z (D) right to the Z

6. Which of the following is not a Dobereiner triad :
(A) H, F, Cl (B) N, O, F (C) P, As, Sb (D) S, Se, Te

7. Select the incorrect statement for Lother Meyer's curve :
(A) Curve is plotted between atomic weight and atomic volume
(B) Alkali metals occupy maxima of curve
(C) Halogens occupy descending portions of the curve
(D) Transition metals occupy bottom portions of the curve

Atomic & Ionic Radii

8. The size of the following species increases in the order:
(A) $Mg^{2+} < Na^+ < F^-$ (B) $F^- < Na^+ < Mg^{2+}$
(C) $Mg^{2+} < F^- < Na^+$ (D) $Na^+ < F^- < Mg^{2+}$

9. Highest size will be of
(A) Br^- (B) I (C) I^- (D) I^+

10. Element Cu has two oxidation states Cu^{+1} & Cu^{+2}. the right order of radii of these ions.
(A) $Cu^{+1} > Cu^{+2}$ (B) $Cu^{+2} > Cu^{+1}$ (C) $Cu^{+1} = Cu^{+2}$ (D) $Cu^{+2} \geq Cu^{+1}$

11. The correct order of increasing atomic size of element N,F, Si & P.
(A) N < F < Si < P (B) F > N < P < Si (C) F < N < P < Si (D) F < N < Si < P

12. The correct order of atomic or ionic size
(A) N < Li < B (B) Cl < Mg < Ca (C) $Ca^{+2} < S^{-2} < Cl^-$ (D) $Na^+ < Mg^{+2} < Cl^-$

13. In isoelectronic series largest difference between size is observed in N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+} :
(A) N^{3-}, Mg^{2+} (B) N^{3-}, O^{2-} (C) Mg^{2+}, Na^+ (D) F^-, Na^+

14. Mg, Mg^{2+}, Al and Al^{3+} are arranged in decreasing order of size 1 > 2 > 3 > 4. Species which are present at 1 and 4 position respectively are :

(A) Al, Mg^{2+} (B) Mg, Al^{3+} (C) Mg^{2+}, Al (D) Al^{3+}, Mg

Ionization Energy or Potential

15. In which of the following electronic configuration, ionisation energy will be maximum in

(A) [Ne] $3s^2\ 3p^1$ (B) [Ne] $3s^2\ 3p^2$ (C) [Ne] $3s^2\ 3p^3$ (D) [Ar] $3d^{10}\ 4s^2\ 4p^3$

16. The correct order of second ionisation potential of C, N, O and F is:

(A) C > N > O > F (B) O > N > F >C (C) O > F > N > C (D) F > O > N > C

17. The ionization energy will be maximum for which process?

(A) $Ba \rightarrow Ba^+$ (B) $Be \rightarrow Be^+$ (C) $Cs \rightarrow Cs^+$ (D) $Li \rightarrow Li^+$

18. Amongst the following, the incorrect statement is

(A) IE_1 (Al) < IE_1 (Mg) (B) IE_1 (Na) < IE_1(Mg)

(C) IE_2 (Mg) > IE_2 (Na) (D) IE_3 (Mg) > IE_3 (Al)

19. Decreasing ionization potential for K, Ca & Ba is

(A) Ba> K > Ca (B) Ca > Ba > K (C) K > Ba > Ca (D) K > Ca > Ba

20. Alkaline earth metals always form dipositive ions due to

(A) $IE_2 - IE_1 > 10$ eV (B) $IE_2 - IE_1 = 17$ eV

(C) $IE_2 - IE_1 < 10$ eV (D) None of these

21. The correct order of second I.P.

(A) Na < Mg > Al < Si (B) Na > Mg < Al > Si

(C) Na > Mg > Al < Si (D) Na > Mg > Al > Si

Electron affinity or Electron Gain Enthalpy

22. The process requires absorption of energy is

(A) $F \rightarrow F^-$ (B) $Cl \rightarrow Cl^-$ (C) $O^- \rightarrow O^{2-}$ (D) $H \rightarrow H^-$

23. Of the following elements, which possesses the highest electron affinity?

(A) As (B) O (C) S (D) Se

24. Electron affinities of O,F,S and Cl are in the order.

(A) O < S < Cl < F (B) O < S < F < Cl

(C) S < O < Cl < F (D) S < O < F < Cl

25. Increasing order of Electron affinity for following configuration.

(a) $1s^2, 2s^2\ 2p^2$ (b) $1s^2, 2s^2\ 2p^4$

(c) $1s^2, 2s^2\ 2p^6\ 3s^2\ 3p^4$ (d) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^3$

(A) d < a < b < c (B) d < a < c < b (C) a < b < c < d (D) a < b < d < c

26. Highest electron affinity is shown by

(A) F^- (B) Cl^- (C) Li^+ (D) Na^+

27. Which of the following statements is not true?

(A) F atom can hold additional electron more tightly than Cl atom

(B) Cl atom can hold additional electron more tightly than F atom

(C) The incoming electron encounters greater repulsion for F atom than for Cl atom

(D) It is easier to remove an electron from F^- than Cl^-.

Electronegativity

28. The outermost electronic configuration of most electronegative element amongst the following is :

(A) $ns^2\ np^3$ (B) $ns^2\ np^4$ (C) $ns^2\ np^5$ (D) $ns^2\ np^6$

29. In the following, which is the correct representation ?

(A) $\overset{\delta+}{C}-\overset{\delta-}{F}$ (B) $\overset{\delta+}{C}-\overset{\delta+}{Cl}$ (C) $\overset{\delta+}{F}-\overset{\delta-}{Cl}$ (D) $\overset{\delta-}{O}-\overset{\delta+}{F}$

30. On the Pauling's electronegativity scale, which element is next to F .

(A) Cl (B) O (C) Br (D) Ne

31. Which one is not correct order of electronegativity.

(A) $F > Cl > Br > I$ (B) $Si > Al > Mg > Na$

(C) $Cl > S > P > Si$ (D) None of these

32. The increasing order of acidic nature of Li_2O, BeO, B_2O_3

(A) $Li_2O > BeO < B_2O_3$ (B) $Li_2O < BeO < B_2O_3$

(C) $Li_2O < BeO > B_2O_3$ (D) $Li_2O > BeO > B_2O_3$

33. The lowest electronegativity of the element from the following atomic number is.

(A) 37 (B) 55 (C) 9 (D) 35

Miscellaneous

34. Which of the following does not reflect the periodicity of element

(A) Bonding behaviour (B) Electronegativity

(C) Ionisation potential (D) Neutron/ Proton ratio

35. Among the following, which species is/are paramagnetic ?

(i) Sr^{2+} (ii) Fe^{3+} (iii) Co^{2+} (iv) S^{2-} (v) Pb^{2+}

(A) i, iv, v (B) i, ii, iii (C) ii, iii (D) iv, v

36. Choose the s-block element from the following:

(A) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^5, 4s^1$ (B) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^{10}, 4s^1$

(C) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$ (D) all of the above

37. False statement for periodic classification of elements is

(A) The properties of the elements are periodic function of their atomic numbers.

(B) No. of non-metallic elements is less than the no. of metallic elements.

(C) First ionization energy of elements does not increase regularly with the increase in atomic number in a period.

(D) d-subshell is filled by final electron with increasing atomic number of inner transition elements.

38. Which of the following order is incorrect against the property indicated :

(A) $Mg < Ar < Na$ (2nd I.E.) (B) $Be < F < Cl$ ($|\Delta H_{eg}|$)

(C) $Rb < Na < K > Ca$ (atomic radius) (D) $P < S < N$ (electronegativity)

39. If each orbital can hold a maximum of three electrons, the number of elements in 9th period of periodic table (long form) will be

(A) 48 (B) 162 (C) 50 (D) 75

40. Which of the following element has highest metallic character .

	Element	IP
(A)	P	17 eV
(B)	Q	2 eV
(C)	R	10 eV
(D)	S	13 eV

41. The electronic configuration of an element is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$. The atomic number and the group number of the element 'X' which is just below the above element in the periodic table are respectively.

(A) 24 & 6 (B) 24 & 15 (C) 34 & 16 (D) 34 & 8

42. The number of d- electrons in Mn^{2+} is equal to that of

(A) p-electrons in N (B) s-electrons in Na

(C) d-electrons in Fe^{+2} (D) p-electrons in O^{-2}

43. Which of the following formula has involved all the energy terms used to calculated ΔH_f^o of $Na_2O_{(s)}$. (ΔH_{Sb} : Sublimation energy ; $I.E_1$: First ionisation energy ; $I.E_2$: Second ionisation energy ; B.D.E.; Bond dissosiation energy ; $E.G.E_1$: First electron gain enthalpy ; $E.G.E_2$: Second electron gain enthalpy; U : Lattice energy)

(A) $+ 2\Delta H_{Sb} + I.E_1 + I.E_2 + \frac{B.D.E.}{2} + E.G.E_1 + E.G.E_2 + U$

(B) $+ 2\Delta H_{Sb} + 2I.E_1 + \frac{B.D.E.}{2} + E.G.E_1 + E.G.E_2 + U$

(C) $+ 2\Delta H_{Sb} + 2I.E_1 + \frac{B.D.E.}{2} + 2E.G.E_1 + U$

(D) $+ 2\Delta H_{Sb} + I.E_1 + \frac{B.D.E.}{2} + E.G.E_1 + E.G.E_2 + U$

44. EN of the element (A) is E_1 and IP is E_2. Hence EA will be according to mulliken

(A) $2E_1 - E_2$ (B) $E_1 - E_2$ (C) $E_1 - 2E_2$ (D) $(E_1 + E_2)/2$

45. Moving from right to left in a periodic table, the atomic size is:

(A) Increased (B) Decreased (C) Remains constant (D) None of these

46. One element has atomic weight 39. Its electronic configuration is $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 4s^1$. The true statement for that element is:

(A) High value of IE (B) Transition element (C) Isotone with ${}_{18}Ar^{38}$ (D) None

47. The number of paired electrons in oxygen atom is:

(A) 6 (B) 16 (C) 8 (D) 32

48. The decreasing size of K^+, Ca^{2+}, Cl^- & S^{2-} follows the order:

(A) $K^+ > Ca^{+2} > S^{-2} > Cl^-$ (B) $K^+ > Ca^{+2} > Cl^- > S^{-2}$

(C) $Ca^{+2} > K^+ > Cl^- > S^{-2}$ (D) $S^{-2} > Cl^- > K^+ > Ca^{+2}$

49. Which of the following has the maximum number of unpaired electrons

(A) Mg^{2+} (B) Ti^{3+} (C) V^{3+} (D) Fe^{2+}

EXERCISE : O-2

Atomic & Ionic Radius

1. Select correct order of size :

(A) $Ti^{2+} < Ti < Zr$ (B) $Ti^{2+} < Ti < Hf$ (C) $Zr^{2+} < Zr \approx Hf$ (D) $Hf^{2+} < Hf \approx Zr$

2. Which of the following orders of atomic / Ionic radius is correct ?

(A) $B < Al \approx Ga$ (B) $Sc > Cu < Zn$ (C) $C < O < N$ (D) $Al^{+3} < Al^{+2} < Al^{+}$

Electron Affinity

3. Which of the following is correct order of EA.

(A) $N < C < O < F$ (B) $F > Cl > Br > I$ (C) $Cl > F > Br > I$ (D) $C < N < O < F$

4. The electron affinity of the members of oxygen family of the periodic table, follows the sequence

(A) $O > S > Se$ (B) $S > O < Se$ (C) $O < S > Se$ (D) $Se > O > S$

Ionisation Energy

5. Considering the following ionisation steps :

$A(g) \rightarrow A^{+}(g) + e^{-}$ $\Delta H = 100$ eV $A(g) \rightarrow A^{2+}(g) + e^{-}$ $\Delta H = 250$ eV

Select the correct statements :

(A) IE_1 of A(g) is 100 eV (B) IE_1 of $A^{+}(g)$ is 150 eV

(C) IE_2 of A(g) is 150 eV (D) IE_2 of A(g) is 250 eV

6. Which of the following are correct ?

(A) $IE_2(Mg) < IE_2(Na)$ (B) EA (N) < EA (P)

(C) Atomic size Mg^{+2} > Atomic size (Li^{+}) (D) IP of Na < Mg < Al

Electronegativty

7. Amongst the following statements, which is / are correct?

(A) Electronegativity of sulphur is greater than that of oxygen.

(B) Electron affinity of oxygen is smaller than that of sulphur.

(C) Electron gain enthalpy of fluorine is most negative

(D) Electron gain enthalpy of chlorine is most negative

8. An element 'E' have IE = x eV/atom and EA = y eV / atom and EN on Pauling scale is 1.2. Find EN of 'E' on Mullikan scale :

(A) $\frac{x+y}{2}$ (B) 1.2×2.8 (C) $\frac{x+y}{2.8} - 1.2$ (D) $1.2 - \frac{x+y}{2}$

Hydration Energy

9. Choose the INCORRECT order of hydrated size of the ions -

(A) $F^{\ominus}_{(aq.)} > Cl^{\ominus}_{(aq.)} > Br^{\ominus}_{(aq.)} > I^{\ominus}_{(aq.)}$ (B) $Rb^{\oplus}_{(aq.)} > K^{\oplus}_{(aq.)} > Na^{\oplus}_{(aq.)} > Li^{\oplus}_{(aq.)}$

(C) $Na^{\oplus}_{(aq.)} > Mg^{2+}_{(aq.)} > Al^{3+}_{(aq.)}$ (D) $Be^{2+}_{(aq.)} > Mg^{2+}_{(aq.)} > Ca^{2+}_{(aq.)} > Sr^{2+}_{(aq.)}$

10. Find the correct ionic mobility order in aqueous solution from the following options-

(A) $Li^{+} < Na^{+}$ (B) $Mg^{2+} < Sr^{2+}$ (C) $Na^{+} < K^{+}$ (D) $F^{-} < Cl^{-}$

Miscellaneous

11. Select the correct statement(s).

(A) The value of electron gain enthalpy of an element can be -ve or +ve.

(B) In the periodic table, metallic character of the elements increases down the group and decreases across the period

(C) The Cl^- & S^{2-} are isoelectronic species but first one is not smaller in size than the second

(D) Ionization enthalpy of an atom is equal to electron gain enthalpy of cation

12. In halogens, which of the following properties increase from iodine to fluroine

(A) Ionisation energy
(B) Electronegativity
(C) Bond length
(D) Electron affinity

13. In which of the following set of elements 1st element is more metallic than second.

(A) Ba, Ca
(B) Sb, Sn
(C) Ge, S
(D) Na, F

14. Which of the following order(s) is / are **CORRECT** :

(A) Li < Be < B < C (IE_1)

(B) HF < HCl < HBr < HI (Bond length)

(C) $Na_2O < MgO < Al_2O_3 < SiO_2 < P_2O_5$ (Acidic)

(D) $Li^+(g) < Na^+(g) < K^+(g) < Cs^+(g)$ (Ionic radius)

15. Which of the following order is correct :

(A) P < Si < Be < Mg < Na (Metallic character)

(B) $Mg^{+2} < Na^+ < F^- < O^{2-}$ (Ionic radius)

(C) Li < B < Be < C < N < O (2nd ionization energy)

(D) $Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$ (Ionic mobility)

16. The ionic compound $A^+ B^-$ is formed easily when the

(A) electron affinity of B is high

(B) ionization energy of A is low

(C) lattice energy of AB is high

(D) lattice energy of AB is low

17. Which of the following is/are correct?

(A) For $A(g) + e^- \longrightarrow A^-(g)$ ΔH may be negative

(B) For $A^-(g) + e^- \longrightarrow A^{2-}(g)$ ΔH may be negative

(C) For $A^-(g) + e^- \longrightarrow A^{2-}(g)$ ΔH must be positive

(D) For $Ne(g) + e^- \longrightarrow Ne^-(g)$ ΔH may be zero

EXERCISE # S-1

1. Find out the atomic number of element whose IUPAC name is Unnilpentium.

Fill your answer as sum of digits (excluding decimal places) till you get the single digit answer.

2. Select total number of acidic compounds out of given below.

$CsOH$, $OC(OH)_2$, $SO_2(OH)_2$, $Sr(OH)_2$, $Ca(OH)_2$, $Ba(OH)_2$, $BrOH$, $NaOH$, O_2NOH

3. Total number of enthalpy(s) (out of given eight) of A(g) which is/are **not** associated with conversion of

$A^{-}_{(g)} \longrightarrow A^{+4}_{(g)}$:

IE_1, IE_2, IE_3, IE_4, IE_5, IE_6, EGE_1, EGE_2

(IE = ionization energy, EGE = electron gain enthalpy)

4. Write the number of pairs in which size of first element or ion is higher as compared to IInd out of following eight pairs.

(O,S) , (He, Ne) , (Kr, Ne), (Na, Na^+), (Cl, Cl^-), (I^- , Cl^-), (Li, Na), (Li^+ , Na^+)

5. Total number of elements which have less IE_1 than that of 'N'.

Be, B, C, F, P, He

6. Size of H^- is smaller than how many elements among these ?

H, Li, Be, B, C, N, O, F, F^-

7. How many orders are **CORRECT**

(a) Acidic strength ⇒ ClOH < BrOH < IOH

(b) Basic strength ⇒ MgO < CaO < SrO < BeO

(c) Electronegativity ⇒ I < Br < N < O < F

(d) Electron affinity ⇒ Mg < Na < Si < S < Cl

(e) % Ionic character ⇒ NaF < KF < RbF

8. ___________ is higher for fluorine as compared to chlorine.

Find the number of properties given below to fill the blank space to make a correct statement.

Atomic mass, Covalent radius, Ionic radius (X^-), Ionization energy,Electron affinity, Electronegativity, Hydration energy of uninegative ion (X^-),

9. Upto argon find the number of elements which have lower IE_1 as compared to He.

10. The number of pairs, in which EA of the second element is more than that of the first element is :

[O, S], [C, N], [O, N], [N, P], [Cl^+, F^+], [K^+, Na^+]

11. For an element the successive ionisation energy values (in eV/atom), are given below.

14.534, 29.601, 47.448, 77.472, 97.888, 552.057, 667.029

Find the number of valence shell electrons in that element.

Subjective :

12. Calculate E.N. of chlorine atom on Pauling scale if I.E. of Cl^- is 4eV & of E.A. of Cl^+ is + 13.0 eV.

13. Increasing order of ionic size :

N^{3-}, Na^+, F^-, O^{2-}, Mg^{2+}

EXERCISE # S-2

Paragraph for Questions 1 to 2

First electron gain enthalpy (in $\frac{kJ}{mol}$) of few elements are given below :

Elements	ΔH_{eg}
I	–60
II	–45
III	–328
IV	–295
V	+ 48

Answer the following questions on the basis of above data:

1. Which element may be an inert gas

(A) I (B) III (C) IV (D) V

2. Which element is most non-metallic among all the elements -

(A) I (B) II (C) III (D) IV

Paragraph for Questions 3 to 4

The IE_1 and the IE_2 in KJ/mol of a few elements designated by U, V, W, X are shown below.

Atom	IE_1	IE_2
U	2464	6110
V	610	7542
W	928	1810
X	1588	3410

Based on the above information answer the follwoing question :-

3. Which of the elements represent a noble gas.

(A) U (B) V (C) W (D) X

4. Which of the following element belongs to group 1 (IA).

(A) U (B) V (C) W (D) X

Paragraph for Question 5 to 7

Nature of bond can be predicted on the basis of electronegativity of bonded atoms, greater difference in electronegativity (X), more will be the polarity of bond, and polar bond are easily broken in polar solvent like water. For hydroxy acids $X_O - X_A$ difference predict the nature of oxide formed by the element A.

$|X_O - X_A| > |X_O - X_H|$ then A–O–H show basic nature (NaOH)

$|X_O - X_A| < |X_O - X_H|$ then A–O–H show acidic nature (H–O–Cl)

With the help of EN values [$EN_A = 1.8$, $EN_B = 2.6$, $EN_C = 1.6$, $EN_D = 2.8$] answer the following questions for the compounds HAO, HBO, HCO, HDO.

5. Compounds whose aqueous solution is acidic and order of their acidic strength

(A) AOH, COH ; AOH < COH (B) HDO, HBO ; HDO > HBO

(C) AOH, COH ; AOH > COH (D) HDO, HBO ; HDO < HBO

6. Compounds whose aqueous solution is basic and order of their basic strength

(A) AOH, COH ; AOH < COH (B) HDO, HBO ; HDO > HBO

(C) AOH, COH ; AOH > COH (D) HDO, HBO ; HDO < HBO

7. Percentage ionic character of compound AB is

(A) 42.42% (B) 24.24% (C) 15.04% (D) None of these

Paragraph for Question 8 to 12

All the simple salt dissolve in water, producing ions and consequently the solution conduct electricity. In this process water molecule surround both the cations and anions & release energy. This process is called hydration & energy released is called hydration energy & it depends on size of gaseous ions.

Answer the following questions with respect to given cations.

Na^+, Mg^{2+}, Al^{3+}

8. Order of extent of hydration

(A) $Na^+ = Mg^{2+} = Al^{3+}$ (B) $Na^+ > Mg^{2+} > Al^{3+}$
(C) $Al^{+3} > Mg^{2+} > Na^+$ (D) $Al^{+3} > Mg^{2+} < Na^+$

9. Order of hydration energy

(A) $Na^+ = Mg^{2+} = Al^{3+}$ (B) $Na^+ > Mg^{2+} > Al^{3+}$
(C) $Al^{+3} > Mg^{2+} > Na^+$ (D) $Al^{+3} > Mg^{2+} < Na^+$

10. Order of size of hydrated ion.

(A) $Na^+ = Mg^{2+} = Al^{3+}$ (B) $Na^+ > Mg^{2+} > Al^{3+}$
(C) $Al^{+3} > Mg^{2+} > Na^+$ (D) $Al^{+3} > Mg^{2+} < Na^+$

11. Order of ionic mobility

(A) $Na^+ = Mg^{2+} = Al^{3+}$ (B) $Na^+ > Mg^{2+} > Al^{3+}$
(C) $Al^{+3} > Mg^{2+} > Na^+$ (D) $Al^{+3} > Mg^{2+} < Na^+$

12. Order of size of gaseous ions.

(A) $Na^+ = Mg^{2+} = Al^{3+}$ (B) $Na^+ > Mg^{2+} > Al^{3+}$
(C) $Al^{+3} > Mg^{2+} > Na^+$ (D) $Al^{+3} > Mg^{2+} < Na^+$

Paragraph for Question 13 to 14 are based on the following information.

Four elements P, Q, R & S have ground state electronic configuration as:

$P \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$ $Q \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
$R \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^3$ $S \rightarrow 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^1$

13. Comment which of the following option represent the correct order of true (T) & false (F) statement.

I size of P < size of Q **II** size of R < size of S
III size of P < size of R (appreciable difference) **IV** size of Q < size of S (appreciable difference)

(A) TTTT (B) TTTF (C) FFTT (D) TTFF

14. Order of IE_1 values among the following is

(A) P > R > S > Q (B) P < R < S < Q (C) R > S > P > Q (D) P > S > R > Q

Matching List

15.

	Column-I **Element**		**Column-II** **Period and group number respectively**
(P)	Si	(1)	4, 3
(Q)	Sc	(2)	3, 14
(R)	Ga	(3)	6, 13
(S)	Tℓ	(4)	4, 13

Code :

	P	**Q**	**R**	**S**
(A)	2	1	3	4
(B)	1	2	4	3
(C)	2	1	4	3
(D)	4	3	1	2

Match the column

16. Match the column :

Column I	**Column II**
(A) Highest density	(P) Lithium
(B) Metallic character	(Q) Osmium
(C) Lightest Metal	(R) Mercury
(D) Liquid at room temperature	(S) Bromine

17. If electrons are filled in the sub shells of an atom in the following order 1s, 2s, 2p, 3s, 3p, 3d, 4s, 4p, 4d, 4f......... then match the following element in List I with block in List II.

List-I	**List-II**
(A) K(19)	(P) s-Block
(B) Fe(26)	(Q) p-Block
(C) Ga(31)	(R) d-Block
(D) Sn(50)	(S) f-block

18. Match the characteristics mentioned in List II with the process in List I.

List I	**List II**
(A) $O\ (g) + e^- \rightarrow O^-\ (g)$	(P) Positve electron gain enthalpy
(B) $O^-\ (g) + e^- \rightarrow O^{2-}\ (g)$	(Q) Negative electron gain enthalpy
(C) $Na^-(g) \rightarrow Na(g) + e^-$	(R) Exothermic
(D) $Mg^+\ (g) + e^- \rightarrow Mg\ (g)$	(S) Endothermic

19. Match the column :

Column I	**Column II**
(A) Cl	(P) Metal
(B) F	(Q) Highest negative electron gain enthalpy
(C) Cu	(R) Most Electronegative element
(D) He	(S) Highest ionisation energy.

EXERCISE # JEE-MAIN

1. According to the Periodic law of elements, the variation in properties of elements is related to their :- **[AIEEE-2003]**

(1) Nuclear masses (2) Atomic numbers

(3) Nuclear neutron-proton number ratio (4) Atomic masses

2. 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 ionization enthalpy? **[AIEEE-2003]**

(1) Cr (2) Mn (3) Fe (4) V

3. Among Al_2O_3, SiO_2, P_2O_5 and SO_3, the correct order of acid strength is :- **[AIEEE-2004]**

(1) $Al_2O_3 < SiO_2 < SO_3 < P_2O_5$ (2) $SiO_2 < SO_3 < Al_2O_3 < P_2O_5$

(3) $SO_3 < P_2O_5 < SiO_2 < Al_2O_3$ (4) $Al_2O_3 < SiO_2 < P_2O_5 < SO_3$

4. The formation of the oxide ion $O^{2-}(g)$ requires first an exothermic and then an endothermic step as shown below :- **[AIEEE-2004]**

$O(g) + e^- = O^-(g), \quad \Delta H° = -142 \text{ kJ mol}^{-1}$

$O^-(g) + e^- = O^{2-}(g), \quad \Delta H° = 844 \text{ kJ mol}^{-1}$

This is because :-

(1) O^- ion will tend to resist the addition of another electron

(2) Oxygen has high electron affinity

(3) Oxygen is more electronegative

(4) O^- ion has comparitively larger size than oxygen atom

5. In which of the following arrangements, the order is NOT according to the property indicated below? **[AIEEE-2005]**

(1) $Al^{3+} < Mg^{2+} < Na^{1+} < F^-$ – increasing ionic size

(2) B < C < N < O - increasing first ionization enthalpy

(3) I < Br < F < Cl - increasing electron gain enthalpy (with negative sign)

(4) Li < Na < K < Rb - increasing metallic radius

6. Which of the following oxides is amphoteric in character ? **[AIEEE-2005]**

(1) SnO_2 (2) SiO_2 (3) CO_2 (4) CaO

7. Which of the following factors may be regarded as the main cause of lanthanide contraction ? **[AIEEE-2005]**

(1) poor shielding of one of 4f electron by another in the subshell

(2) effective shielding of one of 4f electrons by another in the subshell

(3) poorer shielding of 5d electrons by 4f electrons

(4) greater shielding of 5d electrons by 4f electrons

8. The increasing order of the first ionization enthalpies of the elements B, P, S and F (lowest first) is:- **[AIEEE-2006]**

(1) F < S < P < B (2) P < S < B < F (3) B < P < S < F (4) B < S < P < F

9. The set representing the correct order of ionic radius is :- **[AIEEE-2009]**

(1) $Li^+ > Na^+ > Mg^{2+} > Be^{2+}$ (2) $Mg^{2+} > Be^{2+} > Li^+ > Na^+$

(3) $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$ (4) $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$

10. 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 :- **[AIEEE-2011]**

(1) I > Br > Cl > F (2) F > Cl > Br > I (3) Cl > F > Br > I (4) Br > Cl > I > F

11. The increasing order of the ionic radii of the given isoelectronic species is :- **[AIEEE-2012]**

(1) K^+, S^{2-}, Ca^{2+} , Cl^- (2) Cl^-, Ca^{2+}, K^+, S^{2-}

(3) S^{2-}, Cl^-, Ca^{2+}, K^+ (4) Ca^{2+}, K^+, Cl^-, S^{2-}

12. In which of the following arrangements, the sequence is not strictly according to the property written against it ? **[JEE-MAIN 2012]**

(1) $CO_2 < SiO_2 < SnO_2 < PbO_2$: increasing oxidising power **(On-Line)**

(2) B < C < O < N : increasing first ionisation enthalpy

(3) $NH_3 < PH_3 < AsH_3 < SbH_3$: increasing basic strength

(4) HF < HCl < HBr < HI : increasing acid strength

13. Which of the following represents the correct order of increasing first ionization enthalpy for Ca, Ba, S, Se and Ar ? **[JEE-MAIN-2013]**

(1) Ca < S < Ba < Se < Ar (2) S < Se < Ca < Ba < Ar

(3) Ba < Ca < Se < S < Ar (4) Ca < Ba S < Se < Ar

14. The first ionisation potential of Na is 5.1 eV. The value of electron gain enthalpy of Na^+ will be :- **[JEE-MAIN-2013]**

(1) – 2.55 eV (2) – 5.1 eV (3) – 10.2 eV (4) + 2.55 eV

15. Electron gain enthalpy with negative sign of fluorine is less than that of chlorine due to : **[JEE-MAIN 2013 (On-Line)]**

(1) Smaller size of chlorine atom (2) Bigger size of 2p orbital of fluorine

(3) High ionization enthalpy of fluorine (4) Smaller size of fluorine atom

16. Given **[JEE-MAIN 2013 (On-Line)]**

Reaction	Energy Change (in kJ)
$Li(s) \longrightarrow Li(g)$	161
$Li(g) \longrightarrow Li^+(g)$	520
$\frac{1}{2}F_2(g) \longrightarrow F(g)$	77
$F(g) + e^- \longrightarrow F^-(g)$	(Electron gain enthalpy)
$Li^+(g) + F^-(g) \longrightarrow LiF(s)$	–1047
$Li(s) + \frac{1}{2}F_2(g) \longrightarrow Li\,F(s)$	–617

Based on data provided, the value of electron gain enthalpy of fluorine would be :

(1) –300 kJ mol^{-1} (2) –328 kJ mol^{-1} (3) –350 kJ mol^{-1} (4) –228 kJ mol^{-1}

17. The order of increasing sizes of atomic radii among the elements O, S, Se and As is :

[JEE-MAIN 2013 (On-Line)]

(1) As < S < O < Se (2) O < S < As < Se

(3) Se < S < As < O (4) O < S < Se < As

18. Which is the correct order of second ionization potential of C, N, O and F in the following ?

[JEE-MAIN 2013 (On-Line)]

(1) O > F > N > C (2) O > N > F > C

(3) C > N > O > F (4) F > O > N > C

19. Which of the following series correctly represents relations between the elements from X to Y ? **[JEE-MAIN 2014 (On-Line)]**

$X \longrightarrow Y$

(1) ${}_{18}Ar \rightarrow {}_{54}Xe$	Noble character increases
(2) ${}_{3}Li \rightarrow {}_{19}K$	Ionization enthalpy increases
(3) ${}_{6}C \rightarrow {}_{32}Ge$	Atomic radii increases
(4) ${}_{9}F \rightarrow {}_{35}Br$	Electron gain enthalpy with negative sign increases

20. The ionic radii (in Å) of N^{3-}, O^{2-} and F^{-} are respectively :- **[JEE-MAIN 2015 (Off-Line)]**

(1) 1.71, 1.40 and 1.36 (2) 1.71, 1.36 and 1.40

(3) 1.36, 1.40 and 1.71 (4) 1.36, 1.71 and 1.40

21. In the long form of the periodic table, the valence shell electronic configuration of $5s^2\ 5p^4$ corresponds to the element present in: **[JEE-MAIN 2015 (On-Line)]**

(1) Group 16 and period 5 (2) Group 17 and period 6

(3) Group 17 and period 5 (4) Group 16 and period 6

22. Which of the following atoms has the highest first ionization energy ?

[JEE-MAIN 2016 (Off-Line)]

(1) Sc (2) Rb (3) Na (4) K

23. The non-metal that does not exhibit positive oxidation state is :

[JEE-MAIN 2016 (On-Line)]

(1) Oxygen (2) Fluorine (3) Iodine (4) Chlorine

24. The electronic configuration with the highest ionization enthalpy is:- **[JEE-MAIN 2017]**

(1) $[Ar]\ 3d^{10}4s^2\ 4p^3$ (2) $[Ne]\ 3s^2\ 3p^1$

(3) $[Ne]\ 3s^2\ 3p^2$ (4) $[Ne]\ 3s^2\ 3p^3$

25. Consider the following ionization enthalpies of two elements 'A' and 'B' :

[JEE-MAIN 2017]

Element	**Ionization enthalpy (kJ/mol)**		
	1^{st}	**2^{nd}**	**3^{rd}**
A	**899**	**1757**	**14847**
B	**737**	**1450**	**7731**

Which of the following statements is correct ?

(1) Both A and B belong to group-2 where A comes below B

(2) Both A and B belong to group-1 where A comes below B

(3) Both A and B belong to group-1 where B comes below A

(4) Both A and B belong to group-2 where B comes below A

26. In the following reactions, ZnO is respectively acting as a/an : **[JEE-MAIN 2017]**

(a) $ZnO + Na_2O \rightarrow Na_2ZnO_2$

(b) $ZnO + CO_2 \rightarrow ZnCO_3$

(1) base and acid (2) base and base (3) acid and acid (4) acid and base

27. The group having isoelectronic species is :- **[JEE-MAIN 2017]**

(1) O^{2-} , F^{-}, Na^{+} , Mg^{2+} (2) O^{-} , F^{-}, Na , Mg^{+}

(3) O^{2-} , F^{-}, Na , Mg^{2+} (4) O^{-} , F^{-}, Na^{+} , Mg^{2+}

EXERCISE # JEE-ADVANCED

1. **Statement-1 :** F atom has a less negative electron gain enthalpy than Cl atom. **[JEE 2000]**

Statement-2 : Additional electron is repelled more efficiently by 3p electron in Cl atom than by 2p electron in F atom.

(A) Statement-1 is true, statement-2 is true and statement-2 is correct explanation for statement-1.

(B) Statement-1 is true, statement-2 is true and statement-2 is NOT the correct explanation for statement-1.

(C) Statement-1 is true, statement-2 is false.

(D) Statement-1 is false, statement-2 is true.

2. The correct order of radii is: **[JEE 2000]**

(A) N < Be < B (B) $F^- < O^{2-} < N^{3-}$ (C) Na < Li < K (D) $Fe^{3+} < Fe^{2+} < Fe^{4+}$

3. The IE_1 of Be is greater than that of B. **[T/F]** **[JEE 2001]**

4. The set representing correct order of IP_1 is **[JEE 2001]**

(A) K > Na > Li (B) Be > Mg > Ca (C) B > C > N (D) Fe > Si > C

5. Identify the least stable ion amongst the following: **[JEE 2002]**

(A) Li^- (B) Be^- (C) B^- (D) C^-

6. The increasing order of atomic radii of the following group 13 elements is

(A) Al < Ga < In < Tl (B) Ga < Al < In < Tl **[JEE 2016]**
(C) Al < In < Ga < Tl (D) Al < Ga < Tl < In

7. The option(s) with only amphoteric oxides is (are): **[JEE 2017]**

(A) Cr_2O_3, CrO, SnO, PbO

(B) NO, B_2O_3, PbO, SnO_2

(C) Cr_2O_3, BeO, SnO, SnO_2

(D) ZnO, Al_2O_3, PbO, PbO_2

ANSWERS KEY

EXERCISE # O-1

Que.	1	2	3	4	5	6	7	8	9	10
Ans.	D	A	A	D	A	B	C	A	C	A
Que.	11	12	13	14	15	16	17	18	19	20
Ans.	C	B	A	B	C	C	B	C	B	C
Que.	21	22	23	24	25	26	27	28	29	30
Ans.	B	C	C	B	A	C	A	C	A	B
Que.	31	32	33	34	35	36	37	38	39	40
Ans.	D	B	B	D	C	C	D	C	D	B
Que.	41	42	43	44	45	46	47	48	49	
Ans.	C	B	B	A	A	C	A	D	D	

EXERCISE # O-2

Que.	1	2	3	4	5	6	7	8	9	10
Ans.	A,B,C,D	A, B, D	A, C	B, C	A, B, C	A, B	B, D	A, B	B, C	A,B,C,D
Que.	11	12	13	14	15	16	17			
Ans.	A, B, D	A, B	A, C, D	B, C, D	A, B, D	A, B, C	A, C			

EXERCISE # S-1

Que.	1	2	3	4	5	6	7	8	9	10
Ans.	6	4	3	3	4	0	3	3	17	4
Que.	11	12	13							
Ans.	5	3.03 (Pauling)	$Mg^{2+} < Na^{+} < F^{-} < O^{2-} < N^{3-}$							

EXERCISE # S-2

Que.	1	2	3	4	5	6	7	8	9	10
Ans.	D	C	A	B	B	A	C	C	C	C
Que.	11	12	13	14	15	16				
Ans.	B	B	B	A	C	(A)→Q;(B)→P,Q,R ;(C)→P;(D)→R, S				
Que.	17					18				
Ans.	(A)→R ; (B)→R ; (C)→Q ; (D)→S					(A) → Q,R ; (B) → P,S ; (C) → S ; (D) → Q,R				
Que.	19									
Ans.	(A) → Q ; (B) → R ; (C) → P ; (D) → S									

EXERCISE # JEE-MAIN

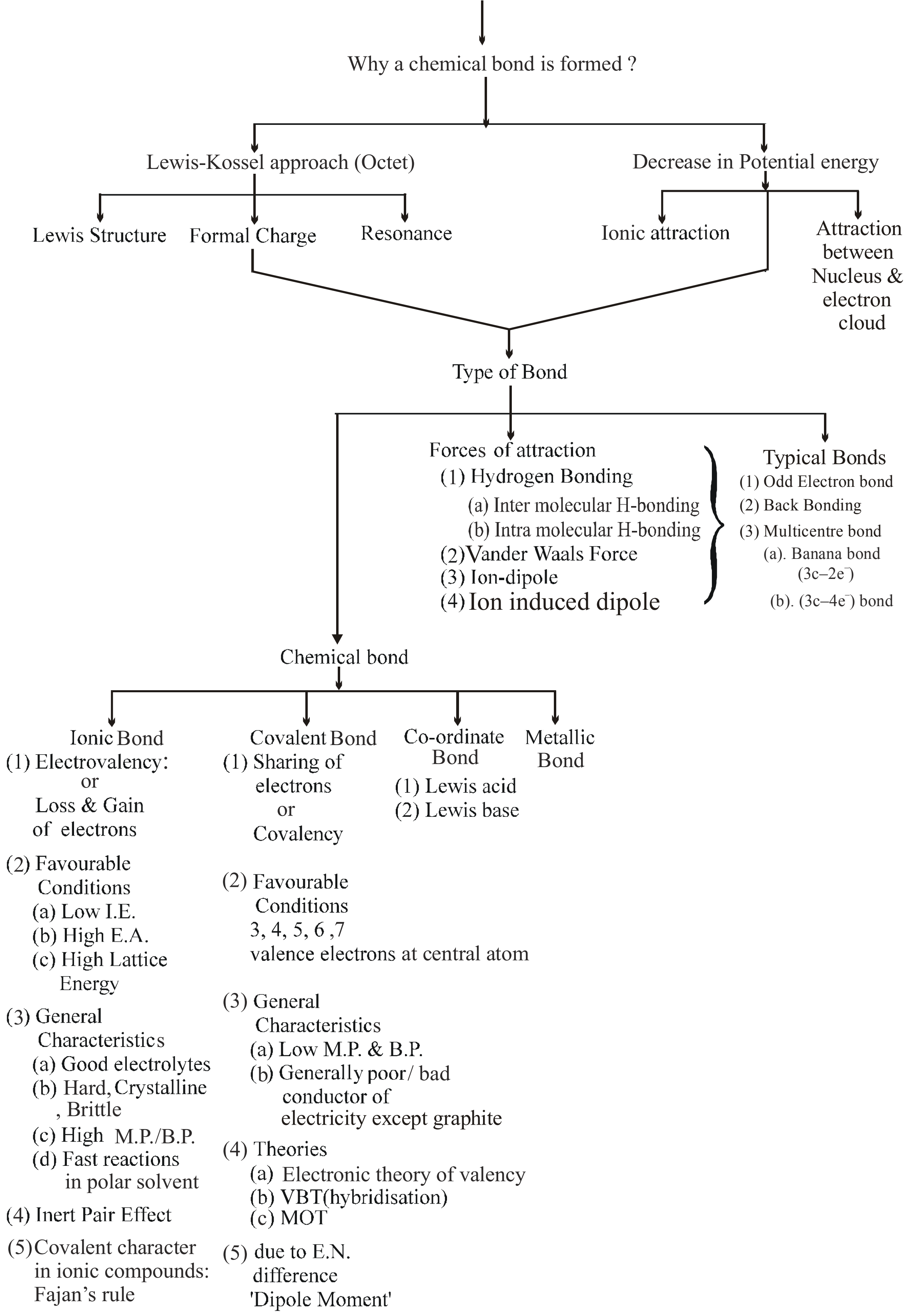
Que.	1	2	3	4	5	6	7	8	9	10
Ans.	2	1	4	1	2	1	3	4	4	3
Que.	11	12	13	14	15	16	17	18	19	20
Ans.	4	3	3	2	4	2	4	1	3	1
Que.	21	22	23	24	25	26	27			
Ans.	1	1	2	4	4	4	1			

EXERCISE # JEE-ADVANCED

Que.	1	2	3	4	5	6	7
Ans.	C	B	T	B	B	B	C, D

CHEMICAL BONDING

FLOWCHART TO CHEMICAL BONDING



KEY CONCEPT

CHEMICAL BOND

(a) The attractive force which holds various constituents (atoms, ions, etc.) together in different chemical species is called a chemical bond.

(b) It is the combination of two or more chemical species involving redistribution of e^- among them.

(c) This process is accompanied with decrease in energy.

(d) Decrease in energy strengthens the bond.

(e) Therefore, molecules are more stable than atoms.

CAUSE OF CHEMICAL COMBINATION

1. **Tendency to acquire minimum energy :**
2. **Tendency to acquire noble gas configuration (Octet rule) :**

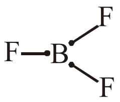
Atoms combine to complete an octet of electrons in their outer most shell. Hence all atoms have a tendency to acquire octet (ns^2np^6) configuration or to attain nearest noble gas configuration in their outermost orbit. This can be achieved by combining with other atom or ion.

LIMITIATION OF OCTET RULE

1. **Contraction of octet (incomplete octet)**

$\underline{Be}F_2$ $\underline{B}F_3$ $\underline{Al}Cl_3$ $\underline{B}Cl_3$

($4e^-$) ($6e^-$) ($6e^-$) ($6e^-$)

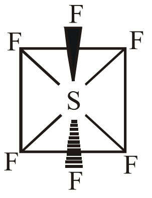


($6\ e^-$)

These compounds are hypovalent.

2. **Expansion of Octet (due to empty d-orbitals)**

PCl_5	SF_6	ClF_3	BrF_5	IF_7
↓	↓	↓	↓	↓
($10e^-$)	($12e^-$)	($10e^-$)	($12e^-$)	($14e^-$)



(12 e)

These compounds are hypervalent.

3. **Odd electron species**

Ex. NO, NO_2, ClO_2 etc.

OTHER EXCEPTIONS OF OCTET RULE

1. **Compounds of Noble gases**

Noble gases which have already completed their octets (or duplet in case of He.) should not form compounds. However, their compounds like XeF_2, XeF_6 & KrF_2 etc., have been actually prepared.

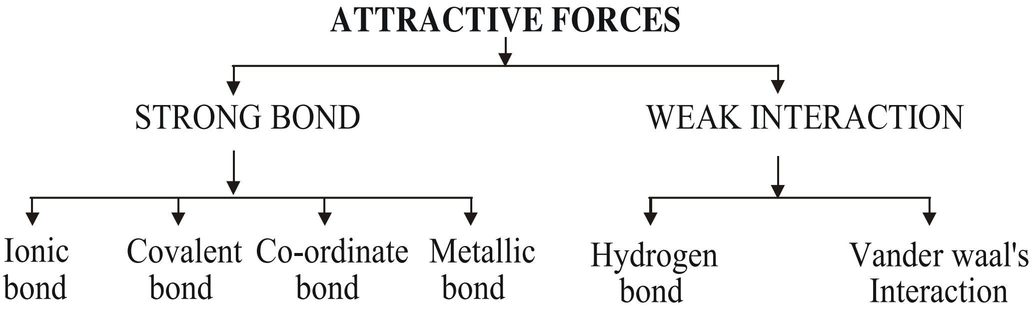
2. **Transition metal ions**

Cr^{3+}	Mn^{2+}	Fe^{2+}
$[Ar]3d^3$	$[Ar]3d^5$	$[Ar]3d^6$
[2, 8, 11]	[2, 8, 13]	[2, 8, 14]

3. **Pseudo inert gas configuration $[(n-1)d^{10}ns^2np^6]$**

Zn^{2+}	Cd^{2+}
$[Ar]3d^{10}$	$[Kr]4d^{10}$

CLASSIFICATION OF BONDS

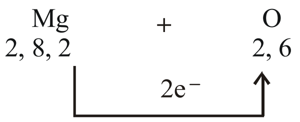


ELECTROVALENT OR IONIC BOND

(a) The chemical bond formed between two or more atoms as a result of the complete transfer of one or more electons from one atom to another is called as Ionic or electrovalent bond.

Example

(i)



$\longrightarrow Mg^{+2} + O^{-2} = MgO$

electrovalency of Mg = 2

electrovalency of O = 2

Note : Ionic bond is non-directional.

FACTORS FAVOURING IONIC BONDING

(a) Less Ionization energy of atom forming cation

(b) Higher electron affinity of atom forming anion

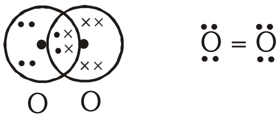
(c) Greater Lattice energy of formed product.

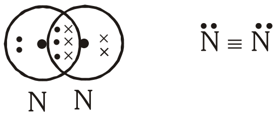
(d) Greater Electronegativity difference between atoms forming cation & anion.

COVALENT BOND

(a) A covalent bond is formed by the mutual sharing of electrons between two atoms.

H_2 molecule



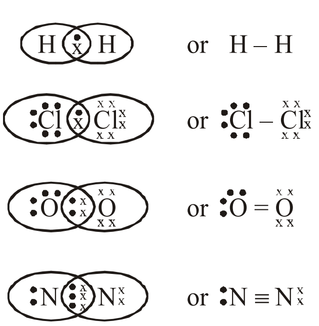


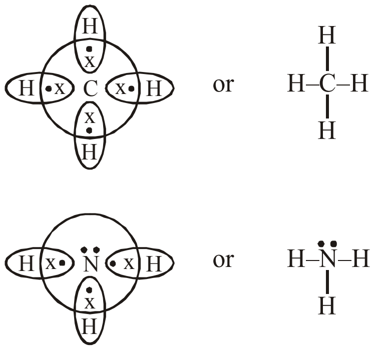
(b) The shared pair of electrons must have opposite spins, and are localised between two atoms concerned.

(c) On the basis of electrons being shared between two atoms the bonds are of three types –

Covalency : Capacity to form covalent bond is known as covalency

It is represented by (–) small line. – means single bond ; = means double bond ; ≡ means triple bond.



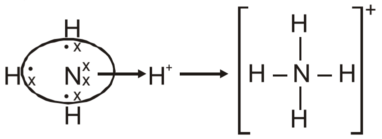


✦ **Covalent bonds are directional in nature**

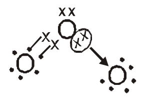
COORDINATE BOND (DATIVE BOND)

The bond formed between two atom in which contribution of an electron pair is made by one of them while the sharing is done by both.

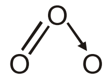
(i) NH_4^+ (ammonium ion)



(ii) O_3 (ozone)



or



Lewis Dot structures:

✦ Arrangement of various atoms in a molecule & types of bonding present in it but no idea of geometry of the molecule.

✦ In most cases we can construct a Lewis structure in three steps :

(1) Decide on the number of electrons that are included in the structure by adding together the number of all the valence electrons provided by the atoms.

(2) Write the chemical symbols of the atoms in the arrangement that shows which atoms are bonded together.

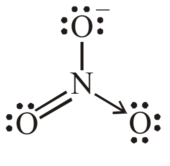
(3) Distribute the electrons in pairs so that there is one pair of electrons forming a single bond between each pair of atom bonded together and then supply electron pairs (to form lone pairs or multiple bonds) until each atom has an octet.

✦ In oxy acids all 'H' atoms are attached to oxygen as –OH groups except in H_3PO_3(dibasic), H_3PO_2(monobasic) & $H_4P_2O_5$ (dibasic). Which are directly bonded to central atom.

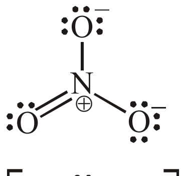
Applications :

✦ To know various linkages present

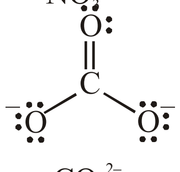
✦ To calculate formal charge of various elements.



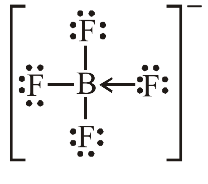
or



NO_3^-

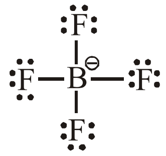


CO_3^{2-}



BF_4^-

or



FORMAL CHARGE

Lewis dot structures, in general, do not represent the actual shapes of the molecules. In case of polyatomic ions, the net charge is possessed by the ion as a whole and not by a particular atom. It is, however, feasible to assign a formal charge on each atom.

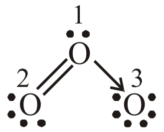
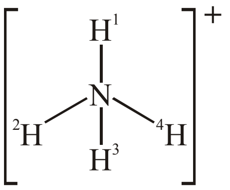
Formal Charge : $Q_F = N_A - N_{L.P.} - \frac{1}{2} N_{B.P.}$

Where :

N_A = Total number of valence electron in the free atom

$N_{L.P.}$ = Total number of non bonding (lone pair) electrons

$N_{B.P.}$ = Total number of bonding(shared) electrons

Molecule	**Structure**	**Formal Charge**
O_3		$O(1) = 6 - 2 - \frac{1}{2}(6) = +1$
		$O(2) = 6 - 4 - \frac{1}{2}(4) = 0$
		$O(3) = 6 - 6 - \frac{1}{2}(2) = -1$
CO		$C = 4 - 2 - \frac{1}{2} \times 6 = -1$
		$O = 6 - 2 - \frac{1}{2} \times 6 = +1$
NH_4^+		$N = 5 - 0 - \frac{1}{2}(8) = +1$
		On each H $= 1 - 0 - \frac{1}{2}(2) = 0$

RESONANCE*

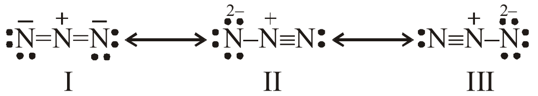
When a molecule cannot be completely represented by a single Lewis structure but it's characteristic properties can be described by two or more different structures, with similar energy position of nuclei, bonding and non-bonding pairs of electrons, then the true structure is said to be resonance hybrid of these structures. The phenomenon is called resonance and different contributing structures are called resonating structures or canonical structures.

- Resonance stabilizes the molecule as the energy of the resonance hybrid is less than the energy of any single canonical structure.
- Resonance averages the bond characteristics as a whole.
- The canonical forms have no real existence.

Resonance Structure of Some Molecules/Ions :

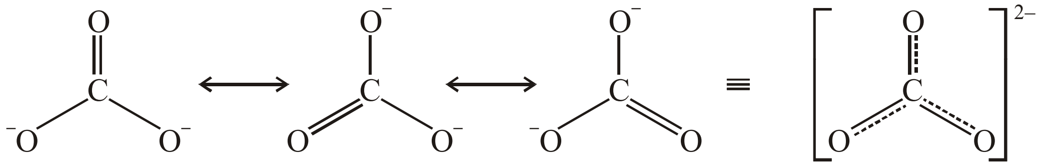
(i) Azide ion, N_3^- :

The azide ion may be represented as -



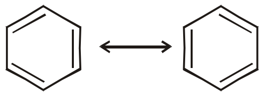
The structures II and III contribute equally and the molecule has almost double bond character in each N–N bond.

(ii)



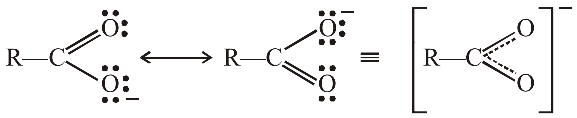
B.O. = 1.33

(iii)



BO = 1.5

(iv)



Bond order = $\frac{2+1}{2}=\frac{3}{2}=1.5$

*Descriptive discussion of concept of resonance will be done in **Organic Chemistry.**

VALENCE BOND THEORY

Valence bond theory was introduced by Heitler and London (1927) and developed further by Pauling and others. A discussion of the valence bond theory is based on the knowledge of atomic orbitals, electronic configurations of elements, the overlap criteria of atomic orbitals and the hybridization of atomic orbitals.

Formation of H_2 molecule :

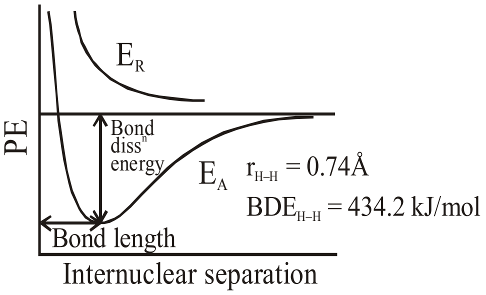
When two 'H'-atoms approaches towards each other for the formation of H_2 molecule. The following interactions takes place.

(i) e-e repulsion

(ii) e-p attraction

(iii) p-p repulsion

Experimentally, it has been found that the magnitude of attractive forces is more than the repulsive forces. As a result, when two atoms approach each other then potential energy decreases. Ultimately a stage is reached where the net force of attraction balances the force of repulsion and system acquires minimum energy.



The potential energy curve for the formation of H_2 molecule as a function of internuclear distance of the H atoms. The minimum in the curve corresponds to the most stable state of H_2.

At this stage two hydrogen atoms are said to be bonded together to form a stable molecule having the bond length of 74 pm. Since the energy gets released when the bond is formed between two hydrogen atoms, the hydrogen molecule is more stable than that of isolated hydrogen atoms.

Bond Length : Internuclear distance at minimum potential energy is called bond length.

Bond Dissociation Energy : The amount of energy released when one mole of same type of bonds are formed is called B.D.E.

434.2 kJ/mol of energy is required to dissociate one mole of H_2 molecule.

$H_2(g) + 434.2\ \text{kJ mol}^{-} \rightarrow H(g) + H(g)$

ORBITAL OVERLAP CONCEPT

In the formation of hydrogen molecule, there is a minimum energy state when two hydrogen atoms are so near that their atomic orbitals undergo partial interpenetration. This partial merging of atomic orbitals is called overlapping of atomic orbitals which results in the pairing of electrons. The extent of overlap decides the strength of a covalent bond. In general, greater the overlap the stronger is the bond formed between two atoms. Therefore, according to orbital overlap concept, the formation of a covalent bond between two atoms results by pairing of electrons, present in the valence shell having opposite spins.

Directional Properties of Bonds :

The valence bond theory explains the formation and directional properties of bonds in polyatomic molecules like CH_4, NH_3 and H_2O , etc. in terms of overlap and hybridisation of atomic orbitals.

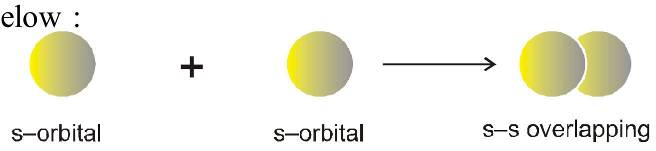
Types of Overlapping and Nature of Covalent Bonds

The covalent bond may be classified into following types depending upon the types of overlapping :-

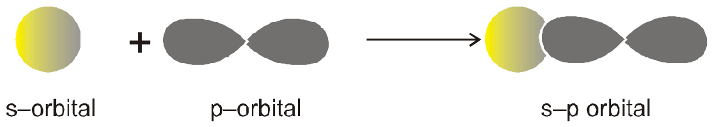
(i) sigma(σ) bond (ii) pi (π) bond (iii) delta(δ) bond

(i) **Sigma (σ) bond :** This type of covalent bond is formed by the end to end (head on or axial) overlap of bonding orbitals along the internuclear axis. This can be formed by any one of the following types of combinations of atomic orbitals.

- s-s overlapping : In this case, there is overlap of two half filled s-orbitals along the internuclear axis as shown below :



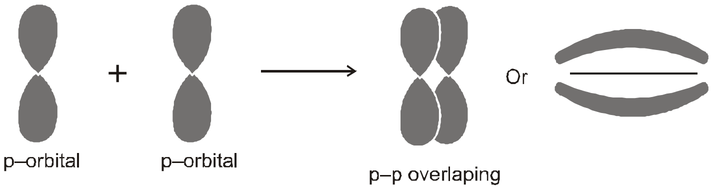
- s-p overlapping: This type of overlap occurs between half filled s-orbital of one atom and half filled p-orbital of another atom.



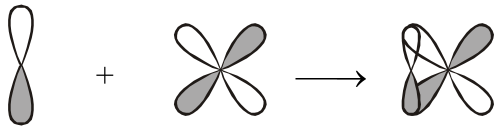
- p-p overlapping : This type of overlap takes place between half filled similar p-orbitals of the two approaching atoms.

(ii) **pi(π) bond :** In the formation of π bond the atomic orbitals overlap in such a way that their axis remain parallel to each other and perpendicular to the internuclear axis. The orbitals formed due to sidewise overlapping consists of two saucer type charged clouds above and below the plane of the participating atoms.

(a) **p_π–p_π**

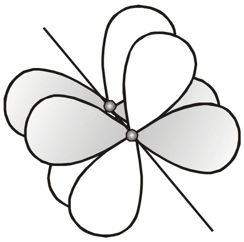


(b) **p_π–d_π**



p-orbital d-orbital p-d orbital overlaping

(iii) **delta (δ) bond :** are the covalent bonds where four lobes of d-orbital of one atom overlap with four lobes of the similar d-orbital of other atom. Except d_{z^2} all d orbitals form δ bond.



Strength of sigma and pi Bonds :

Basically the strength of a bond depends upon the extent of overlapping- In case of sigma bond, the overlapping of orbitals takes place to a larger extent. Hence, it is stronger as compared to the pi bond where the extent of overlapping occurs to a smaller extent. Further, it is important to note that pi bond between two atoms is formed in addition to a sigma bond. It is always present in the molecules containing multiple bond (double or triple bonds)

Advantages of VBT :

❖ It explain various bond characteristics e.g., bond length, bond strength.

❖ It explains the quantitative relationship between the extent of overlapping and bond dissociation energy.

❖ This theory accounts for shape and nature of bonding of the molecule whose covalency is not according to the number of half-filled orbitals present in the ground state.

❖ This theory redefined the stability of molecules e.g. BF_3, $AlCl_3$, PCl_5, SF_6 etc which are exception to octet rule.

Disadvantages of VBT :

❖ According to this theory three bond angle in CH_4 should be 90°, as these are formed by p–p overlapping, but actually it has four 109°28 angles. In NH_3 & H_2O, angle should be 90°. This is in disagreement with the actual bond angles of 107° & 104.5° in NH_3 & H_2O molecules respectively.

❖ In order to explain the characteristic geometrical shapes of polyatomic molecules like CH_4, NH_3, H_2O etc. Pauling introduced the concept of hybridisation.

HYBRIDISATION

The phenomenon of intermixing of the orbitals of slightly different energies so as to redistribute their energies, resulting in the formation of new set of orbitals of equivalent energies and shape. For example when one 2s and three 2p-orbitals of carbon hybridise, there is the formation of four new sp^3 hybrid orbitals.

Salient features of hybridisation: The main features of hybridisation are as under :

1. The number of hybrid orbitals is equal to the number of the atomic orbitals that get hybridised.
2. The hybridised orbitals are always equivalent in energy and shape.
3. The hybrid orbitals are more effective in forming stable bonds than the pure atomic orbitals.
4. These hybrid orbitals are directed in space in some preferred direction to have minimum repulsion between electron pairs and thus a stable arrangement. Therefore, the type of hybridisation indicates the geometry of the molecules.

Important conditions for hybridisation

(i) The orbitals present in the valence shell of the atom are hybridised.

(ii) The orbitals undergoing hybridisation should have almost equal energy.

(iii) Promotion of electron is not essential condition prior to hybridisation.

(iv) It is not necessary that only half filled orbitals participate in hybridisation. In some cases, even filled orbitals of valence shell take part in hybridisation.

Types of Hybridisation

There are various types of hybridisation involving s, p and d orbitals. The different types of hybridisation are as under :

(i) *sp hybridisation* (ii) *sp^2 hybridisation* (iii) *sp^3 hybridisation*

(iv) *sp^3d hybridisation:* (v) *sp^3d^2 hybridisation:* (vi) *sp^3d^3 hybridisation:*

Determination of hybridisation state –

To predict hybridisation following formula may be used :

Number of hybrid orbital = Steric Number (S.N.) = Number of σ-bond arround that atom + Number of lone pair on that atom.

Molecule	Method	Hybridisation
NH_4^+	S.N. = 4 + 0 = 4	sp^3 hybridisation.

Number of hybrid orbitals	Hybridisation
two	sp
three	sp^2
four	sp^3
five	sp^3d
six	sp^3d^2
seven	sp^3d^3

Hybridisation in Ionic solid species :

Species	Cationic part	Anionic part
PCl_5	PCl_4^+ (sp^3)	PCl_6^- (sp^3d^2)
PBr_5	PBr_4^+ (sp^3)	Br^-
XeF_6	XeF_5^+ (sp^3d^2)	F^-
N_2O_5	NO_2^+ (sp)	NO_3^- (sp^2)
I_2Cl_6 (liquid)	ICl_2^+ (sp^3)	ICl_4^-(sp^3d^2)
Cl_2O_6	ClO_2^+ (sp^2)	ClO_4^- (sp^3)

VALENCE SHELL ELECTRON PAIR REPULSION THEORY (VSEPR)

(a) Molecules having covalent bond have definite geometry as covalent bonding has directional characteristics. A simple theory was given for the molecular shape of the covalent molecules by Gillespie and Nyholm in 1957.

(b) This theory predicts the shape of the molecule by considering the most stable configuration of the bond angles in the molecule. This theory states

(i) Electron pairs in the valence shell of the central atom of a molecule, whether bonding or lone pairs are regarded as occupying localised orbitals. These orbitals arange themselves in so as to minimize the mutual electronic repulsions.

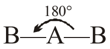
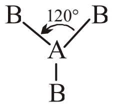
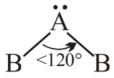
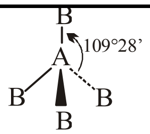
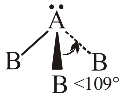
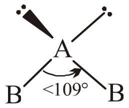
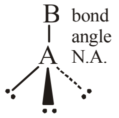
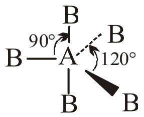
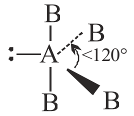
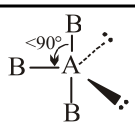
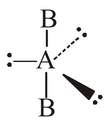
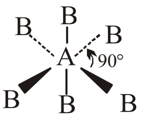
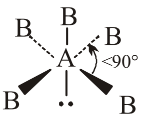
(ii) The magnitude of the different types of electronic repulsions follows the order given below:

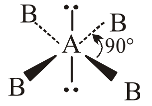
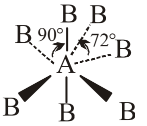
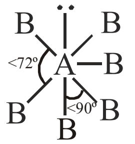
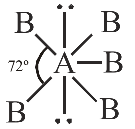
lone pair-lone pair > lone pair - bonded pair > bonded pair - bonded pair

(iii) The electronic repulsion between two pairs of electrons will be minimum if they are as far apart as possible.

(iv) The actual shape of the molecules containing lone pairs is a little distorted from the basic shape as in the NH_3 and H_2O molecules, the bond angles are not 109°28' but 107° and 104.5° respectively due to presence of one lone pair in NH_3 and two lone pairs in H_2O.

SHAPES OF MOLECULES BASED ON VSEPR THEORY

Total no. of hybrid orbitals	No. of b.p. (bond pairs)	No. of unshared pair i.e. l_p	General formula	Type of hybridisation	Stereo chemical formula/str.	Shape	Exam.
2	2	0	AB_2	sp		linear	$BeCl_2$
3	3	0	AB_3	sp^2		Trigonal planar	BCl_3, NO_3^- GaF_3, CO_3^{--}
3	2	1	AB_2	sp^2		V or Bent or angular	$SnCl_2$, O_3,SO_2
4	4	0	AB_4	sp^3		Tetrahedron	CH_4, SiF_4, $\overset{+}{N}H_4$
4	3	1	AB_3	sp^3		Trigonal pyramidal	NH_3, CH_3^-
4	2	2	AB_2	sp^3		V or Bent or angular	H_2O, SF_2
4	1	3	AB	sp^3		linear	ClO^-
5	5	0	AB_5	sp^3d		Trigonal bipyramidal	PF_5, SF_5^+, $SbBr_5$, XeO_3F_2
5	4	1	AB_4	sp^3d		Seesaw	SF_4
5	3	2	AB_3	sp^3d		T-shaped	ClF_3, BrF_3
5	2	3	AB_2	sp^3d		Linear	ICl_2^-, XeF_2, I_3^-
6	6	0	AB_6	sp^3d^2	 All bond ∠ = 90°	Octahedral or Square bipyramidal	SF_6, IF_6^+
6	5	1	AB_5	sp^3d^2	 All less than 90°	Square pyramidal	IF_5, $XeOF_4$, BrF_5

6	4	2	AB_4	sp^3d^2		Square planar	IF_4^-, XeF_4, ICl_4^-
7	7	0	AB_7	sp^3d^3		Pentagonal bipyramidal	IF_7
7	6	1	AB_6	sp^3d^3		Distorted octahedral	XeF_6, IF_6^-
7	5	2	AB_5	sp^3d^3		Pentagonal planar	XeF_5^-

BOND PARAMETERS

(I) Bond order ***(II) Bond Length (Bond distance)***

(III) Bond Angle ***(IV) Bond Enthalpy***

(I) Bond order *:*

The Bond Order is given by the number of bonds between the two atoms in a molecule.

The bond order, for example in H_2 (with a single shared electron pair), in O_2 (with two shared electron pairs) and in N_2 (with three shared electron pairs) is 1,2,3 respectively.

Note : Isoelectronic molecules and ions have identical bond orders; for example, F_2 and O_2^{2-} have bond order 1. N_2, CO and NO^+ have bond order 3.

Note : A general correlation useful for understanding the stablities of molecules is that: with increase in bond order, bond enthalpy increases and bond length decreases.

(II) Bond Length :-

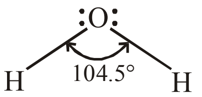
Bond length is defined as the equilibrium distance between the nuclei of two bonded atoms in a molecule. Bond lengths are measured by spectroscopic, X-ray diffraction and electron-diffraction techniques about which you will learn in higher classes.

Factors affecting bond length

(a) Size of atoms
(b) Effect of bond order or number of bonds
(c) Effect of Resonance
(d) Effect of Electronegativity difference
(f) Effect of Hybridisation

(III) Bond angle :

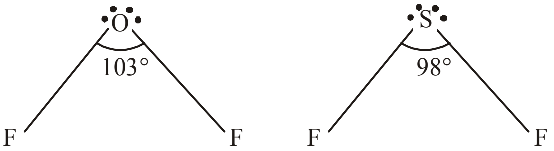
It is defined as the angle between the orbitals containing bonding electron pairs around the central atom in a molecule/complex ion. Bond angle is expressed in degree which can be experimentally determined by spectroscopic methods. For example $H-\hat{O}-H$ bond angle in water can be represented as under :



Comparison of bond angles.

(a) If central atoms are in different hybridisations then it can be compared.

(b) If same hybridisation but different central atoms then bond angle would be more of the molecule in which central atom is more E.N. eg. OF_2 , SF_2



(c) Molecules having same C.A. but different substituent then bond angle increases as the size of attached atom increases except in symmetrical molecules and H_2O / F_2O , NH_3 / NF_3

Factors affecting bond angle

(a) State of Hybridisation

(b) Presence of lone pair

(c) Electronegativity of central atom

(d) Electronegativity of surrounding atom

(e) Size of surrounding atom

(f) Lone pairs may sometimes be transferred from a filled shell of one atom to an unfilled shell of another bonded atom, causing less repulsion.

(g) Multiple bond orbital repel other orbitals more strongly than single bond orbitals.

(IV) Bond Enthalpy

It is defined as the amount of energy required to break one mole of bonds of a particular type between two atoms in a gaseous state. The unit of bond enthalpy is kJ mol^{-1}. For example, the H – H bond enthalpy in hydrogen molecule is 435.8 kJ mol^{-1}.

$H_2(g) \longrightarrow H(g) + H(g); \Delta_a H^{\ominus} = 435.8 \text{ kJ mol}^{-1}$

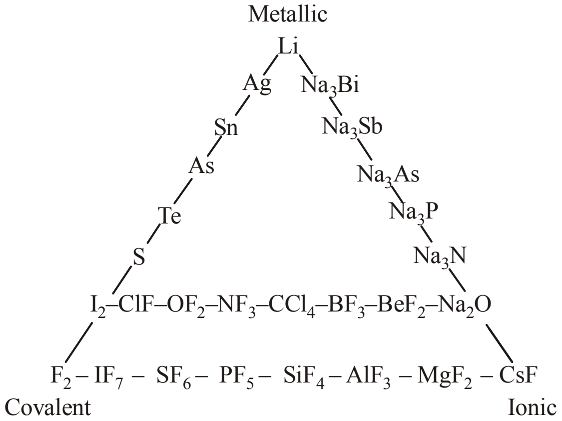
⇒ In polyatomic molecules the **term mean or average bond enthalpy** is used. It is obtained by dividing total bond dissociation enthalpy by the number of bonds broken as explained below in case of water molecule,

$H_2O(g) \longrightarrow H(g) + OH(g); \Delta_a H_1^{\ominus} = 502 \text{ kJ mol}^{-1}$

$OH(g) \longrightarrow H(g) + O(g); \Delta_a H_2^{\ominus} = 427 \text{ kJ mol}^{-1}$

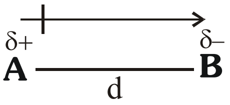
$$\text{Average bond enthalpy} = \frac{502+427}{2} = 464.5 \text{ kJ mol}^{-1}$$

SUMMARY OF THE THREE MAIN TYPES OF BONDS



DIPOLE MOMENT

- The degree of polarity of covalent bond is given by the dipole moment (μ)
- It is the product of the either charge and the distance between them.



Electronegativity of A < Electronegativity of B

$$\mu = \delta \times d$$

- Dipole moment is a vector quantity.
  The direction of dipole moment is represented by $+\!\!\!\longrightarrow$
- Units = Cm (S.I.) or esu cm(CGS) or Debye (common unit)
- 1 D = 10^{-18} esu cm = 3.33×10^{-30} coulomb m

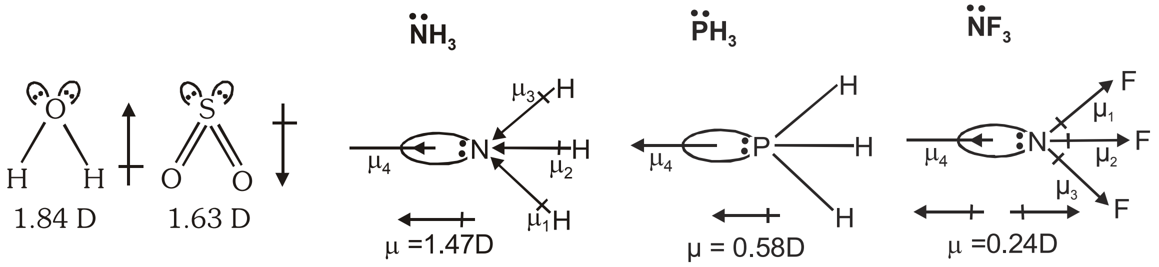
Dipole moment depends on

- Electronegativity difference between bonded atoms
- Direction of bond dipole moment
- Angle between various bonds
- Influence of unshared e^- pairs
- Magnitude of polarity of the molecule
- Symmetrical / Unsymmetrical shape.

Application of dipole moment

- To determine the polarity and geometry of molecules

Example :



- To the calculate the percentage of ionic character

$$\%\ \text{ionic character} = \frac{\text{observed } \mu}{\text{calculated } \mu \text{ (for 100 \% ionic)}} \times 100\ \%$$

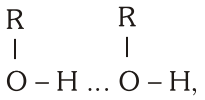
HYDROGEN BONDING

✦ **Hydrogen bonding:** When a hydrogen atom is bonded to a highly electronegative atom (like F, O or N) comes under the influence of another strongly electronegative atom, then a weak electrostatic force of attraction is developed between them, which is called as hydrogen bond.

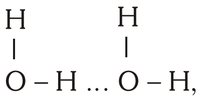
Types of H-bonding:

✦ Intermolecular H-bond

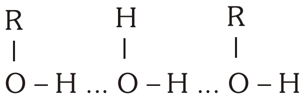
(i) This type of H-bonding takes place between two moleules. **Ex.** ROH, H_2O, R - OH & H_2O



$\overset{\overset{R}{|}}{O}-H\ldots\overset{\overset{R}{|}}{O}-H,$



$\overset{\overset{H}{|}}{O}-H\ldots\overset{\overset{H}{|}}{O}-H,$



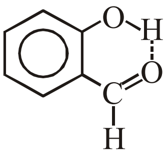
$\overset{\overset{R}{|}}{O}-H\ldots\overset{\overset{H}{|}}{O}-H\ldots\overset{\overset{R}{|}}{O}-H$

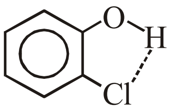
(ii) In such compounds molecular wt., M.P, & B.P. are high.

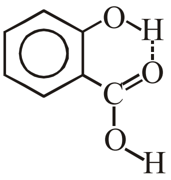
(iii) Extent of Inter molecular H-bonding ↑ viscosity & density ↑.

✦ Intramolecular H-bond

Normally when 2 hydroxyl groups are present on the same carbon atom i.e. gem diols are unstable, but "chloralhydrate" is a stable molecule due to formation of H-bond.







Properties influenced by hydrogen bonding

(a) Abnormal behaviour of water.

(b) Association of molecules eg. dimersiation of CH_3COOH, HCOOH

(c) Dissociation of a polar species.

(d) Abnormal melting point & boiling point.

(e) Enhanced solubility in water.

✦ Metallic bonds :

Electron gas model or sea model, with metal atom existing as kernels along with less firmly held valence electrons & bonds between various kernels (at the lattice site) & valence electrons are known as metallic bonds.

ALLOTROPES OF CARBON FAMILY

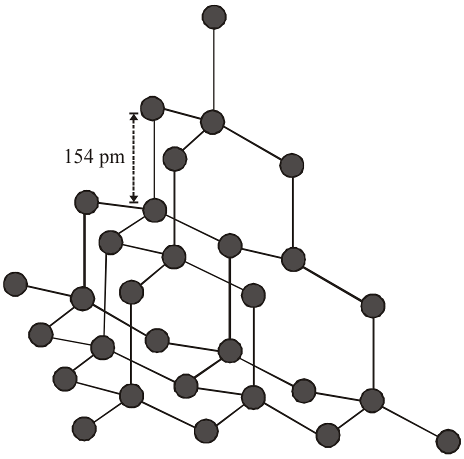
Carbon exhibits many allotropic forms; both crystalline as well as amorphous. Diamond and graphite are two well-known crystalline forms of carbon. In 1985, third form of carbon known as **fullerenes** was discovered by H.W.Kroto, E.Smalley and R.F.Curl. For this discovery they were awarded the Nobel Prize in 1996.

Si, Ge and Sn have a diamond type of structure, though Sn exists as a metallic form. Pb exists only in the metallic form. Ge is unusual because the liquid expands when it forms the solid. This property is unique to Ga, Ge and Bi.

$$\underset{\text{(Diamond structure)}}{\underset{\text{grey tin}}{\alpha - \text{Sn}}} \underset{}{\overset{13.2^\circ\text{C}}{\rightleftharpoons}} \underset{\text{(metallic)}}{\underset{\text{white tin}}{\beta - \text{Sn}}}$$

Diamond *(kinetically most stable allotrope of carbon, meta stable phase of carbon)*

It has a crystalline lattice. In diamond each carbon atom undergoes sp^3 hybridisation and linked to four other carbon atoms by using hybridised orbitals in tetrahedral fashion. The C–C bond length is 154 pm. The structure extends in space and produces a rigid three dimensional network of carbon atoms.



The structure of diamond

In this structure, directional covalent bonds are present throughout the lattice. It is very difficult to break extended covalent bonding and, therefore, diamond is a very hard substance. It is used as an abrasive for sharpening hard tools, in making dyes and in the manufacture of tungsten filaments for electric light bulbs.

Problem

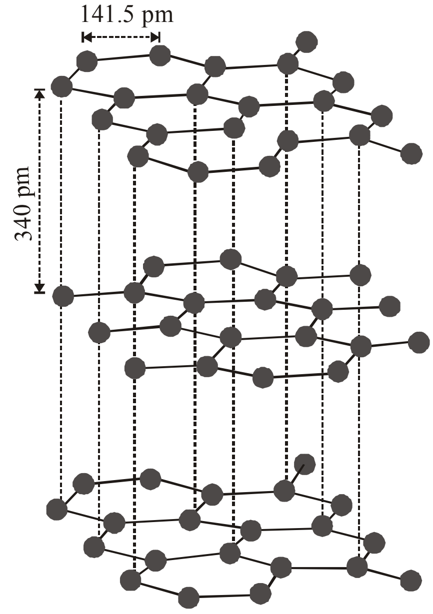
Diamond is covalent, yet it has high melting point. Why ?

Solution

Diamond has a three-dimensional network involving strong C—C bonds, which are very difficult to break and, in turn has high melting point.

Graphite *(Thermodynamically most stable allotrope of carbon)*

Graphite has layered structure. Layers are held by van der Waals forces and distance between two layers is 340 pm. Each layer is composed of planar hexagonal rings of carbon atoms. C—C bond length within the layer is 141.5 pm. Each carbon atom in hexagonal ring undergoes sp^2 hybridisation and makes three sigma bonds with three neighbouring carbon atoms. Fourth electron forms a π bond. The electrons are delocalised over the whole sheet. Thus, graphite is lusturous. Electrons are mobile and, therefore, graphite conducts electricity along the sheet. Graphite cleaves easily between the layers and, therefore, it is very soft and slippery. For this reason graphite is used as a dry lubricant in machines running at high temperature, where oil cannot be used as a lubricant.



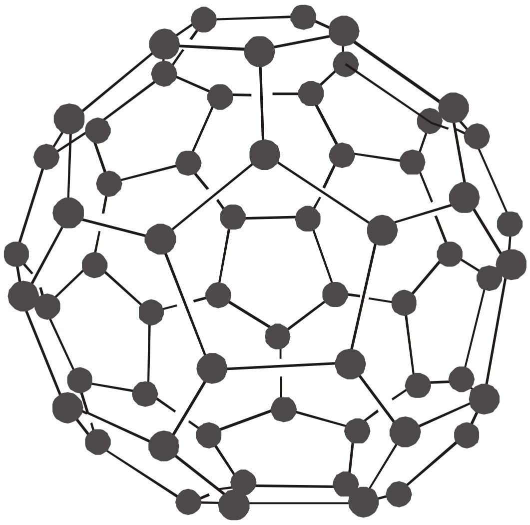
The structure of graphite

Fullerenes

Fullerenes are made by the heating of graphite in an electric arc in the presence of inert gases such as helium or argon. The sooty material formed by condensation of vapourised C_n small molecules consists of mainly C_{60} with smaller quantity of C_{70} and traces of fullerenes consisting of even number of carbon atoms up to 350 or above. Fullerenes are the only pure form of carbon because they have smooth structure without having ‘dangling’ bonds. Fullerenes are cage like molecules. C_{60} molecule has a shape like soccer ball and called **Buckminsterfullerene.**

It contains twenty, six- membered rings and twelve, five membered rings. A six membered ring is fused with six or five membered rings but a five membered ring can only fuse with six membered rings. All the carbon atoms are equal and they undergo *sp*^{2} hybridisation. Each carbon atom forms three sigma bonds with other three carbon atoms. The remaining electron at each carbon is delocalised in molecular orbitals, which in turn give aromatic character to molecule. However, because of non-planar nature of fullerenes, their aromatic character gets diminished and reactivity increases.

This ball shaped molecule has 60 vertices and each one is occupied by one carbon atom and it also contains both single and double bonds with C–C distances of 143.5 pm and 138.3 pm respectively. Spherical fullerenes are also called *bucky balls* in short.



[The structure of C_{60}, Buckminsterfullerene : Note that molecule has the shape of a soccer ball (football)]

It is very important to know that graphite is thermodynamically most stable allotrope of carbon and, therefore, ΔH_f of graphite is taken as zero. ΔH_f values of diamond and fullerene, C_{60} are 1.90 and 38.1 kJ mol^{-1}, respectively.

Other forms of elemental carbon like carbon black, coke, and charcoal are all impure forms of graphite or fullerenes. Carbon black is obtained by burning hydrocarbons in a limited supply of air. Charcoal and coke are obtained by heating wood or coal respectively at high temperatures in the absence of air.

Uses of Carbon

Graphite fibres embedded in plastic material form high strength, lightweight composites. The composites are used in products such as tennis rackets, fishing rods, aircrafts and canoes. Being good conductor, graphite is used for electrodes in batteries and industrial electrolysis. Crucibles made from graphite are inert to dilute acids and alkalies. Being highly porous, activated charcoal is used in adsorbing poisonous gases; also used in water filters to remove organic contaminators and in airconditioning system to control odour. Carbon black is used as black pigment in black ink and as filler in automobile tyres. Coke is used as a fuel and largely as a reducing agent in metallurgy. Diamond is a precious stone and used in jewellery. It is measured in carats (1 carat = 200 mg).

ALLOTROPES OF OXYGEN FAMILY

Oxygen occurs as two non-metallic forms, dioxygen O_2 and ozone O_3. Dioxygen O_2 is stable as a diatomic molecule. O_3 molecule is diamagnetic while O_2 is paramagnetic.

Six allotropes of selenium are known.

Tellurium has only one crystalline form, which is silvery white and semimetallic.

This is similar to grey Se, but has stronger metallic interaction.

Sulphur Allotropic forms :

Sulphur forms numerous allotropes of which the yellow rhombic (α-sulphur) and monoclinic (β -sulphur) forms are the most important. The stable form at room temperature is rhombic sulphur, which transforms to monoclinic sulphur when heated above 369 K.

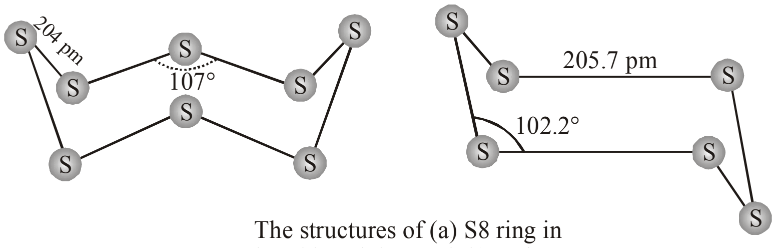
Rhombic sulphur (α-sulphur) (*Most stable sulphur allotrope at room temperature*)

This allotrope is yellow in colour, m.p. 385.8 K and specific gravity 2.06. Rhombic sulphur crystals are formed on evaporating the solution of roll sulphur in CS_2. It is insoluble in water but dissolves to some extent in benzene, alcohol and ether. It is readily soluble in CS_2.

Monoclinic sulphur (β-sulphur)

Its m.p. is 393 K and specific gravity 1.98. It is soluble in CS_2 while water insoluble. This form of sulphur is prepared by melting rhombic sulphur in a dish and cooling, till crust is formed. Two holes are made in the crust and the remaining liquid poured out. On removing the crust, colourless needle shaped crystals of β-sulphur are formed. It is stable above 369 K and transforms into α-sulphur below it. Conversely, α-sulphur is stable below 369 K and transforms into β-sulphur above this. At 369 K, both the forms are stable. This temperature is called transition temperature.

Both rhombic and monoclinic sulphur have S_8 molecules. These S_8 molecules are packed to give different crystal structures. The S_8 ring in both the forms is puckered and has a crown shape. The molecular dimensions are given in the figure.



The structures of (a) S8 ring in rhombic sulphur and (b) S_6 form

Several other modifications of sulphur containing 6-20 sulphur atoms per ring have been synthesised in the last two decades. In cyclo-S_6, the ring adopts the chair form and the molecular dimensions are as shown in figure. At elevated temperatures (~1000 K), S_2 is the dominant species and is paramagnetic just like O_2.

Note : Engel's sulphur (ε-sulphur) is unstable and contains S_6 rings arranged in the chair conformation. It is made by pouring $Na_2S_2O_3$ solution into concentrated HCl and extracting the S with toluene. Plastic sulphur-(χ) is obtained by pouring liquid sulphur into water.

Example

Which form of sulphur shows paramagnetic behaviour ?

Solution

In vapour state sulphur partly exists as S_2 molecule which has two unpaired electrons in the antibonding π^* orbitals like O_2 and, hence, exhibits paramagnetism.

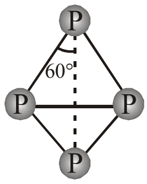
Allotropes of Nitrogen Family :

Solid As, Sb and Bi each exist in several allotropic forms. Arsenic vapour contains tetrahedral As_4 molecules. A reactive yellow form of the solid resembles white phosphorus and is thought to contain tetrahedral As_4 units. Sb also has a yellow form. All three elements have much less reactive metallic or α-forms.

PHOSPHORUS ALLOTROPIC FORMS

Phosphorus is found in many allotropic forms, the important ones being white, red and black. White phosphorus is a translucent, white, waxy solid. It is poisonous, insoluble in water but soluble in carbon disulphide and glows in dark (chemiluminescence). It dissolves in boiling NaOH solution in an inert atmosphere giving PH_3.

$$P_4 + 3NaOH + 3H_2O \longrightarrow PH_3 + \underset{\text{(Sodium hypophosphite)}}{3NaH_2PO_2}$$



White Phosphorus

White phosphorus is less stable and therefore, more reactive than the other solid phases under normal conditions because of angular strain in the P_4 molecule where the angles are only 60°. It readily catches fire in air to give dense white fumes of P_4O_{10}.

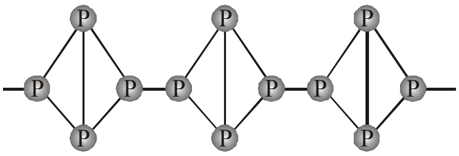
$$P_4 + 5O_2 \rightarrow P_4O_{10}$$

It consists of discrete tetrahedral P_4 molecule as shown in Fig.

Red phosphorus is obtained by heating white phosphorus at 573K in an inert atmosphere for several days. When red phosphorus is heated under high pressure, a series of phases of black phosphorus is formed.

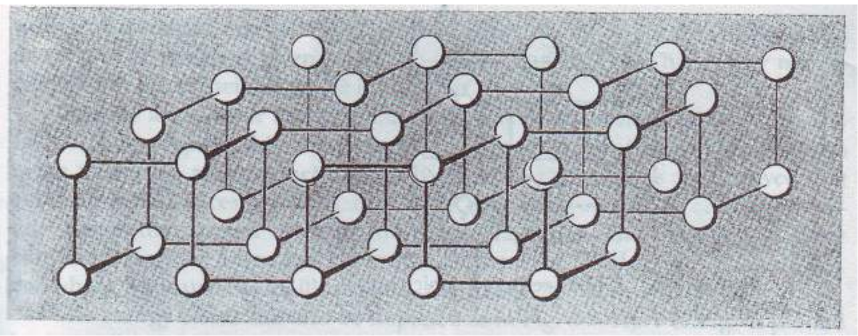
Red phosphorus possesses iron grey lustre. It is odourless, nonpoisonous and insoluble in water as well as in carbon disulphide. Chemically, red phosphorus is much less reactive than white phosphorus. It does not glow in the dark.

It is polymeric, consisting of chains of P_4 tetrahedra linked together in the manner as shown in Fig.



Red Phosphorus

Black phosphorus has two forms α-black phosphorus and β-black phosphorus. α-Black phosphorus is formed when red phosphorus is heated in a sealed tube at 803K. It can be sublimed in air and has opaque monoclinic or rhombohedral crystals. It does not oxidise in air. β-Black phosphorus is prepared by heating white phosphorus at 473 K under high pressure. It does not burn in air upto 673 K.



Thermodynamic stability order : Black phosphorus > Red phosphorus > White phosphorus

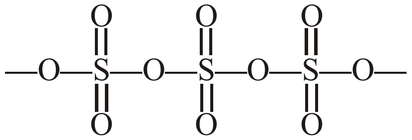
Reactivity order : Black phosphorus < Red phosphorus < White phosphorus

ALLOTROPIC FORM OF SO_3

SO_3 have three allotropic forms α-SO_3 , β-SO_3 and γ-SO_3

α-SO_3
Cross linked solid

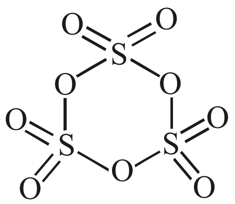
,



$$-O-\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}-O-\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}-O-\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}-O-$$

Polymeric chain structure
β-SO_3

,



γ-SO_3
(S_3O_9)
cylic trimer

QUESTION OF HYDROGEN BONDING

Q.1 Explain the structure of Boric acid in solid state.

Q.2 Boiling point of o-Nitrophenol is less than meta and para nitrophenol. Why?

Q.3 Maleic acid is more acidic than fumaric acid. Why?

Q.4 H – F is only liquid among halogen acid. Why?

Q.5 Ammonia is more easily liquefied than HCl, explain.

Q.6 Why ice floats on water?

Q.7 Water shows maximum density at 4°C. Why?

Q.8 HI is the strongest halogen acid, whereas H–F is the weakest. Why?

Q.9 Wood pieces are used to hold ice-cream. Why?

Q.10 KHF_2 is possible but not $KHBr_2$ or KHI_2. Why?

Q.11 O – Nitrophenol is less soluble in H_2O than p – Nitrophenol. Why?

Q.12 o-Hydroxy benzaldehyde is a liquid at room temperature while p-hydroxy benzaldehyde is a high melting solid.

Q.13 Glycerol is more viscous than ethanol. Explain.

Q.14 CH_4 and H_2O have nearly same molecular weight. Yet CH_4 has a boiling point 112 K and water 373 K. Explain.

Q.15 The experimental molecular weight of acetic acid in just double than theoretical molecular weight of acetic acid. Why?

Q.16 Although chlorine has same electronegativity as nitrogen but the former does not form effective H-bonding. Explain.

Q.17 Molar entropy change of vapourization of acetic acid is less than that of water. Explain

Q.18 Heat of vapourization of water is higher than HF, however strength of H-bond in HF is higher than water. Explain

EXERCISE # O-1

IONIC BOND

(Only one option is correct)

1. The compound which contains ionic as well as covalent bonds is

(A) $C_2H_4Cl_2$ (B) CH_3I (C) KCN (D) H_2O_2

2. A bond formed between two like atoms cannot be

(A) ionic (B) covalent (C) coordinate (D) metallic

3. An ionic bond A^+B^- is most likely to be formed when :

(A) the ionization energy of A is high and the electron gain enthalpy of B is low

(B) the ionization energy of A is low and the electron gain enthalpy of B is high

(C) the ionization energy of A and the electron gain enthalpy of B both are high

(D) the ionization energy of A and the electron gain enthalpy of B both are low

COVALENT BOND , CO-ORDINATE BOND & LEWIS STRUCTURE

4. PCl_5 exists but NCl_5 does not because :

(A) Nitrogen has no vacant 2*d*-orbitals (B) NCl_5 is unstable

(C) Nitrogen atom is much smaller than P (D) Nitrogen is highly inert

5. Which of the following species are hypervalent?

(1) PCl_5, (2) BF_3, (3) XeF_2 , (4) CO_3^{2-}

(A) 1, 2, 3 (B) 1, 3 (C) 3, 4 (D) 1, 2

6. The types of bond present in $N_2O_{5(g)}$ are

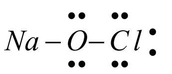
(A) only covalent (B) only ionic (C) ionic and covalent (D) covalent & coordinate

7. Which of the following molecule does not have coordinate bonds?

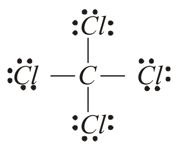
(A) $CH_3–NC$ (B) CO (C) O_3 (D) CO_3^{2-}

8. Which of the following Lewis dot diagrams is(are) incorrect ?

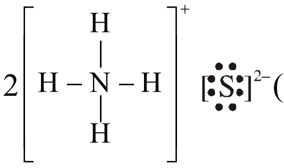
(A)



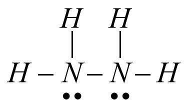
(B)



(C)

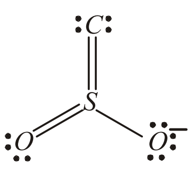


(D)

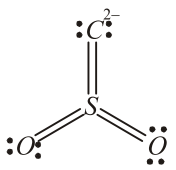


9. The possible structure(s) of monothiocarbonate ion is :

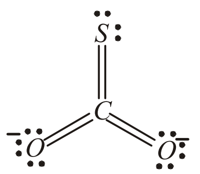
(A)



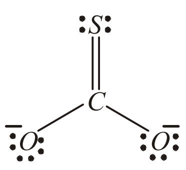
(B)



(C)



(D)



10. The valency of sulphur in sulphuric acid is :

(A) 2 (B) 8 (C) 4 (D) 6

V.B.T., HYBRIDISATION & VSEPR THEORY

11. In the following compound $\overset{1}{C}H_2 = \overset{2}{C}H - \overset{3}{C}H_2 - C \equiv CH$, the $C_2 - C_3$ bond is of the type :

(A) $sp–sp^2$ (B) $sp^3–sp^3$ (C) $sp–sp^3$ (D) $sp^2–sp^3$

12. Which of the following has a geometry different from the other three species (having the same geometry)?

(A) BF_4^- (B) SO_4^{2-} (C) XeF_4 (D) PH_4^+

13. Maximum bond energy is in :

(A) F_2 (B) N_2 (C) O_2 (D) equal

14. The hybridisation and shape of BrF_3 molecule are :

(A) sp^3d and bent *T* shape (B) sp^2d^2 and tetragonal

(C) sp^3d and bent (D) none of these

15. The shape of methyl cation (CH_3^+) is likely to be:

(A) linear (B) pyramidal (C) planar (D) spherical

16. The structure of XeF_2 involves hybridization of the type :

(A) sp^3 (B) sp^2 (C) sp^3d (D) sp^3d^2

17. Which of the following has been arranged in increasing order of length of the hybrid orbitals ?

(A) $sp < sp^2 < sp^3$ (B) $sp^3 < sp^2 < sp$ (C) $sp^2 < sp^3 < sp$ (D) $sp^2 < sp < sp^3$

18. In the context of carbon, which of the following is arranged in the correct order of electronegativity :

(A) $sp > sp^2 > sp^3$ (B) $sp^3 > sp^2 > sp$ (C) $sp^2 > sp > sp^3$ (D) $sp^3 < sp < sp^2$

19. When 2s–2s, 2p–2p and 2p–2s orbitals overlap, the bond strength decreases in the order :

(A) $p - p > s - s > p - s$ (B) $p - p > p - s > s - s$

(C) $s - s > p - p > p - s$ (D) $s - s > p - s > p - p$

20. Carbon atoms in $C_2(CN)_4$ are :

(A) *sp*-hybridized (B) sp^2-hybridized

(C) *sp*- and sp^2 hybridized (D) *sp*, sp^2 and sp^3- hybridized

21. The bond angle and hybridization in ether (CH_3OCH_3) is :

(A) 106°51', sp^3 (B) 104°31', sp^3 (C) > 109° 28', sp^3 (D) None of these

22. The structure of XeF_6 in vapour phase is

(A) pentagonal bipyramidal (B) trigonal bipyramidal

(C) capped octahedron (D) square bipyramidal

DIPOLE MOMENT & HYDROGEN BONDING

23. Which of the following has been arranged in order of decreasing dipole moment ?

(A) $CH_3Cl > CH_3F > CH_3Br > CH_3I$ (B) $CH_3F > CH_3Cl > CH_3Br > CH_3I$

(C) $CH_3Cl > CH_3Br > CH_3I > CH_3F$ (D) $CH_3F > CH_3Cl > CH_3I > CH_3Br$

24. The experimental value of the dipole moment of *HCl* is 1.03 *D*. The length of the *H–Cl* bond is 1.275 Å. The percentage of ionic character in *HCl* is nearly :

(A) 43 (B) 21 (C) 17 (D) 7

25. Ethanol has a higher boiling point than dimethyl ether though they have the same molecular weight. This is due to :

(A) resonance (B) coordinate bonding

(C) hydrogen bonding (D) ionic bonding

26. Which of the following compounds would have significant intermolecular hydrogen bonding ? HF, CH_3OH, N_2O_4, CH_4

(A) HF, N_2O_4 (B) HF, CH_4, CH_3OH (C) HF, CH_3OH (D) CH_3OH, CH_4

27. For H_2O_2, H_2S, H_2O and HF, the correct order of decreasing strength of hydrogen bonding is :

(A) $H_2O > HF > H_2O_2 > H_2S$ (B) $HF > H_2O_2 > H_2O > H_2S$

(C) $HF > H_2O > H_2O_2 > H_2S$ (D) $H_2O_2 > H_2O > HF > H_2S$

28. Which one of the following does not have intermolecular H-bonding?

(A) H_2O (B) *o*-nitro phenol (C) HF (D) CH_3COOH

29. The order of strength of hydrogen bonds is:

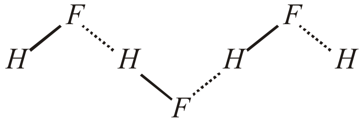
(A) NH...N > OH...O > FH...F (B) NH...N < OH...O < FH...F

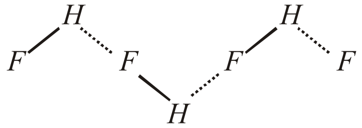
(C) NH...N > OH...O < FH...F (D) NH...N < OH...O > FH...F

30. The *H* bond in solid *HF* can be best represented as:

(A) H – F....H – F....H – F

(B)


(C)


(D)


MISCELLEANEOUS

31. The formal charges on the three O-atoms in O_3 molecule are

(A) 0, 0, 0 (B) 0, 0, –1 (C) 0, 0, +1 (D) 0, +1, –1

32. The types of bonds present in $CuSO_4 \cdot 5H_2O$ are

(A) electrovalent and covalent

(B) electrovalent , coordinate covalent & H-bond

(C) covalent, coordinate covalent & H-bonds

(D) electrovalent, covalent, coordinate covalent & H-bond

33. If ethylene molecule lies in X-Y plane then nodal planes of the π-bond will lie in

(A) XZ plane (B) YZ plane

(C) In a plane that bisects C–C axis (D) XY plane

34. Two ice cubes are pressed over each other and unite to form one cube. Which force is responsible for holding them together :

(A) van der Waal's forces (B) Covalent attraction

(C) Hydrogen bond formation (D) Dipole-dipole attraction

EXERCISE # O-2

IONIC BOND , COVALENT BOND & COORDINATE BOND

1. Which of the following compounds contain ionic, covalent and co-ordinate bonds?

(A) NH_4Cl (B) KCN (C) $NaBF_4$ (D) NaOH

2. Which of the following statements is/are true?

(A) Covalent bonds are directional

(B) Ionic bonds are nondirectional

(C) A polar bond is formed between two atoms which have the same electronegativity value.

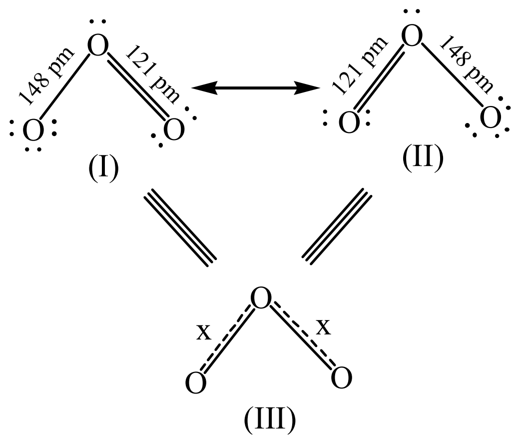
(D) The presence of polar bonds in a polyatomic molecule suggests that it has zero dipole moment

3. To which of the following species octet rule is not applicable :

(A) BrF_5 (B) SF_6 (C) IF_7 (D) CO

LEWIS STRUCTURE & BOND LENGTH

4. O_3(ozone) molecule can be equally represented by the structures I and II shown below



which of the following values of x are not correct

(A) 148 pm (B) 121 pm (C) less than 121 pm (D) more than 148 pm

5. Select the correct statement -

(A) C–O bond length of CO molecule is higher than the C–O bond length of CO_2

(B) C–O bond length of CO molecule is lower than the C–O bond length of CO_2

(C) N–O bond length of NO_3^- ion is higher than the N–O bond length of NO_2^-

(D) N–O bond length of NO_3^- ion is lower than the N–O bond length of NO_2^-

VBT, HYBRIDISATION & VSEPR THEORY

6. Which of the following statement(s) is / are not correct?

(A) Hybridization is the mixing of atomic orbitals of large energy difference.

(B) sp^2–hybrid orbitals are formed from two *p* - atomic orbitals and one *s*- atomic orbital

(C) sp^3d^2–hybrid orbitals are all at 90º to one another

(D) sp^3–hybrid orbitals are directed towards the corners of a regular tetrahedron

7. Which of the following species are linear ?

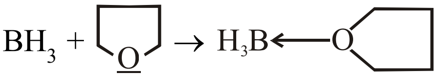
(A) ICl_2^- (B) I_3^- (C) N_3^- (D) ClO_2^+

8. Which of the following molecules have bent shape.
Where A is central atom , B is surrounding atom and E is lone pair -

(A) AB_2E_2 (B) AB_2E (C) AB_3E (D) AB_2

9. In which of the following process(s) hybridisation of underlined atom does not change –

(A) $\underline{N}H_3 + BF_3 \rightarrow H_3N.BF_3$ (B) $\underline{Si}F_4 + 2F^{\ominus} \rightarrow [SiF_6]^{2-}$

(C) 

(D) $H_3\underline{B}O_3 \xrightarrow{OH^-} [B(OH)_4]^-$

10. Which of the following 'd' orbital(s) participate in the hybridistaion for Xe in the cationic part of XeF_6(s)

(A) d_{xy} (B) $d_{x^2-y^2}$ (C) d_{z^2} (D) d_{y^2}

DIPOLE MOMENT & HYDROGEN BOND

11. Which of the following compounds possesses zero dipole moment?

(A) Water (B) Benzene

(C) Carbon tetrachloride (D) Boron trifluoride

12. Which of the following statements are correct?

(A) The crystal lattice of ice is formed by covalent as well as hydrogen bonds

(B) The density of water increases when heated from 0° C to 4° C

(C) Above 4° *C* the thermal agitation of water molecules increases. Therefore, intermolecular distance increases and water starts expanding

(D) The density of water decreases from 0° C to a maximum at 4° C

13. In which of the following chemical species hydrogen bonding takes places -

(A) $Na_2HPO_3(s)$ (B) $K_2HPO_4(s)$ (C) $KHCO_3(s)$ (D) $NaHCO_3(s)$

MISCELLEANEOUS & STRUCTURE

14. Which of the following oxyacids of sulphur contain $S-S$ bonds ?

(A) $H_2S_2O_8$ (B) $H_2S_2O_6$ (C) $H_2S_2O_4$ (D) $H_2S_2O_5$

15. Molecule(s) having both polar and non polar bonds is / are

(A) O_2F_2 (B) S_2Cl_2 (C) N_2H_4 (D) S_2F_{10}

16. Which of the following molecule is/are planar -

(A) CH_2Cl_2 (B) $B_3N_3H_6$ (C) SO_3 (D) NF_3

EXERCISE # S-1

Integer Type Answer :

1. In solid $\overset{+}{Na}Cl^{-}$ one Na^+ is surrounded by six Cl^- ions find out total number of directional bonds formed by each Na^+ with Cl^- ions.

2. Find out total number of $p\pi - d\pi$ bonds present in SO_2 molecule

3. Find out total number of $p\pi - p\pi$ bonds present in SO_4^{2-}

4. How many number of atomic orbitals are required, so that their mixing produces hybrid orbitals each having % of s-character nearly equal to 17%.

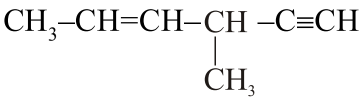
5. Calculate the maximum number of atoms lying in one plane of PCl_3F_2.

6. Find the lone pair-bond pair ratio for $[SiF_6]^{2-}$

7. A diatomic molecule has a dipole moment of 1.2 D. If bond length is 1.0 Å, what percentage of an electronic charge exists on each atom.

[Add the digits untill you are getting single digit. e.g. If the answer is 537, then 5 + 3 + 7 = 15] again 1 + 5 = 6 and 6 is to be written as answer.

8. Total number of sp^3 hybridised atom(s) in given hydrocarbon.



$CH_3{-}CH{=}CH{-}\underset{\displaystyle CH_3}{\underset{|}{CH}}{-}C{\equiv}CH$

9. Find the number of lone pair on central atom of ICl_4^-

10. Find the number of acid(s) which are having hypo prefix in it's name from the following.

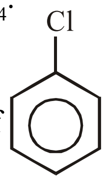
H_3PO_4, H_3BO_3, H_3PO_3, H_3PO_2, $HClO_3$, $HClO$, $(HNO)_2$, $H_4P_2O_6$

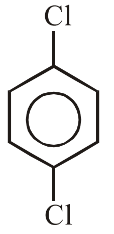
11. Find the number of planar molecules/ions which are cyclic –

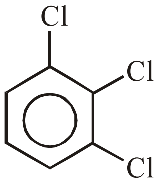
$B_3N_3H_6$, $C_3N_3Cl_3$, $(BO_2^-)_3$, $S_2O_8^{2-}$

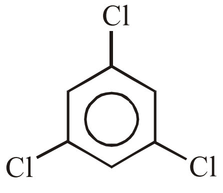
12. Find the maximum number of plane having maximum number of atoms in CH_4.

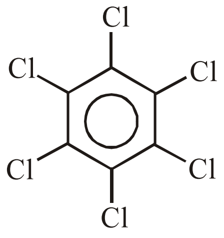
13. Find the number of molecule having dipole moment less than the dipole moment of











14. Find the number of molecules having intramolecular hydrogen bonding in the following :

ice, o-boric acid, o-nitrophenol, o-chlorophenol, chloral hydrate, o-salicyaldehyde

15. Find the number of chemical species which are planar from in the following.

$XeF_5^{\ominus}$, XeF_5^{+}, HCN, CO_2, SO_2

EXERCISE # S-2

(Comprehension) (Q.1 to Q.2)

Bond angle and bond polarity are important parameters associated for bonding in molecule.

1. Which of the following bond angle order is incorrect.

(A) $NH_3 > PH_3$
(B) $BF_3 < BH_3$
(C) $O(CH_3)_2 < O(C_2H_5)_2$
(D) $CO_3^{2-} < CO_2$

2. Which of the following molecule is polar and non-planar

(A) CH_2Cl_2
(B) PCl_5
(C) SO_2
(D) XeF_4

(Comprehension) (Q.3 to Q.4)

During the formation of a covalent bond each participating atom usually acquires electronic configuration to that of nearest noble gas, and the shared electron pair remains localized between the bonded nuclei. Besides these shared electrons there are certain electron pairs which remains localized on C.A. or on substituent or both and are called non bonding or lone pair electrons.

3. Which of the following species does / do not follow octet rule:

(A) Hypovalent
(B) Hypervalent
(C) Odd electron molecules
(D) All

4. In SO_3^{2-} the total number of lone pairs and bond pairs are respectively

(A) 18 and 4
(B) 9 and 8
(C) 9 and 4
(D) 7 and 4

(Comprehension) (Q.5 to Q.7)

Hybridisation is a theoretical concept, as state of hybridisation cannot be detected even by spectroscopically; unlike intermediates or transition state in various reactions. but it corrects the predictions which are based simple on overlapping of pure atomic orbitals. VSEPR theory predicts precisely shape and bond angle in a given molecule.

5. In which pair of molecules bond angles are not same:

(A) CCl_4 & $SiCl_4$
(B) NH_4^+ & $\overset{+}{N}F_4$
(C) ClF_6^+ & SF_6
(D) None

6. The molecules / ions which are planar as well as polar.

(A) BF_3 , H_2O, H F , NH_2^-
(B) $SnCl_2$, I_3^+ , NH_2^-, IF_3
(C) CO_3^{2-} , I_3^-, SF_2, XeF_2
(D) NO_2^-, XeF_4, ICl_4^-, NH_2^+

7. The correct order of energy levels of hybrid orbitals.

(A) $sp > sp^2 > sp^3$
(B) $sp < sp^2 < sp^3$
(C) $sp^2 > sp^3 > sp$
(D) $sp^3 > sp > sp^2$

(Comprehension) (Q.8 to Q.10)

As one moves from sp hybridisation to sp^3. % of s-character in hybrid orbital decreases from 50% to 25% and p-character increases from 50% to 75% and in any hybrid orbital total % of s & p- character remains 100%. By increasing p-character the hybrid orbitals become elongated hence, their overlapping extent decreases that is results into weak bond energy also bond angle decreases.

8. **Statement-1 :** On decreasing s-character in hybrid orbitals, bond angle decreases.

Statement-2 : p-orbitals are at 90° to one another.

(A) Statement-1 is true, statement-2 is true and statement-2 is correct explanation for statement-1.
(B) Statement-1 is true, statement-2 is true and statement-2 is NOT the correct explanation for statement-1.
(C) Statement-1 is true, statement-2 is false.
(D) Statement-1 is false, statement-2 is true.

9. The type of overlapping which produces bond of maximum bond energy is:

(A) $sp^3 - 1s$ (B) $sp^2 - 1s$

(C) $sp - 1s$ (D) All have same bond energy

10. The molecule having bond pair in which % of s character is minimum for the central atom's hybrid orbital.

(A) SiF_4 (B) BCl_3 (C) CH_3^+ (D) CO_2

Match the Column :

11. Match the column

Column-I (molecules)	Column-II (correct about molecule)
(A) $\underline{P}Cl_5(\theta = 120^\circ), (\theta' = 90^\circ)$	(P) Number of the given bond angle (θ or θ') is = One
(B) $\underline{N}H_3(\theta = 107^\circ)$	(Q) Number of the given bond angle (θ or θ') is = Three
(C) $\underline{Si}H_4(\theta = 109^\circ\ 28')$ (R)	(R) Number of the given bond angle (θ or θ') is = Six
(D) $\underline{S}O_2(\theta = 116^\circ)$	(S) d-orbital(s) is/are involved in hybrdisation of central atom
	(T) Lone pair at underlined atom is present

12.

Column I	Column II
(A) BrF_3	(P) Atleast one angle $\leq 90^\circ$
(B) TeF_5^-	(Q) Central atom is sp^3d^2 hybridised
(C) IF_7	(R) Non planar
(D) XeF_4	(S) Polar

13.

Column I	Column II
(A) $I(CN)_2^-$	(P) Having $p\pi-p\pi$ bond & $\mu_D = 0$
(B) CO_3^{2-}	(Q) Having $p\pi-d\pi$ bond & $\mu_D \neq 0$
(C) XeO_2F_2	(R) Planar
(D) SOF_4	(S) Central atom is sp^3d

14.

Column I	Column II
(A) $H_2S_2O_5$	(P) Central atom is sp^3 hybridised
(B) $H_6B_2O_7^{2-}$	(Q) M–O–M i.e. oxo linkage is present
(C) $H_4P_2O_6$(tetrabasic)	(R) M–M linkage is present
(D) $H_6Si_2O_7$	(S) Non planar

15.

Column–I	Column–II
(P) XeF_4	(1) sp^3d hybridised with 2 lone pair on central atom
(Q) BrF_5	(2) sp^3d^2 hybridised with 2 lone pair on central atom
(R) ClF_3	(3) sp^3 hybridised with 1 lone pair on central atom
(S) SO_3^{2-}	(4) sp^3d^2 hybridised with 1 lone pair on central atom

Code :

	P	Q	R	S
(A)	1	2	3	4
(B)	2	4	1	3
(C)	2	4	3	1
(D)	3	1	4	2

16. **Column–I** **Column–II**

	Column–I		Column–II
(P)	NH_4Cl	(1)	diamond like structure
(Q)	$K[HF_2]$	(2)	Benzene like structure
(R)	$B_3N_3H_6$	(3)	Both hydrogen bond and ionic bond exist
(S)	SiC	(4)	Both covalent bond and ionic bond exist

Code :

	P	Q	R	S
(A)	4	3	2	1
(B)	3	4	2	1
(C)	3	4	1	2
(D)	4	1	3	2

MATCHING LIST TYPE 1 × 3 Q. (THREE LIST TYPE Q.)

Column - I	Column - II	Column - III
(P) d_{yz}	(1) p_y	(i) 1 lobe – 1 lobe overlap
(Q) s	(2) p_x	(ii) 2 lobe – 2 lobe overlap
(R) d_{xz}	(3) d_{yz}	(iii) 4 lobe – 4 lobe overlap
(S) p_z	(4) s	(iv) Zero overlap

17. Which of the following combination results in π-bond formation if internuclear axis is z-axis

(A) (P), (1), (ii) (B) (P), (4), (i) (C) (R), (4), (iv) (D) (R), (3), (ii)

18. Which of the following combination resultsin σ-bond formation if internuclear axis is z-axis

(A) (P), (1), (ii) (B) (Q), (3), (ii) (C) (Q), (4), (i) (D) (R), (2), (ii)

19. Which of the following combination results in δ-bond formation if internuclear axis is x-axis

(A) (P), (1), (i) (B) (P), (3), (iii) (C) (R), (3), (iv) (D) (P), (2), (ii)

EXERCISE # JEE-MAIN

1. Which of the following statements is true ? **[AIEEE-2002]**

(1) HF is less polar than HBr

(2) Water does not contain any ions

(3) Chemical bond formation takes place when forces of attraction overcome the forces of repulsion

(4) In covalent bond, transfer of electrons takes place

2. Which one of the following pairs of molecules will have permanent dipole moments for both members

(1) NO_2 and CO_2 (2) NO_2 and O_3 **[AIEEE-2003]**

(3) SiF_4 and CO_2 (4) SiF_4 and NO_2

3. The pair of species having identical shapes for molecules of both species is **[AIEEE-2003]**

(1) XeF_2, CO_2 (2) BF_3, PCl_3 (3) PF_5, IF_5 (4) CF_4, SF_4

4. The correct order of bond angles (smallest first) in H_2S, NH_3, BF_3 and SiH_4 is :–**[AIEEE–2004]**

(1) $H_2S < NH_3 < SiH_4 < BF_3$ (2) $NH_3 < H_2S < SiH_4 < BF_3$

(3) $H_2S < SiH_4 < NH_3 < BF_3$ (4) $H_2S < NH_3 < BF_3 < SiH_4$

5. Lattice energy of an ionic compound depends upon **[AIEEE-2005]**

(1) charge on the ion only (2) size of the ion only

(3) packing of the ion only (4) charge and size of the ion

6. Of the following sets which one does not contain isoelectronic species ? **[AIEEE-2005]**

(1) PO_4^{3-}, SO_4^{2-}, ClO_4^- (2) CN^-, N_2, C_2^{2-}

(3) SO_3^{2-}, CO_3^{2-}, NO_3^- (4) BO_3^{3-}, CO_3^{2-}, NO_3^-

7. In which of the following molecules/ions are all the bonds **not** equal? **[AIEEE-2006]**

(1) XeF_4 (2) BF_4^- (3) SF_4 (4) SiF_4

8. Which of the following hydrogen bonds is the strongest **[AIEEE-2007]**

(1) F–H.....F (2) O–H.....O (3) O–H.....F (4) O–H.....N

9. The hybridisation of orbitals of N atom in NO_3^-, NO_2^+ and NH_4^+ are respectively:- **[AIEEE-2011]**

(1) sp, sp^3, sp^2 (2) sp^2, sp^3, sp (3) sp, sp^2, sp^3 (4) sp^2, sp, sp^3

10. The structure of IF_7 is :- **[AIEEE-2011]**

(1) octahedral (2) pentagonal bipyramid

(3) square pyramid (4) trigonal bipyramid

11. Which of the following has maximum number of lone pairs associated with Xe **[AIEEE-2011]**

(1) XeO_3 (2) XeF_4 (3) XeF_6 (4) XeF_2

12. The molecule having smallest bond angle is :- **[AIEEE-2012]**

(1) PCl_3 (2) NCl_3 (3) $AsCl_3$ (4) $SbCl_3$

13. In which of the following pairs the two species are not isostructural ? **[AIEEE-2012]**

(1) AlF_6^{3-} and SF_6 (2) CO_3^{2-} and NO_3^- (3) PCl_4^+ and $SiCl_4$ (4) PF_5 and BrF_5

14. The number of S–S bonds in SO_3, $S_2O_3^{2-}$, $S_2O_6^{2-}$ and $S_2O_8^{2-}$ respectively are :- **[JEE-MAINS-2012]**

(1) 1, 0, 1, 0 (2) 0, 1, 1, 0 (3) 1, 0, 0, 1 (4) 0, 1, 0, 1 **(On line)**

15. Among the following species which two have trigonal bipyramidal shape ? **[JEE-MAINS-2012]**

(I) NI_3 (II) I_3^- (III) SO_3^{2-} (IV) NO_3^- **(On line)**

(1) II and III (2) III and IV (3) I and IV (4) I and III

16. Which of the following has the square planar structure :- **[JEE-MAINS-2012]**

(1) NH_4^+ (2) CCl_4 (3) XeF_4 (4) BF_4^- **(On line)**

17. The compound of Xenon with zero dipole moment is :- **[JEE-MAINS-2012]**

(1) XeO_3 (2) XeO_2 (3) XeF_4 (4) $XeOF_4$ **(On line)**

18. Among the following the molecule with the lowest dipole moment is :- **[JEE-MAINS-2012]**

(1) $CHCl_3$ (2) CH_2Cl_2 (3) CCl_4 (4) CH_3Cl **(On line)**

19. The formation of molecular complex $BF_3 - NH_3$ results in a change in hybridisation of boron :- **[JEE-MAINS-2012]**

(1) from sp^3 to sp^3d (2) from sp^2 to dsp^2 **(On line)**

(3) from sp^3 to sp^2 (4) from sp^2 to sp^3

20. Trigonal bipyramidal geometry is shown by: **[JEE-MAINS-2013]**

(1) XeO_3F_2 (2) $XeOF_2$ (3) XeO_3 (4) $FXeOSO_2F$ **(On line)**

21. Which one of the following molecules is polar? **[JEE-MAINS-2013]**

(1) CF_4 (2) SbF_5 (3) IF_5 (4) XeF_4 **(On line)**

22. Oxidation state of sulphur in anions SO_3^{2-}, $S_2O_4^{2-}$ and $S_2O_6^{2-}$ increases in the orders : **[JEE-MAINS-2013]**

(1) $S_2O_6^{2-} < S_2O_4^{2-} < SO_3^{2-}$ (2) $SO_3^{2-} < S_2O_4^{2-} < S_2O_6^{2-}$ **(On line)**

(3) $S_2O_4^{2-} < SO_3^{2-} < S_2O_6^{2-}$ (4) $S_2O_4^{2-} < S_2O_6^{2-} < SO_3^{2-}$

23. XeO_4 molecule is tetrahedral having : **[JEE-MAINS-2013]**

(1) Two pπ-dπ bonds (2) Four pπ-dπ bonds **(On line)**

(3) One pπ-dπ bond (4) Three pπ-dπ bonds

24. Bond distance in HF is 9.17×10^{-11} m. Dipole moment of HF is 6.104×10^{-30} Cm. The percent ionic character in HF will be : (electron charge = 1.60×10^{-19} C) **[JEE-MAINS-2013]**

(1) 61.0% (2) 38.0% **(On line)**

(3) 35.5% (4) 41.5%

25. The shape of IF_6^- is : **[JEE-MAINS-2013]**

(1) distorted octahedron (2) Pyramidal **(On line)**

(3) Octahedral (4) Square antiprism

26. Which has trigonal bipyramidal shape ? **[JEE-MAINS-2013]**

(1) $XeOF_4$ (2) XeO_3 (3) XeO_3F_2 (4) $XeOF_2$ **(On line)**

27. The catenation tendency of C, Si and Ge is in the order Ge < Si < C. The bond energies (in kJ mol^{-1}) of C — C, Si —Si and Ge—Ge bonds are respectively : **[JEE-MAINS-2013 (On line)]**

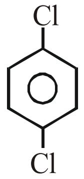
(1) 348, 260, 297 (2) 348, 297, 260 (3) 297, 348, 260 (4) 260, 297, 348

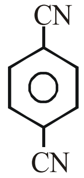
28. In which of the following sets, all the given species are isostructural ? **[JEE-MAINS-2013 (On line)]**

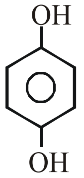
(1) BF_3, NF_3, PF_3, AlF_3 (2) PCl_3, $AlCl_3$, BCl_3, $SbCl_3$

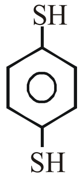
(3) BF_4^-, CCl_4, NH_4^+, PCl_4^+ (4) CO_2, NO_2, ClO_2, SiO_2

29. For which of the following molecule significant $\mu \neq 0$ **[JEE-M-2014]**

(a) 

(b) 

(c) 

(d) 

(1) Only (c) (2) (c) and (d) (3) Only (a) (4) (a) and (b)

30. Among the following oxoacids, the correct decreasing order of acid strength is :- **[JEE-M-2014]**

(1) $HClO_4 > HClO_3 > HClO_2 > HOCl$ (2) $HClO_2 > HClO_4 > HClO_3 > HOCl$

(3) $HOCl > HClO_2 > HClO_3 > HClO_4$ (4) $HClO_4 > HOCl > HClO_2 > HClO_3$

31. The number and type of bonds in C_2^{2-} ion in CaC_2 are:

[AIEEE-2005, AIEEE-2011, JEE-MAINS-2014 (On line)]

(1) Two σ bonds and one π – bond (2) Two σ bonds and two π – bonds

(3) One σ bond and two π – bonds (4) One σ bond and one π – bond

32. For the compounds **[JEE-MAINS-2014]**

CH_3Cl, CH_3Br, CH_3I and CH_3F, **(On line)**

the correct order of increasing C-halogen bond length is :

(1) $CH_3F < CH_3Br < CH_3Cl < CH_3I$ (2) $CH_3F < CH_3Cl < CH_3Br < CH_3I$

(3) $CH_3Cl < CH_3Br < CH_3F < CH_3I$ (4) $CH_3F < CH_3I < CH_3Br < CH_3Cl$

33. Shapes of certain interhalogen compounds are stated below. Which one of them is not correctly stated? **[JEE-MAINS-2014]**

(1) IF_7 : Pentagonal bipyramid (2) BrF_5 : Trigonal bipyramid **(On line)**

(3) ICl_3 : Planar dimeric (4) BrF_3 : Planar T-shaped

34. Which of the following molecules has two sigma(σ) and two pi(π) bonds :- **[JEE-MAINS-2014]**

(1) HCN (2) $C_2H_2Cl_2$ (3) N_2F_2 (4) C_2H_4 **(On line)**

35. The species in which the N atom is in a state of sp hybridization is **[JEE-MAINS(offline)-2016]**

(1) NO_2 (2) NO_2^+ (3) NO_2^- (4) NO_3^-

36. The pair in which phosphorous atoms have a formal oxidation state of + 3 is :-

(1) Pyrophosphorous and pyrophosphoric acids **[JEE-MAINS(offline)-2016]**

(2) Orthophosphorous and pyrophosphorous acids

(3) Pyrophosphorous and hypophosphoric acids

(4) Orthophosphorous and hypophosphoric acids

37. The group of molecules having identical shape is : **[JEE-MAINS(online)-2016]**

(1) SF_4, XeF_4, CCl_4 (2) ClF_3, $XeOF_2$, XeF_3^+

(3) PCl_5, IF_5, XeO_2F_2 (4) BF_3, PCl_3, XeO_3

38. ***Assertion :*** Among the carbon allotropes, diamond is an insulator, whereas, graphite is a good conductor of electricity. **[JEE-MAINS(online)-2016]**

Reason : Hybridization of carbon in diamond and graphite are sp^3 and sp^2, respectively.

(1) Assertion is incorrect statement, but the reason is correct.

(2)Both assertion and reason are correct, and the reason is the correct explanation for the assertion.

(3) Both assertion and reason are incorrect.

(4) Both assertion and reson are correct, but the reason is not the correct explanation for the assertion.

39. Aqueous solution of which salt will not contain ions with the electronic configuration $1s^22s^22p^63s^23p^6$? **[JEE-MAINS(online)-2016]**

(1) NaCl (2) CaI_2 (3) KBr (4) NaF

40. The correct sequence of decreasing number of π-bonds in the structures of H_2SO_3, H_2SO_4 and $H_2S_2O_7$ is:- **[JEE-MAINS(online)-2017]**

(1) $H_2S_2O_7 > H_2SO_4 > H_2SO_3$ (2) $H_2SO_3 > H_2SO_4 > H_2S_2O_7$

(3) $H_2S_2O_7 > H_2SO_3 > H_2SO_4$ (4) $H_2SO_4 > H_2S_2O_7 > H_2SO_3$

41. The increasing order of the boiling points for the following compounds is:-

(I) C_2H_5OH (II) C_2H_5Cl **[JEE-MAINS(online)-2017]**

(III) $C_2H_5CH_3$ (IV) $C_2H_5OCH_3$

(1) (III) < (II) < (I) < (IV) (2) (II) < (III) < (IV) < (I)

(3) (IV) < (III) < (I) < (II) (4) (III) < (IV) < (II) < (I)

42. The number of P–OH bonds and the oxidation state of phosphorus atom in pyrophosphoric acid ($H_4P_2O_7$) respectively are:- **[JEE-MAINS(online)-2017]**

(1) five and four (2) five and five

(3) four and five (4) four and four

43. The group having triangular planar structures is :- **[JEE-MAINS(online)-2017]**

(1) CO_3^{2-}, NO_3^-, SO_3 (2) NCl_3, BCl_3, SO_3

(3) NH_3, SO_3, CO_3^{2-} (4) BF_3, NF_3, CO_3^{2-}

44. The number of S=O and S–OH bonds present in peroxodisulphuric acid and pyrosulphuric acid respectively are : **[JEE-MAINS(online)-2017]**

(1) (2 and 4) and (2 and 4) (2) (4 and 2) and (2 and 4)

(3) (2 and 2) and (2 and 2) (4) (4 and 2) and (4 and 2)

EXERCISE # JEE-ADVANCED

Integer Type :

1. The number of water molecule(s) directly bonded to the metal centre in $CuSO_4.5H_2O$ is **[JEE 2009]**

2. Based on VSEPR theory, the number of 90 degree F–Br–F angles in BrF_5 is **[JEE 2010]**

3. The difference in the oxidation numbers of the two types of sulphur atoms in $Na_2S_4O_6$ is. **[JEE 2011]**

4. The total number of lone-pairs of electrons in melamine is **[JEE Adv. 2013]**

5. The sum of the number of lone pairs of electrons on each central atom in the following species is. **[JEE 2017]**

$[TeBr_6]^{2-}$, $[BrF_2]^+$, SNF_3 and $[XeF_3]^-$

[Atomic number : N = 7, F = 9, S = 16, Br = 35, Te = 52, Xe = 54]

One or more than one correct :

6. The linear structure is assumed by : **[IIT–1991]**

(A) $SnCl_4$ (B) NCO^- (C) CS_2 (D) NO_2^+

7. Which of the following has zero dipole moment ? **[IIT–1993]**

(A) ClF (B) PCl_3 (C) SiF_4 (D) $CFCl_3$

8. Among the following molecules, which one is planar ? **[IIT–1994]**

(A) BCl_3 (B) SO_2Cl_2 (C) NH_3 (D) NF_3

9. Which of the following species is / are non-linear ? **[IIT–1995]**

(A) H_2S (B) NH_3 (C) CO_2 (D) SO_2

10. Which contains both polar and non-polar bonds ? **[IIT–1997]**

(A) CH_4 (B) HCN (C) H_2O_2 (D) NH_4Cl

11. Which of the following compounds has sp^2 hybridisation ? **[IIT–1997]**

(A) CO_2 (B) SO_2 (C) N_2O (D) CO

12. Which one of the following molecules is planar ? **[IIT–1996]**

(A) NF_3 (B) NCl_3 (C) BF_3 (D) PH_3

13. The number and type of bonds between two carbon atoms in CaC_2 are : **[IIT–1996]**

(A) 1σ and 1π (B) 1σ and 2π (C) 1σ and 1.5π (D) 1σ

14. Among the species NF_3, NO_3^-, BF_3, H_3O^+ and HN_3, identify the isostructural species : **[IIT–1996]**

(A) (NF_3 , NO_3^-) and (BF_3 , H_3O^+) (B) (NF_3 , HN_3) and (NO_3^- , BF_3)

(C) (NF_3 , H_3O^+) and (NO_3^- , BF_3) (D) (NF_3 , H_3O^+) and (HN_3 , BF_3)

15. Among the following the one that is polar and has the central atom with sp^2 hybridisation is :

(A) H_2CO_3 (B) SiF_4 (C) BF_3 (D) $HClO_2$ **[IIT–1997]**

16. The geometry & the type of hybrid orbitals present about the central atom in BF_3 is : **[JEE '98]**

(A) linear, sp (B) trigonal planar, sp^2

(C) tetrahedral sp^3 (D) pyramidal, sp^3

17. Which one of the following statement (s) is (are) correct? **[JEE 1998]**

(A) The electronic configuration of Cr is [Ar] $3d^5\,4s^1$.(Atomic No. of Cr = 24)

(B) The magnetic quantum number may have a negative value

(C) In silver atom, 23 electrons have a spin of one type and 24 of the opposite type. (Atomic No. of Ag = 47)

(D) The oxidation state of nitrogen in HN_3 is –3.

18. Element X is strongly electronegative and element Y is strongly electropositive. Both are univalent. The compound formed would be : **[JEE 1998]**

(A) X→ Y (B) X^-Y^+ (C) X^+Y^- (D) Y→ X

19. The correct order of increasing C - O bond length of, CO, CO_3^{2-} , CO_2 is **[JEE '99]**

(A) $CO_3^{2-} < CO_2 < CO$ (B) $CO_2 < CO_3^{2-} < CO$

(C) $CO < CO_3^{2-} < CO_2$ (D) $CO < CO_2 < CO_3^{2-}$

20. The geometry of H_2S and its dipole moment are **[JEE '99]**

(A) angular & non zero (B) angular & zero

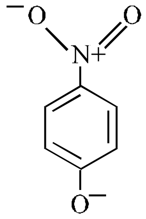
(C) linear & non zero (D) linear & zero

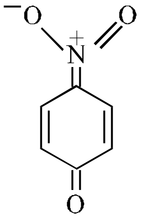
21. In compounds type E Cl_3, where E = B, P, As or Bi, the angles Cl - E - Cl for different E are in the order **[JEE '99]**

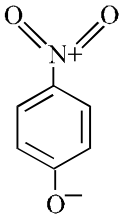
(A) B > P = As = Bi (B) B > P > As > Bi

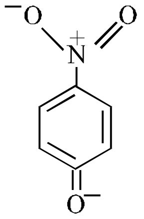
(C) B < P = As = Bi (D) B < P < As < Bi

22. The most likely representation of resonance structure of p–nitrophenoxide is:

(A) 

(B) 

(C) 

(D) 

23. Amongst H_2O, H_2S , H_2Se and H_2Te, the one with the highest boiling point is **[JEE 2000]**

(A) H_2O because of hydrogen bonding

(B) H_2Te because of higher molecular weight

(C) H_2S because of hydrogen bonding

(D) H_2Se because of lower molecular weight

24. The hybridization of atomic orbitals of nitrogen in NO_2^+, NO_3^- and NH_4^+ are **[JEE 2000]**

(A) sp^2, sp^3 and sp^2 respectively (B) sp, sp^2 and sp^3 respectively

(C) sp^2, sp and sp^3 respectively (D) sp^2, sp^3 and sp respectively

25. The correct order of hybridization of the central atom in the following species NH_3, $PtCl_4^{-2}$, PCl_5 and BCl_3 is **[JEE 2001]**

(A) dsp^2, sp^3d, sp^2 and sp^3 (B) sp^3 , dsp^2 , sp^3d, sp^2

(C) dsp^2, sp^2, sp^3, sp^3d (D) dsp^2, sp^3, sp^2, sp^3d

26. Specify hybridization of N and B atoms in a 1 : 1 complex of BF_3 and NH_3 **[JEE 2002]**

(A) N : tetrahedral, sp^3 ; B : tetrahedral, sp^3

(B) N : pyramidal, sp^3; B : pyramidal, sp^3

(C) N : pyramidal, sp^3 ; B : planar, sp^2

(D) N : pyramidal, sp^3; B : tetrahedral, sp^3

27. The nodal plane in the π-bond of ethene is located in **[JEE 2002]**

(A) the molecular plane

(B) a plane parallel to the molecular plane

(C) a plane perpendicular to the molecular plane which bisects, the carbon-carbon σ bond at right angle

(D) a plane perpendicular to the molecular plane which contains, the carbon-carbon bond.

28. Which of the following are isoelectronic and isostructural ? $NO_3^-, CO_3^{2-}, ClO_3^-, SO_3$ **[JEE 2003]**

(A) NO_3^-, CO_3^{2-} (B) SO_3, NO_3^- (C) ClO_3^-, CO_3^{2-} (D) CO_3^{2-}, SO_3

29. Which species has the maximum number of lone pair of electrons on the central atom ? **[JEE 2005]**

(A) ClO_3^- (B) XeF_4 (C) SF_4 (D) I_3^-

30. The percentage of p-character in the orbitals forming P–P bonds in P_4 is **[JEE 2007]**

(A) 25 (B) 33 (C) 50 (D) 75

31. The structure of XeO_3 is **[JEE 2007]**

(A) linear (B) planar (C) pyramidal (D) T-shaped

32. **Statement-1** : p-Hydroxybenzoic acid has a lower boiling point than o-hydroxybenzoic acid.

because

Statement-2 : o-Hydroxybenzoic acid has intramolecular hydrogen bonding. **[JEE 2007]**

(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.

33. **Statement-1** : In water, orthoboric acid behaves as a weak monobasic acid.

because

Statement-2 : In water, orthoboric acid acts as a proton donor. **[JEE 2007]**

(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.

34. The nitrogen oxide(s) that contain(s) N–N bond(s) is/are **[JEE 2009]**

(A) N_2O (B) N_2O_3 (C) N_2O_4 (D) N_2O_5

35. The species having pyramidal shape is/are : **[JEE 2010]**

(A) SO_3 (B) BrF_3 (C) SiO_3^{--} (D) OSF_2

36. Which ordering of compounds is according to the decreasing order of the oxidation state of nitrogen-

(A) HNO_3, NO, NH_4Cl, N_2 (B) HNO_3, NO, N_2, NH_4Cl **[JEE 2012]**

(C) HNO_3, NH_4Cl, NO, N_2 (D) NO, HNO_3, NH_4Cl, N_2

37. The shape of XeO_2F_2 molecule is : **[JEE 2012]**

(A) Trigonal bipyramidal (B) Square planar

(C) tetrahedral (D) see-saw

38. The compound(s) with TWO lone pairs of electrons on the central atom is(are) **[JEE 2016]**

(A) BrF_5 (B) ClF_3 (C) XeF_4 (D) SF_4

39. The crystalline form of borax has **[JEE 2016]**

(A) Tetranuclear $[B_4O_5(OH)_4]^{2-}$ unit

(B) All boron atoms in the same plane

(C) Equal number of sp^2 and sp^3 hybridized boron atoms

(D) One terminal hydroxide per boron atom

40. The order of the oxidation state of the phosphorus atom in H_3PO_2, H_3PO_4, H_3PO_3 and $H_4P_2O_6$ is

(A) $H_3PO_4 > H_4P_2O_6 > H_3PO_3 > H_3PO_2$ **[JEE 2017]**

(B) $H_3PO_3 > H_3PO_2 > H_3PO_4 > H_4P_2O_6$

(C) $H_3PO_2 > H_3PO_3 > H_4P_2O_6 > H_3PO_4$

(D) $H_3PO_4 > H_3PO_2 > H_3PO_3 > H_4P_2O_6$

ANSWERS KEY

EXERCISE # O-1

Que.	1	2	3	4	5	6	7	8	9	10
Ans.	C	A	B	A	B	D	D	A	D	D
Que.	11	12	13	14	15	16	17	18	19	20
Ans.	D	C	B	A	C	C	A	A	B	C
Que.	21	22	23	24	25	26	27	28	29	30
Ans.	C	C	A	C	C	C	C	B	B	C
Que.	31	32	33	34						
Ans.	D	D	D	C						

EXERCISE # O-2

Que.	1	2	3	4	5	6	7	8	9	10
Ans.	A, C	A, B	A, B, C	A,B,C,D	B, C	A, C	A, B, C	A, B	A, C	B, C
Que.	11	12	13	14	15	16				
Ans.	B, C, D	A, B, C	B, C, D	B, C, D	A,B,C,D	B, C				

EXERCISE # S-1

Que.	1	2	3	4	5	6	7	8	9	10
Ans.	0	1	0	6	4	3	7 (25%)	3	2	4
Que.	11	12	13	14	15					
Ans.	3	10	3	4	4					

EXERCISE # S-2

Que.	1	2	3	4	5	6	7	8	9	10
Ans.	B	A	D	C	D	B	B	B	C	A
Que.	11				12					
Ans.	(A)→Q,R,S;(B)→Q,T;(C)→R;(D)→P,T				(A)→P,S (B)→P,Q,R,S (C)→P,R (D)→P,Q					
Que.	13					14				
Ans.	(A) → P,R,S (B) → P,R (C) → Q,S (D) → Q,S					(A) → P,R,S (B) → P,Q,S (C) → P,R,S (D) → P,Q,S				
Que.	15	16	17	18	19					
Ans.	B	A	A	C	B					

EXERCISE # JEE-MAIN

Que.	1	2	3	4	5	6	7	8	9	10
Ans.	3	2	1	1	4	3	3	1	4	2
Que.	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
Ans.	4	4	4	2	BONUS	3	3	3	4	1
Que.	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**
Ans.	3	3	2	4	1	3	2	3	2	1
Que.	**31**	**32**	**33**	**34**	**35**	**36**	**37**	**38**	**39**	**40**
Ans.	3	2	2	1	2	2	2	4	4	1
Que.	**41**	**42**	**43**	**44**						
Ans.	4	3	1	4						

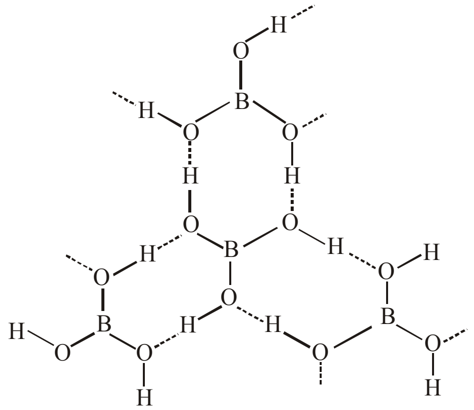
EXERCISE # JEE-ADVANCED

1. Ans. 4
2. Ans. 0
3. Ans. 5
4. Ans. 6
5. Ans. 6

Que.	6	7	8	9	10	11	12	13	14	15
Ans.	B, C, D	C	A	A, B, D	C	B	C	B	C	A
Que.	**16**	**17**	**18**	**19**	**20**	**21**	**22**	**23**	**24**	**25**
Ans.	B	A,B,C	B	D	A	B	A	A	B	B
Que.	**26**	**27**	**28**	**29**	**30**	**31**	**32**	**33**	**34**	**35**
Ans.	A	A	A	D	D	C	D	C	A,B,C or A,C	D
Que.	**36**	**37**	**38**	**39**	**40**					
Ans.	B	D	**B,C**	**A,C,D**	**A**					

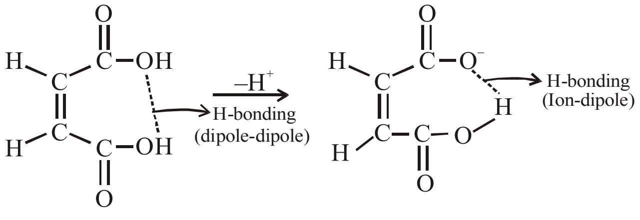
ANSWER OF HYDROGEN BONDING

1.

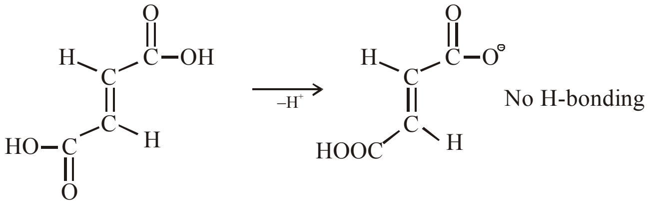


Structure of boric acid; the dotted lines represent hydrogen bonds

2. Inter molecular H-Bonding in case of para nitro phenol is strong enough to increase boiling point which is not favoured in o-Nitro phenol.
3. Maleic acid anion gets stabilize due to intra molecular H-bonding which is not possible in anion of Fumaric acid. So H^+ releasing tendency is more in case of maleic acid

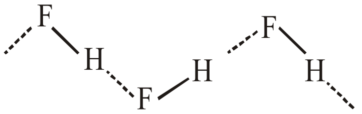


Maleic acid

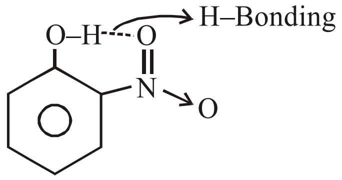


Fumaric acid

4. Due to strong H-Bonding in H–F



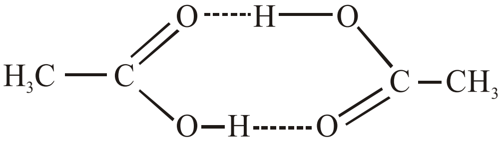
5. Due to H-Bonding in NH_3.
6. Density of ice is higher than that of water due to cage like structure of ice volume of ice increases and density decreases.
7. On heating ice in temperature range 0–4°C H-bonds of ice break down, which decreases volume, hence density of H_2O increases. On heating ice after 4°C thermal vibrations of H_2O molecules increases which increases effective volume of ice, so density of ice again decreases.
8. H-I bond is weak as compare to H-F so it can be dissociated easily and can give H^+ easily.
9. The organic matter in wood participate in H-bonding with water molecules in ice-cream.
10. $K^+[F-H\underset{\underset{\text{H-bonding}}{\downarrow}}{\cdots\cdots}F^-]$ But $KHBr_2$ & KHI_2 can't form H-Bond.
11. In o-Nitrophenol intra molecular H-bond is present which decreases it's solubility in water.



Ortho Nitro-phenol

12. In o-hydroxy benzeldehyde intramolecular H-bond is present which is less stronger than intermolecular H-bond in p-hydroxy benzeldehyde.
13. Extent of H-bonding in glycerol is more as compare to ethanol so it is more viscous.
14. Due to H-bonding in H_2O

15. Due to H-bonding in acetic acid it forms dimer so overall molecular weight is just double of original



Dimer of CH_3COOH

16. Due to bigger size of 'Cl' atom it's interaction is not enough to evolve the aomount of energy which lies in the range of H-bond.

17. $\underset{\text{liquid}}{(AcOH)_n} \xrightarrow{\text{on vapour}} \frac{n}{2}(AcOH)_2$

} Not completely converted

But in $\underset{\text{liquid}}{(H_2O)_n} \xrightarrow{\text{on vapour}} n(H_2O)$

Here all H_2O molecule gets vapourised. So entropy change is more

18. As extent of H-bond is more in H_2O as compare to HF, heat of vaporisation of water is higher than HF.